Experiment date:

**Experiment 22 - Standardization of a Basic Solution and Determination of Equivalent Mass of Acid**

**Goal:** To determine the equivalent mass of an unknown acid via titration.

\[
\text{equivalent mass (EM) } = \frac{g_{\text{acid}}}{\text{mol } H^+}
\]

it is analogous to molar mass (MM) = \(\frac{g_{\text{acid}}}{\text{mol acid}}\)

for a monoprotic acid, HX (1 mol \(H^+\) per mol HX), EM = MM
for a diprotic acid, \(H_2Y\), (2 mol \(H^+\) per mol \(H_2Y\)), EM = MM/2
for a triprotic acid, \(H_3Z\), (3 mol \(H^+\) per mol \(H_3Z\)), EM = MM/3

**A. Standardization**

In this experiment, standardization is used to determine the concentration of NaOH, using HCl of known concentration.

As with any titration, we want to stop the experiment at the endpoint. In this case the endpoint occurs when mol \(H^+\) = mol OH\(^-\). In this experiment, we will use phenolphthalein to tell us the endpoint has been reached. Phenolphthalein is colorless in acidic solution and pink in basic solution. We will start with the acid and indicator in a flask and the base in the buret. We will add the base to the solution in the flask until one drop turns the solution pink.

ex. 16.32 mL of NaOH solution was required to neutralize 19.55 mL of a 0.2047 M HCl solution. Calculate the molarity of the NaOH.

\[
(1) \quad M_{H^+} = M_{HCl} = 0.2047 \text{ M}
\]

\[
(2) \quad M_{OH^-} = \frac{(M_{H^+})(V_{H^+})}{V_{OH^-}} = \frac{(0.2047 \text{ M})(19.55 \text{ mL})}{16.32 \text{ mL}} = 0.2452 \text{ M}
\]

\[
(3) \quad M_{NaOH} = M_{OH^-} = 0.2452 \text{ M}
\]

***Instructors editorial: The lab book's author should receive 1000 lashes with a wet noodle for solving the above problem by this method. The only proper use of the relationship \(M_1V_1 = M_2V_2\) is for dilution problems. The only reason it works in this case is that HCl and NaOH react in a 1:1 ratio. It would not work for any other ratios, and thus should not be used as a general practice.

**B. Determination of Equivalent Mass of an Unknown Acid**

As defined earlier, EM = \(\frac{g_{\text{acid}}}{\text{mol } H^+}\) 

\(\text{Weighed out} \rightarrow \text{Determined by Titration}\)
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A typical problem with titration is that we overshoot the endpoint. In the case of this experiment, that means adding too much NaOH from the buret into the flask with the unknown acid. In this experiment, we can perform a process called **back-titration**. Standardized HCl is titrated into the mixture until the endpoint is reached.

ex. A 0.9123 g sample of unknown solid acid requires 48.22 mL of 0.1204 M NaOH for neutralization. There are 2.01 mL of 0.1002 M HCl used for back-titration. What is the equivalent mass of the unknown acid?

\[
\begin{align*}
(1) \text{ mol OH}^- &= \frac{0.04822 \text{ L} \times 0.1204 \text{ mol}}{1 \text{ L}} = 0.005806 \text{ mol OH}^- \\
(2) \text{ mol HCl} &= \frac{0.00201 \text{ L} \times 0.1002 \text{ mol}}{1 \text{ L}} = 0.0002014 \text{ mol HCl} \\
(3) \text{ mol H}^+ \text{ in unknown solid acid} &= 0.005806 \text{ mol OH}^- - 0.0002014 \text{ mol HCl} = 0.005605 \text{ mol H}^+ \\
(4) \text{ EM} &= \frac{0.9123 \text{ g}}{0.005605 \text{ mol H}^+} = 162.8 \text{ g/mol}
\end{align*}
\]

**Procedure:**

Part A.

(1) Omit first paragraph. A stock solution of dilute NaOH has already been prepared.

(2) Starting on page 135, paragraph 3. "Draw about 20 mL HCl..." This will also change the last paragraph to "...about 20, 30 and 40 mL."

Omit part C (optional section)

**End-of-Lab Protocols:** (add these to your written procedure)

(1) Collect all solutions and excess/unused chemicals in a large beaker. Test the pH.

(2) If the pH is below 5, add approximately 1 g increments of sodium bicarbonate (NaHCO₃) to the solution until the pH is above 5. Proceed to Step 4.

(3) If the pH is above 11.5, add approximately 1 g increments of sodium bicarbonate (NaHCO₃) to the solution until the pH is below 11.5. Proceed to Step 4.

(4) Pour the resulting mixture down the drain with plenty of water.

**ASA:** do all of it