

Lab 3: Equilibria

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Introduction

In most chemical reactions, a dynamic state is reached between products and reactants where products are formed continuously while reactants are formed at the same time. This state is referred to as equilibrium and is signified by the symbol (\rightleftharpoons) in a chemical reaction to show the production of products and reactants at the same time, or. Are being produced in a reversible reaction. The rate of formation for products and reactants does not have to be the same, and are, in most cases, have a favored side where one side of the reaction process produces more product or reactant depending on the chemical reaction. To determine which side will be favored in a chemical reaction, the equilibrium constant expression is used:

This expression uses an example of a chemical reaction, ($aA + bB \rightleftharpoons cC + dD + \text{heat}$) and allows us to mathematically predict and determine which side of the reaction will be favored. This is especially important when trying to acquire the most amount of product or yield in certain industries to maintain high efficiency.

When a reaction reaches equilibrium, certain factors come to play that may cause a shift in the reaction's equilibrium. Equilibrium shifts can be caused by a change in temperature or pressure in the environment, or by adding or removing products or reactants in a reaction. To monitor such shifts, *Le Chatelier's principle* is observed which states that a reaction will shift to the direction that will allow for the least effect change on the entire system. This principle ensures that a chemical reaction will always try to return to equilibrium. For example, If more products are added to the reaction. then the equilibrium will shift towards the reactions to compensate for the additional product and vice versa. For some reactions, one side can be heavily favored, and this may be caused due to a strong acid or base that is involved. To determine whether an acid or base is strong, you observe its ability to dissociate in a reaction completely and use the acid dissociation constant to determine the acidity strength. pH is particularly important in this case, as it indicates the acidity level of a reaction or element using the pH scale. To determine this, any value under 7 in the pH scale is considered acidic, where anything above 7 is considered a base. An expression of $pH = -\log [H^+ (aq)]$ can be used to determine where an ion or molecule lies in such scale.

A popular example of where equilibrium is achieved and is vital is human body. Many of its processes stresses the importance of having equilibrium maintained. The body tries to keep a pH level between 7.35 and 7.45, but is challenged frequently by what we consume, the temperature, pressure and our breathing habits. The body uses chemical buffer systems to keep the pH of blood at equilibrium, where carbonic acid and hydrogen carbonate (bicarbonate) alternate in production to maintain equilibrium. The respiratory system plays a part in removing or adding carbon dioxide levels which affects the pH level in the body and ultimately equilibrium. The amount of carbon dioxide that each person exhales isn't constant and the level varies depending on whether a person is exercising is breathing normally. However, if a person is breathing too rapidly and hyperventilating, they are releasing too much carbon dioxide and as the Le Chatelier's principle suggests, the body will combat that by shifting towards the production of more carbon dioxide using carbonic acid. If a person is hypoventilating, carbon dioxide is removed through producing more H_3O^+ . All these processes revolve around equilibrium and the and applying it's principle to showcase its importance in our bodies and the real world.

- Reactions being studied in the case of this experiment:

I	$[Cu(H_2O)_4]^{2+} (aq) + 4 NH_3 (aq) \rightleftharpoons [Cu(NH_3)_4]^{2+} (aq) + 4 H_2O (l)$
II	$2AgNO_3 (aq) + Na_2CO_3 (aq) \rightleftharpoons Ag_2CO_3 (s) + 2 NaNO_3 (aq)$
III	$2 H^+ (aq) + CO_3^{2-} (aq) \rightleftharpoons H_2CO_3 (aq) H_2O (l) + CO_2 (g)$
IV	$Ag^+ (aq) + Cl^- (aq) \rightleftharpoons AgCl (s)$
V	$Ag^+ (aq) + 2NH_3 (aq) \rightleftharpoons [Ag(NH_3)_2]^+ (aq)$
VI	$H^+ (aq) + NH_3 (aq) \rightleftharpoons NH_4^+ (aq)$
VII	$Ag^+ (aq) + I^- (aq) \rightleftharpoons AgI (s)$
VII	$2Ag^+ (aq) + S^{2-} (aq) \rightleftharpoons Ag_2S (s)$
I	
IX	$2 H_2O (l) + CO_2 (g) \rightleftharpoons H_2CO_3 (aq) + H_2O (l) \rightleftharpoons HCO_3^- (aq) + H_3O^+ (aq)$

Procedure

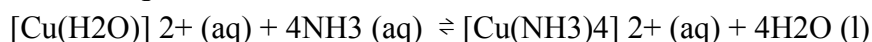
As described in the lab manual ("VIOLENT FIRES SOON BURN OUT THEMSELVES, SMALL SHOW'RS LAST LONG, BUT SUDDEN STORMS ARE SHORT" EQUILIBRIA, Dr. Rashmi Venkateswaran, Exp. 3).

Equipment

As described in the lab manual ("VIOLENT FIRES SOON BURN OUT THEMSELVES, SMALL SHOW'RS LAST LONG, BUT SUDDEN STORMS ARE SHORT" EQUILIBRIA, Dr. Rashmi Venkateswaran, Exp. 3).

Observations and Discussion

Part 1: Equilibrium Shift



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

CuSO_4 is a transparent, light blue liquid. The copper precipitate within the solution is what caused a light blue appearance of the solution.

B. What changes do you observe? From the chemical reactions provided, what chemistry explains the observations you made? What ions cause the colour?

After only 1 drop of NH_3 was added to the CuSO_4 the solution turned a dark blue and intensified with each drop following. In order to maintain equilibrium, a shift to the right occurred. This consumed the excess NH_3 ions.

C. 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?

After 25 drops of HCl was added, the solution returned to light blue colour seen previously. This is a result of the H^+ ions being increased.

1. 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?

When step 4 was performed, the same changes were seen. I would expect them to be same because the same shifts were made and the same solutions were added, having the same effect on the overall observations.

Part 2: Multiple Equilibria

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

Na₂CO₃ is a colourless, transparent liquid. AgNO₃ is a colourless, transparent liquid. HNO₃ is also a colourless, transparent liquid.

E. What changes do you observe? From the chemical reactions provided, what chemistry explains the observations you made? What ions cause this colour?

When 10 drops of AgNO₃ was added, the solution turned brown and cloudy. The observation can be explained with the chemical reaction $2\text{AgNO}_3(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightleftharpoons \text{Ag}_2\text{CO}_3(\text{s}) + 2\text{NaNO}_3(\text{aq})$. The colour change is a result of the precipitate formed during the reaction. The equilibrium is shifting to the right. A double displacement occurred during the reaction, producing two new compounds, comprising of a solid and an aqueous solution.

F. 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?

5 drops of HNO₃ was added, the solution returned to a colourless, transparent liquid. The observation can be explained using the chemical reaction $2\text{H}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightleftharpoons \text{H}_2\text{CO}_3(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2$. The increase of H⁺ ions cause a reaction between the solid and production of excess product. Therefore, the equilibrium shifted to the left. The high concentration of H⁺ caused the products to reverse resulting in the observation of the solution returning to a colourless, transparent liquid.

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

When HCl was added, the solution turned white and cloudy. The observation can be explained using the chemical reaction $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightleftharpoons \text{AgCl}(\text{s})$. The equilibrium shifted to the right as a precipitate was formed. The reaction shows the product AgCl, a solid.

H. 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?

17 drops of NH₃ caused the solution to return colourless, transparent solution. The observation can be explained using the chemical reaction $\text{Ag}^+(\text{aq}) + 2\text{NH}_3(\text{aq}) \rightleftharpoons [\text{Ag}(\text{NH}_3)_2]^+(\text{aq})$. The NH₃ dissolved AgCl (seen above), forming $[\text{Ag}(\text{NH}_3)_2]^+(\text{aq})$. Equilibrium was shifted to the left.

2. Can you observe the same changes if you repeat steps 7 & 9?

No, the same changes were not observed when steps 7 and 9 were repeated. This is a result of step 8 being skipped, the HCl was not added into the solution.

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

When KI was added the solution turned a yellow/cream colour. The reaction that can be used to explain this observation is $\text{Ag}^+ (\text{aq}) + \text{I}^- (\text{aq}) \rightleftharpoons \text{AgI} (\text{s})$. Equilibrium was shifted to the right as the precipitate ($\text{AgI} (\text{s})$) was formed.

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

When Na_2S was added to the solution, it turned a dark brown/grey and was room temperature. The reaction that can be used to explain this observation is $2 \text{Ag}^+ (\text{aq}) + \text{S}^{2-} (\text{aq}) \rightleftharpoons \text{Ag}_2\text{S} (\text{s})$.

Part 3: Buffers

K. 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 pH of the water was 6.16. I would expect the water to be closer to 7, a more neutral pH. The pH could be different than what was expected because the lab required distilled water opposite to tap water. The pH level of distilled water has a slightly more acidic pH.

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

The pH observed was 8.22, concluding it had increased from the initial pH. The ions responsible for this are H^+ ions. The low concentration of H^+ ions increased the pH of the solution.

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

When HCl was added, the solution became neutralized. The acid was able to break apart the hydrogen carbonate ions to create water and carbon dioxide.

N. 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 based on the theory?

When HCl was added, bubbling started to occur. The bubbling is a result of the production of CO₂. The pH was recorded to be 7.02. This is higher than the pH of the distilled water but lower than the pH of blood, which is normally around 7.35. The species present in the solution are H₂O, CO₂ and NaCl. Based on the observations recorded, there seems to be higher species of CO₂ present due to the occurrence of bubbles.

O. 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.

When 10ml of 0.85% lactic acid was added, the pH increased to 7.07. Lactic acid is formed in the body when muscles don't get enough oxygen. A large amount of lactic acid is more the result of intense exercise. If the addition of lactic acid exceeds your body's clearance, your blood pH decreases. Therefore, the pH in this case, should have decreased instead of increasing.

P. 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.

As the stirring rate was increase, a rise in bubbles produced by CO₂ was observed. Therefore, CO₂ is being expelled. The increased rate of respiration results in more CO₂ expelled has the overall effect of an increase pH. This explains the observation of the rate of stirring was increased from 8 to 10 resulting in the pH increase of 7.07.

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

When 0.50g of sodium bicarbonate was added, the pH was increased to 7.19. Sodium bicarbonate occurs naturally in the blood. Generally, it will act as a buffer, meaning it helps to stabilize the blood. In this case, it increased the pH to a more neutral state.

R. 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?

When another 0.50g of sodium bicarbonate was added, the pH was further increased to 7.27. Maining the stability of the blood pH is important. By adding more sodium bicarbonate, the blood pH becomes more stable attempting the reach the normal pH level of 7.35.

S. 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.

When a pellet of CO₂ (dry ice) was added, the pH was decreased to 7.26. When CO₂ is not exhaled it begins to build up in your blood. This leads to acidosis. By adding a pellet of dry ice, the limit of exhalation of CO₂ was simulated. This resulted in a decrease in the pH.

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

When the final step of adding 0.40g of NH₄Cl was complete, the pH was increased to 7.26. NH₄Cl is formed by reaction of HCl with NH₄OH. HCl is strong acid and NH₄OH is a weak base, so their salt is a weak acid. NH₄Cl is normally injected into patients to lower the pH level. In this case, the pH was increased.

End of Report Questions

A. Ammonium ion is used as the acid source instead of hydrochloric acid as the ammonium ion will dissociate into H⁺ and NH₃ ions. Using Hydrochloric acid would not dissociate into OH⁻ ions which will completely lower the pH level to levels more than the body ideally allows. Ammonium uses the NH₃ ions as a base to maintain an equilibrium where the blood pH lowers but does not drop significantly by countering it with a NH₃ base.

B. NH₄Cl is used instead of other ammonium compounds such as (NH₄NO₃ or NH₄I) because it contains a chlorine. This ion will combine with excess H⁺ and create HCl, which will lower the pH level of the blood.

2)

Acid is added = pH level decreases. Addition of an acidic substance causes the pH to go down. This is because H⁺ ions are present and they are responsible for lowering the pH level.

3) Acidosis occurs when the pH of the blood contains significant amount of acidity levels, making it lower than the pH of 7. As many H⁺ ions are present, they begin to bond with the hemoglobin and oxygen is as a result inhibited to bind with the hemoglobin. This affects the rate and oxygen level that is transported in the body, making it much lower and less effective.

4) The solution on the left is represented with a bright red color, which is the expected colour of blood normally. The solution on the left created a darker shade, and the only difference between both solutions was HCl was added to the latter. This allows us to conclude the HCl darkened the solution by increasing the acidity. The excess H^+ might have reacted with the hemoglobin and oxygen in blood, which created less oxygen as H^+ binded to the hemoglobin. The increase of acidity and less oxygen resulted in a colour change in the solution.

5) Soda contains carbonic acid which makes the acidity level high and the pH low. Stirring such soda will dissociate the carbon dioxide in the solution, and this will affect the carbonic acid levels. Removing some of the cause of the acidity will cause an equilibrium shift, and as *Le Chatelier's principle* tells us, the system will try to minimize the effect and that will create an equilibrium with less acidity, raising the pH level of the soda after stirring.

6) The painting of the chicken resembles humans who are hyperventilating, in which they are breathing rapidly and too much carbon dioxide is leaving the body. As a result, the pH level is increased due to less production of carbonic acid and bicarbonate acid buffer, and equilibrium is achieved with higher pH to compensate. Less carbon dioxide results in less carbonate to combine with calcium ions to make calcium carbonate, which affects the egg shells by having less calcium within them. This might make the egg shell fragile and not strong enough.

Conclusion

Equilibrium is crucial to life around us, and allows us to explain the nature of chemical reactions in our real world. It explores the concept of maintaining balance within reactant and products, which ultimately affects the way we operate in our jobs and many industries. It uses principles such as *Le Chatelier's principle* to explain how ions operate in reactions and to explain how our human bodies need equilibrium to function in an ideal manner. Addition of reactant or product, and change in temperature and pressure could all be explained using equilibrium and it's expressions, and shows how the equilibrium shifts to compensate for such changes.

Work Cited/ References

- ("VIOLENT FIRES SOON BURN OUT THEMSELVES, SMALL SHOW'RS LAST LONG, BUT SUDDEN STORMS ARE SHORT" EQUILIBRIA, Dr. Rashmi Venkateswaran, Exp. 3)
- (Orenda Technologies, <http://blog.orendatech.com/total-alkalinity-role-water-chemistry>, October 5 2016)

