Reversable Reactions and Chemical Equilibria

Partner: Gina Dazzo

Purpose:

- **Study 5 different reversible reactions and their equilibrium**
- **Observe the effect of Le Chatelier's principle and the common ion effect**
- **Observe the formation of a precipitate in over saturated solution and the conditions that forms the precipitate**
- Reason out an explanation when an unexpected result occurs in an experiment

Reference:

(1) Kateley, L. J., *Introduction to Chemistry in the Laboratory*, *20th Ed.*, Lake Forest College, **2021,** Experiment 6_Reversable_Reactions, Appendix

Weak Acid Equilibrium

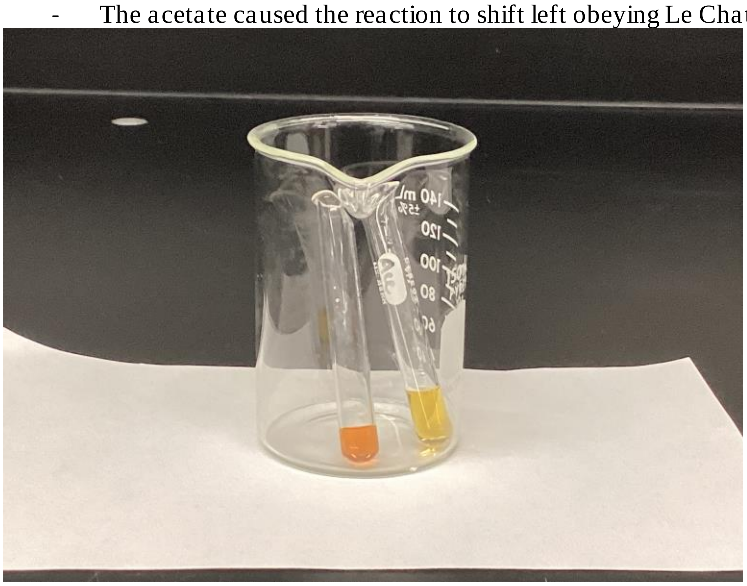
$CH_3COOH(aq) \leftrightarrows CH_3COO^-(aq) + H^+(aq)$

Acetic acid $\leftrightarrows$ acetate ion + hydrogen ion/proton

- Acetic acid (CH_3COOH) of 0.1M is a clear colorless liquid
- Orange methyl orange indicator added
- The orange gold color indicated that the pH is between 3 and 4
- 5 drops of 1.0M sodium acetate (CH3COONa) was added
- Acetate ions/ $CH3COO^-$ are added and the solution turned the solution golden yellow
- The golden yellow color indicated that the pH is greater than 4.4

Conclusion #1:

- Equilibrium shifted left
- The acetate caused the reaction to shift left obeying Le Chatelier's principle (the common ion effect)



Weak Base Reaction

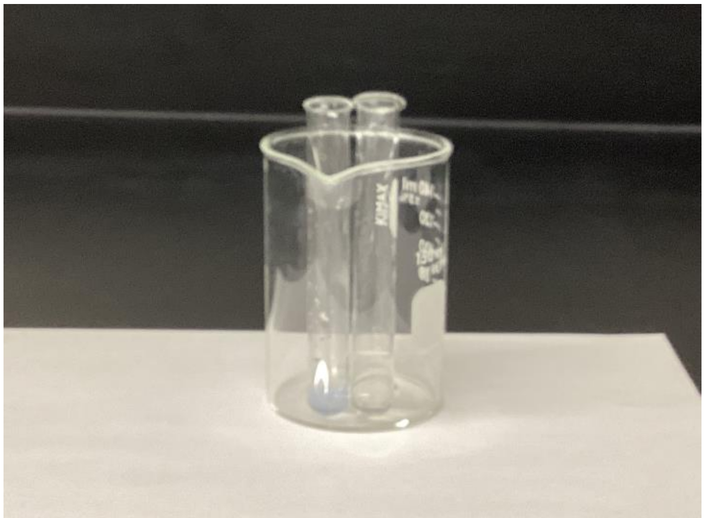
$NH_3(aq) + H_2O(l) \leftrightarrows NH_4^+ + OH^-$

Ammonia + water $\leftrightarrows$ ammonium + hydroxide

- 10 drops of 0.1M ammonia is a clear colorless solution
- 2 drop of Thymolphthalein indicator added
- The lack of color change indicated that the pH is below 9.3
- 10 drops of ammonium chloride (NH_4Cl) was added Molarity???

- Ammonium ions/ NH_4^+ are added and the solution did not change color
- The lack of change in color indicated that the pH is still below 9.3

Conclusion #2:
- Based on Le Chatelier's principle the reaction should have shifted left and the pH should have increased turning the solution blue
- The equilibrium did not shift enough to change the pH
- The NH_4^+ was not concentrated enough or the indicator was not working

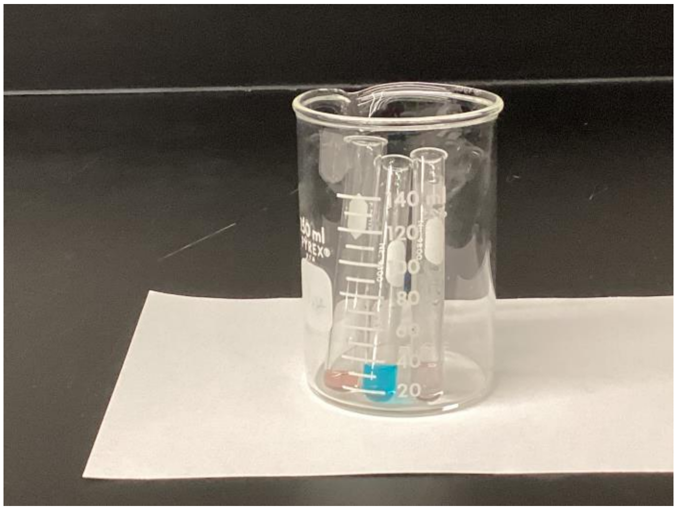


Complex Ion Equilibrium Reaction

$Co(H_2O)_6^{2+}(aq) + 4Cl^-(aq) \leftrightarrows CoCl_4^{2-}(aq) + 6H_2O(l)$
Cobalt chloride + chlorine ions ⇆ tetrachlorocobaltate + water

- 0.1M cobalt chloride is a clear orange pink solution
- 5 drops of 12M hydrochloric acid from the hood created a layer of blue solution on the bottom of the test tube
- When shaken the solution turned clear
- 2 more drops added turning the solution violet
- 5 more drops added turning the solution an intense blue
- The equilibrium shifted right
- Adding water reversed the solution back to pink
- The equilibrium was reversed and shifted left

Conclusion #3:
- The equilibrium shifted right when hydrochloric acid was added and left when water was added
- The hydrochloric acid caused the reaction to shift left obeying Le Chatelier's principle (the common ion effect)
- The water caused the acid to shift left obeying Le Chatelier's principle



Equilibrium of a Saturated Solution of Sodium Chloride

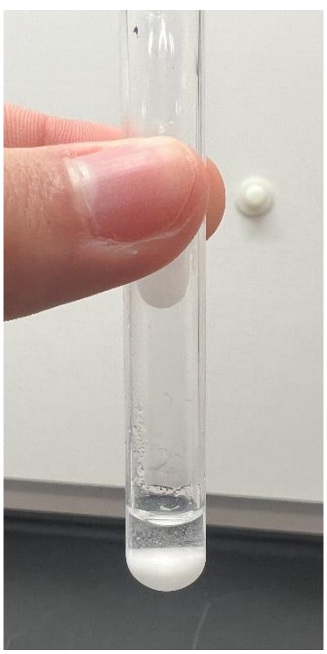
$Na^{+}(aq) + Cl^{-}(aq) \leftrightarrows NaCl(s)$
Sodium ions + chlorine ions $\leftrightarrows$ sodium chloride precipitate

- 10 drops of 5.41M clear colorless NaCl solution were added to a test tube
- A few drops of clear colorless 12M HCl added creating a white precipitate of NaCl

Calculations:
1358g (1mL/1.200g)(1x10-3L/1mL) = 1.132L
358g(1mol/58.5g) = 6.12mol
6.12mol/1.132L = 5.41M of saturated NaCl

Conclusions #4:

- The addition of HCl oversaturated the solution creating a physical change
- The oversaturation of Cl^{-} ions caused NaCl to form and be knocked out of the solution forming a precipitate, right shift?



Simultaneous Chemical Equilibria and Maximizing Precipitate

Precipitation of Calcium Oxalate
$Ca^{2+}(aq) + C_2O_4^{2-}(aq) \leftrightarrows Ca_2O_4(s)$
Calcium ion + oxalate ion ⇆ calcium oxalate

- This reaction was optimized to maximize precipitate

Ionization of Oxalic Acid EQ2
$H_2\ C_2O_4(s) \leftrightarrows 2H^+(aq) + C_2O_4^{2-}$
Oxalic acid ⇆ hydrogen ions + oxalate ions

Complete Ionization of a Soluble Salt
$(NH_4)_2C_2O_4(aq) + H_2O(l) \rightarrow 2NH_4^+(aq) + C_2O_4^{2-}$
Ammonium oxalate + water ⇆ ammonium ions + oxalate ions

- 5 drops of 0.1M clear colorless calcium chlorate were added to 2 test tubes
- 5 drops of 0.5M clear colorless ammonium oxalate is added to one test tube (Tube #1)
- The solution turned cloudy white
- 5 drops of clear colorless 0.5M oxalic acid is added to the other test tube (Tube #2)
- The solution formed a cloudy white precipitate at the bottom of the test tube
- 5 drops of 5M hydrochloric acid were added to Tube #2
- The cloudiness of the solution cleared and it returned to a clear colorless solution

Conclusion #5:

- Ammonium oxalate is a better choice to precipitate calcium oxalate
- Adding HCl shifts the equations, Eq2 to the left following Le Chatelier's principle,