Experiment #2: Acids and Bases

Learning Objectives:

- Demonstrate how to safely handle strong acids/ bases.
- Demonstrate how to properly weigh solids by difference.
- Demonstrate how to accurately use and measure the volume of a solution with a buret and volumetric pipet.
- Demonstrate how to precisely standardize a solution.
- Demonstrate how to precisely determine the concentration of an analyte in a solution.
- Demonstrate how to assess the quality of your data through error analysis.

Introduction:

In Bronsted–Lowry theory, a strong acid is a species that dissociates 100% to produce H^+(aq) or H_3O^+ (aq). Examples are: HCl, HNO_3, H_2SO_4, HBr and HI
A strong base will dissociate 100% to produce OH^- (aq). Examples are KOH and NaOH.
pH is a measure of the H_3O^+ (aq) concentration: pH = – log [H_3O^+], where [ ] represents concentration in units of mol/L

At 25 °C, when pH = 7, the sample is neutral, and [H_3O^+] = [OH^-]
pH < 7, the sample is acidic and [H_3O^+] > [OH^-]
pH > 7, the sample is basic and [H_3O^+] < [OH^-]
pH is often measured with a pH meter.

Water can vary in pH, depending on its source:
Tap water is regulated by the Environmental Protection Agency (EPA) and typically has a pH ~ *ca.*7.5
Water that has gone through distilled reverse osmosis has a pH = 5 – 7
A modern health fad is alkaline water which has a pH = 9 – 9.5
Common bottled water has a pH range of 6.5 – 7.5
The pH of globally averaged ocean water is ~8 and acid rain has a pH ~5 – 5.5

Naturally occurring rainwater is always slightly acidic because CO_2 (g) in the atmosphere dissolves in water and produces carbonic acid.
CO_2 (g) + H_2O (l) $\leftrightarrow$ H_2CO_3 (aq)
H_2CO_3 (aq) + H_2O(l) $\leftrightarrow$ H_3O^+ (aq) + HCO_3^- (aq)
The double headed arrows, $\leftrightarrow$, represent a reaction in equilibrium, in which the forward and reverse reactions are occurring simultaneously.

When coal or wood is burned, or a volcano erupts, sulfur and nitrogen oxides are released into the atmosphere. These compounds react with water to produce nitric acid (HNO_3) and sulfuric acid (H_2SO_4), which then acidifies rain.

To determine the concentration (mol/L) of an acid or base in a sample, a titration can be performed based upon an acid – base reaction.

Figure 1: The buret contains the titrant, NaOH. The concentration of the tirant needs to be accurately determined through a standardization. The Erlenmeyer flask contains the analyte (HCl) and a visual indicator. The goal is to add titrant until the equivalence point (moles NaOH = moles HCl) is reached. At the equivalence point, the indicator changes color.

Consider the titration of HCl with NaOH, the reaction is:
$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$
Since HCl is a strong acid, NaOH is a strong base, and NaCl is ionic, when in solution, they dissociate into their corresponding ions:
$H^+\ (aq) + Cl^-\ (aq) + Na^+\ (aq) + OH^-\ (aq) \rightarrow Na^+\ (aq) + Cl^-\ (aq) + H_2O\ (l)$
The net ionic equation is $H^+\ (aq) + OH^-\ (aq) \rightarrow H_2O\ (l)$ since $Na^+\ (aq)$ and $Cl^-\ (aq)$ are spectator ions. When moles NaOH = moles HCl, all of the analyte is consumed and the equivalence point is reached.

If NaOH is standardized, its concentration is precisely known. From the concentration of NaOH (mol/L) and the volume used (L) to reach the equivalence point, we can find the moles of NaOH. Since NaOH and HCl react in a 1:1 mole ratio, moles NaOH = moles HCl. We will know the volume of the analyte (HCl) we placed in the Erlenmeyer flask. Molarity of HCl = moles HCl / volume (L).

In order to know when to stop adding titrant from the buret, a visual indicator is added which changes color (endpoint) at a pH closest to the equivalent point. For the NaOH (strong base) and HCl (strong acid) titration, the equivalence point occurs at a pH = 7 since only water is present in the solution. For a strong acid/strong base titration, phenolphthalein is often used.

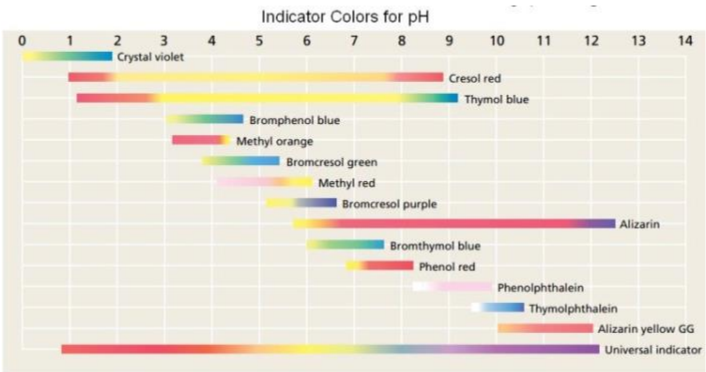

Figure 2: Indicator chart showing the colors of indicators as a function of pH.

In order to determine the concentration of the analyte, the concentration of the titrant in the buret must be precisely known. This is often accomplished by titrating the titrant against a primary standard. Primary standards have the following characteristics:

1. High molecular or equivalent weights
2. Available in high purity
3. Stable, non hygroscopic (does not readily absorb water)
4. Readily available and inexpensive.

For the standardization of NaOH, we will use KHP (potassium hydrogen phthalate):

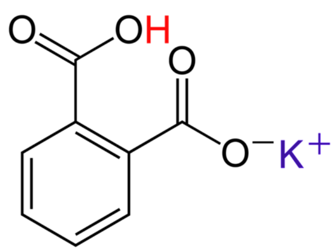

Figure 3: The structural formula for KHP (molecular weight = 204. 223 g/mol)

$$KHP + NaOH \rightarrow KP^{-} + Na^{+} + H_2O$$

To standardize the NaOH solution, we use the mass of KHP and divide by its molar mass to calculate the moles of KHP. Since KHP and NaOH react in a 1:1 mole ratio, moles KHP = moles NaOH at the equivalence point. To find the molarity of NaOH, moles of NaOH is divided by the volume of NaOH used to reach the equivalence point.

Pre-lab assignment:

Watch the following videos for demonstrations of the following techniques:
Weighing by difference (video located in CHE 133.30, Course Documents, Acids and Bases folder)
Using volumetric pipets
https://www.youtube.com/watch?v=HC44xjs7dho
Using burets- rinsing, filling and reading:
https://www.youtube.com/watch?v=P9mj0lNwfwY
Titrations
https://www.youtube.com/watch?v=sFpFCPTDv2w

Prepare lab notebook

Procedure: Assigned Groups of 2

Part A: Standardization of NaOH 1st student in group

Weigh out by difference ~0.25 g of dried KHP into a 250 ml Erlenmeyer flask and add 50 mL of distilled water and 2-3 drops of phenolphthalein. Rinse and then fill your buret with the nominal (approximate) 0.05 M NaOH. Record the actual concentration of NaOH from the reagent bottle. Record the initial volume of NaOH. (To how many decimal places should the volume be recorded to?) Titrate the KHP solution to a persistent **faint** pink color. Record the final volume of NaOH. Complete two additional trials and compute the average molarity of NaOH and the average deviation.

Average Deviation from the mean = $\sum_i |(x_i - \bar{x})|/n$

Data Table 1: Standardization of NaOH

	Trial 1	Trial 2	Trial 3
Weight of vial and initial KHP			
Weight of vial and remaining KHP			
Weight KHP			
Moles KHP			
Moles NaOH			
Initial buret reading			
Final buret reading			
Volume NaOH			
Molarity NaOH			

Average molarity = ____________________
Average deviation = ___________________

Based on your average deviation, is your data precise?

How would you determine whether your data is accurate?

Part B: Titration of Acid Rain

2nd student in group

Transfer 25.00 ml of the acid rain sample, with a volumetric pipet, to a 250 mL Erlenmeyer flask and 2-3 drops of phenolphthalein. Record the initial volume of NaOH. Titrate the acid rain sample to a persistent **faint** pink color. Record the final volume of NaOH. Assume that all of the acid in the sample is HNO_3. Complete two additional trials.
Compute the average molarity of nitric acid. Report the average deviation.

Data Table 2: Determination of Nitric Acid Concentration

Concentration of standardized NaOH = _______________

25.00 mL of acid rain

	Trial 1	Trial 2	Trial 3
Initial buret reading			
Final buret reading			
Volume NaOH			
Moles NaOH			
Moles nitric acid			
Molarity nitric acid			

Average nitric acid molarity ______________________

Average deviation _____________________

Based on your average deviation, is your data precise?

How does the error in the standardization of NaOH (Part A) affect the accuracy in the determination of the concentration of nitric acid in Part B?

Post-lab Quiz: In your lab section of Brightspace (CHE133.Lxx) under Quizzes, complete the Post-Lab Quiz for Experiment #2.