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Analytical Chemistry
Acid-Base Equilibrium and Buffers

Purpose: The purpose of this experiment was to find the pH value of 8 different aqueous solutions using indicators. As well as constructing a buffer in order to find the pH value of an unknown solution. This is done using colorimetric analysis and using the HH equation.

Procedure:

Part 1:

You will be working with nine aqueous solutions (table 3) and your unknown solution for this session. The equilibria, equilibrium constants, and the pKs provided should be used in all of the calculations. All solutions are 0.100 M.

Part 2:

1. In each of the 7 test tubes squirt an eyedropper full of 0.100 M HCH3O.
2. Add 2-3 drops of the seven indicators, one of the each tube.
3. Dispose of the contents of the tubes down the sink and rinse with water. They don't have to be dry.

Part 3:

1. Clean 9 small test tubes; they do not need to be dry. Place a piece of paper beneath the test tube rack so that you can easily label the pH for each tube and the tubes that represent the unknown.
2. Using the eyedropper squirt a dropper full of each of the 9 available buffers to the 8 test tubes. Try to have roughly the same level of liquid in each tube. Add your unknown to the 9th test tube to the same height as the others.
3. Add 2-3 drops of your selected indicator to each tube.
4. Compare the color of your unknown solution with each of the buffers.
5. Dispose of the contents of the test tubes in the sink and rinse.

Part 4:

1. Your instructor will demonstrate how to calibrate and use a pH meter and will have scintillation vials of each of the 9 aqueous solutions from Lab 6 set up for you to use.
2. After calibrating the pH with the standardized buffer solutions (pH 4, 7, 10).
3. Take a look at the calculated pH and indicator pH range you got for each of the solutions last week.
4. Next measure the pH of your unknown.
5. Does your pH meter data for your unknown make sense based on the data you already have.
6. Rinse the electrode with distilled water.

Part 5:

1. Using the method above calculate amounts of the solution necessary to prepare 50.0 of the assigned buffer. (9)
2. Using a graduated cylinder, obtain sufficient volumes of your chosen pair to prepare 50 mL of buffer.
3. Using a separate graduated pipet for each solution add the appropriate amounts of the two components of the buffer.
4. Calibrate a pH meter using the three calibration solutions (pH 4, 7, 10).
5. Using calibrated pH meter determine the pH of your buffer and prevent cross-contamination.
6. Using the calibrated pH meter determine the pH of your buffer and record the result in your lab notebook.
7. If your buffer is at the desired pH, great; you're done.

Data Section:

Table 3:

| | HC2H3O2 | C2H3O2^- | H3PO4 | H3PO4^- | HPO4^2- | PO4^3- | NH3 | NH4^+ | NH4C2H3O2 | | |
|---|---|---|---|---|---|---|---|---|---|---|---|
| TB | Yellow (3-7) | Yellow (3-7) | Red 0 | (3-7) | Red (0) | Blue (10-12) | Blue (10-12) | Yellow (3-7) | Yellow(3-7) | Yellow(3-7) | |
| BB | Yellow (1-6) | Blue(5-12) | Yellow (0-2) | Blueish Purple(5-12) | Yellow (0-7) | Blue (5-12) | Blue (5-12) | Blue (5-12) | Blue (5-12) | Blue (5-12) | |
| BG | Orange (1-2) | Blue (6-12) | Yellow (0-2) | Blue (6-12) | Yellow (0-1) | Blue (6-12) | Blue (6-12) | Blue (6-12) | Blue (7-12) | Blue (6-12) | |
| BP | Green (3-5) | Blue (7-12) | Yellow (0-4) | Yellow (0-4) | Yellow (0-4) | Purple(7-12) | Purple(7-12) | Purple(7-12) | Purple (7-12) | Purple(7-12) | |
| BB | Yellow (0-5) | Blue (8-12) | Yellow (0-5) | Purple (6-7) | Yellow (0-5) | Purple (8-12) | Purple (8-12) | Purple (8-12) | Purple (6-7) | Purple (6-7) | |
| CR | Colorless (2-6) | Blue (9-12) | Red 1,7-8 | Yellow (0-6) | Colorless (1-7) | Red (9-12) | Red (9-12) | Red (2-6) | Red (2-6) | Red (2-6) | |
| P | Colorless (0-7) | Colorless(0-7) | Colorless (0-7) | Colorless (0-7) | Colorless (0-7) | Pink (8-9) | Pink 9 | Colorless(0-7) | Colorless 0-7 | Colorless 0-7 | |
| PH | 2.88 | 3.74 | 1.62 | 4.1 | 10.1 | 12.97 | 11.1 | 5.13 | 7 | N/A | |

| PH Range | (0-7) | (0-7) | (0-7) | (0-7) | (0-7) | (5-12) | (5-12) | (5-12) | (0-7) | (5-12) | |
|---|---|---|---|---|---|---|---|---|---|---|---|
| PH Meter | 2.63 | 7.62 | 1.05 | 4.19 | 8.62 | 11.81 | 10 | 6.7 | 7.37 | 7.12 | |
| | | | | | | | | | | | |

Table 4: Standard Buffers and Colorimetric Analysis Data

Indicator Used in Experimentation: HP04^2-

| pH | Color |
|---|---|
| 6.00 | Colorless |
| 6.40 | Colorless |
| 6.80 | Colorless |
| 7.20 | Colorless |
| 7.60 | Colorless |
| 8.00 | Colorless |
| 8.40 | Colorless |
| 8.80 | Colorless |
| Unknown (#) | B |

Estimated unknown pH based on Color: ______7___________

Results and Conclusions:

- The goal of this experiment was to determine the weak acids and bases pH using pH indicators, standard buffer and pH meter. Another goal of this experiment was to find the pH of an unknown aqueous sample.
- The technique of the experiment was rather well. The results of the pH matched our predictions for each of the aqueous solutions. We made sure we added the right amount aqueous solution and resetting the pH meter.
- The sample pH's were for the most part correct. The HPO4^-2 aqueous solution pH didn't match up for our theoretical pH value but it wasn't too off. As well as our theoretical pH value for PO4^3- which was off but was still able to identify it as an acid because of the theoretical pH calculated and wasn't off by too much.
- Using the different indicators we did a colorimetric analysis to narrow down the pH. Since the color appears in a certain pH range using multiple indicators can help narrow down the possible pH values.
- There is a difference in value for the experimental and calculated pHs. However while the difference is slight this is because the theoretical calculated value may be different from the experiment because not everything reacted since the calculated value assumes that it is a 100 percent reactive reaction which often it isn't.
- The process in which the pH is determined is called colorimetry. Since we used a pH value that is known which was the buffer. We added an indicator which we can compare the colors to the unknown with the same indicator and narrow down the pH estimation.
- Our best estimate for the pH of the unknown solution is between 6-7 since that is the narrowest pH range from our indicator tests using colorimetric analysis. Since the Bromothymol Blue indicator changed the solution to a purple color which means the pH must be within the range of 6-7.
- The solutions that we decided to go with was HPO4^2-1 and H2PO4 this is because this was the closest pH value to our unknown value since that would create a buffer around 7.21. Using the Henderson Hasselbalch equation we find the pKa which is the closest to our estimated value for our unknown. Using M1V1 = M2V2 we can find the volume needed to reach the pH of our unknown. We were lower and did not meet the 7.21 and thus had to add more acid to our solution to adjust the pH. It was 7.12 so there was a absolute difference of 0.09.
- During the duration of the experiment it was observed that the pH meter would be calibrated with different amounts of the buffer and thus may produce random error in the experiment. This can affect the pH values of the experiment.
- One way to improve the experiment is to ensure that 100 percent of the solution has reacted. This would make our actual experimental pH values closer to the theoretical pHs calculated. As well as making sure that the calibration of the pH meter is done automatically rather than by hand reducing random error in the data that can affect pH values.
- I have learned about the HH equation which can be used to create a buffer and determine the amount of volume a solution to create a particular pH. As well as using colorimetry with 8 different indicators to determine pH values. One of the major conclusions for the experiment.

Sample Calculations:

1) ka*kb = kw

1.76*10^-5 * ko = 1.0 * 10^-14
Kb = 5.68 * 10^-10

2) 5.64*10^-10 = (x)(x)/(.1)-x
7.54*10^-6
pH = -log(oH)
pH = 3.73

3) 7.52 * 10^-3 = (x)^2/.1-x
X = .0239
pH = -log(2.39x10^-2)
pH = 1.62

4) 6.23*10^-8=(x^2)/(1-x)
x= 7.89x10^-5
pH = 4.10

5) (x^2)/(.1-x) = 1.61 * 10^-1
x= 1.27 * 10^-7
p[oH] = 10.11

PO4^3- + H2O -> HPO4^2- + OH-
Ka = 2.2*10^-3
Kb = .045
X = .093
pOH = -log(OH-)
pOH = 1.03
pH = 12.47

6) 1.79*10^-5 = x^2/.1-x
X = 1.34*10^-3
pOH = -log(1.34*10^-3)
pH = 11.1

7) NH3
Kb = 1.79*10^-5
pKb = 4.75
1.79^-5 = (x^2)/(.1-x)

8)
kb = 1.79*10^-5
Ka = 5.59 * 10 ^-10
5.59*10^-10 = (x^2)/(.100)
X = 7.48*10^-6
pH = -log(7.4*10^-1)
pH = 5.13

9)
pH of NH4 = 5.8
pH of nitrate = 8.8
(5.13 + 8.8)/2=7

Unknown Solution:

[HPO4^-2]/[H2PO4]=10^0.79
61.66 + H2PO4 = HPO4^3-
50 = 61.66z
X = 49.186 mL