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Lab Section

CHEM101L.416**Introduction**

Figure 1. Graphical analysis of a gross trial of a conductometric titration of $\text{Ca}(\text{OH})_2$ with 0.0014 M H_3PO_4 resulting in a minimum conductivity value of 44 $\mu\text{S}/\text{cm}$ and a minimum volume measurement of 28.36 mL, calculated using linear trend-lines

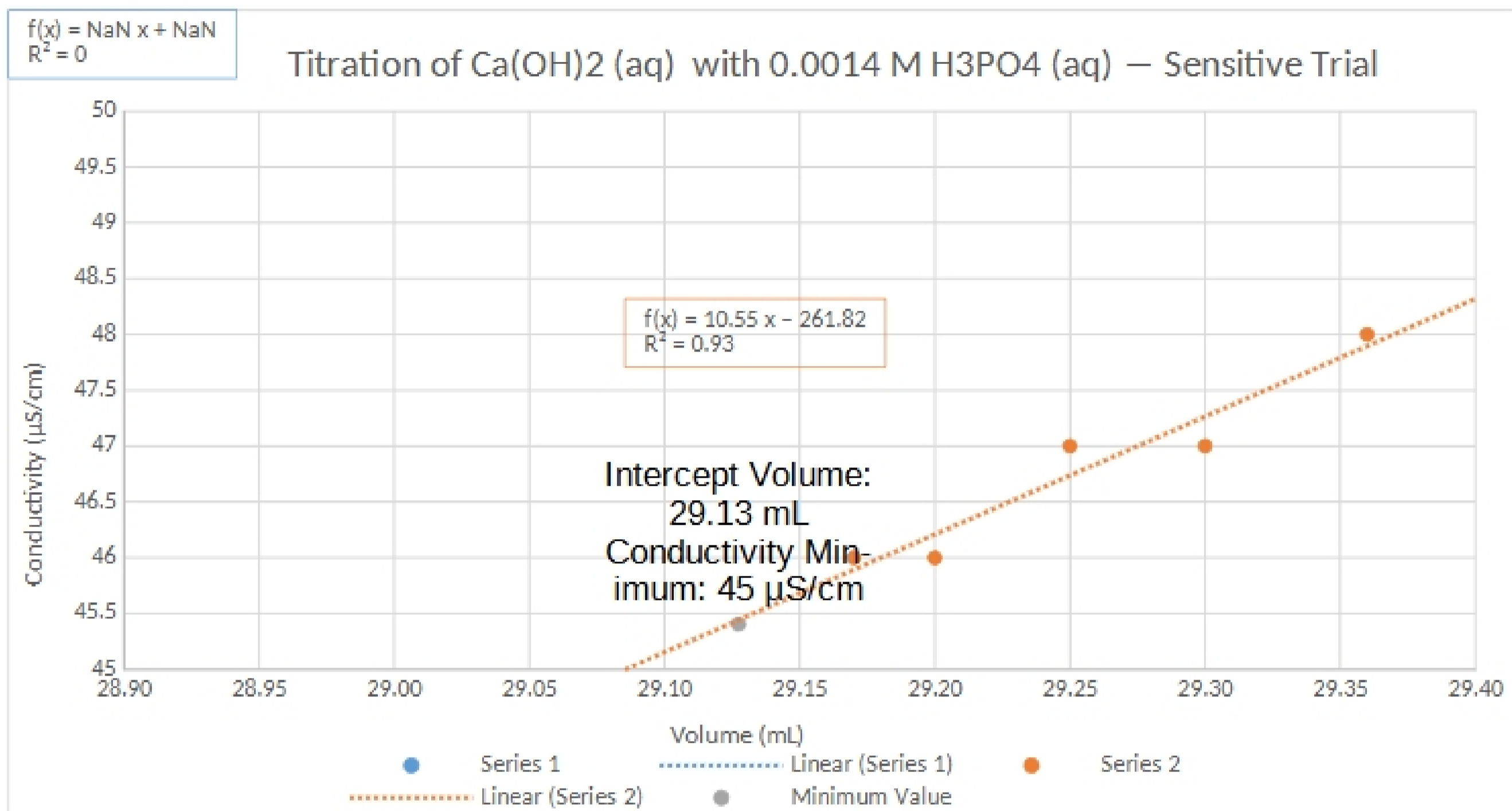


Figure 2. Graphical analysis of a sensitive trial of a conductometric titration of $\text{Ca}(\text{OH})_2$ with 0.0014 M H_3PO_4 resulting in a minimum conductivity value of 45 $\mu\text{S}/\text{cm}$ and a minimum volume measurement of 29.13 mL, calculated using linear trend-lines

Conclusion:

Within the experiment, two conductometric titrations (a gross trial—for finding a rough target volume for the sensitive trial—and a sensitive trial—for narrowing the results and determining a more accurate equilibrium point) were performed in order to determine the mole ratio of phosphoric acid (H_3PO_4 (aq)) to calcium hydroxide [$\text{Ca}(\text{OH})_2$ (aq)]. Calcium hydroxide, when dissociated in water, acts as an electrolyte; by using a conductivity meter, the conductance of the solution can be measured in $\mu\text{S}/\text{cm}$ and thus be used as an indicator for the progression of the reaction. As phosphoric acid is added to the solution, a reaction occurs:



According to the above reaction, as this reaction proceeds, one mole of a precipitate (calcium phosphate) and six moles of water form when three moles of calcium hydroxide and two moles of phosphoric acid react. Because the precipitate is, by definition, not dissolved in solution and does not contribute to the conductivity of the solution, and also because water is intrinsically a poor conductor of electricity, we can observe this reaction occurring by measuring the change in conductivity through this process (thus the name “conductometric titration”). Using the data collected from the sensitive trial (as graphically detailed in **Figure 2**), it was calculated with the empirical data that 4.1×10^{-5} mol. of PO_4^{3-} was reacting with 1.1×10^{-4} mol. of Ca^{2+} , which would result in an empirical formula of $\text{Ca}_8(\text{PO}_4)_3$ (to see how moles of each substance, estimated and empirical, were calculated, see “Calculations” section following conclusion). The calculated empirical formula, as opposed to the correct formula of $\text{Ca}_3(\text{PO}_4)_2$, results from the difference in the theoretical value of used $\text{Ca}(\text{OH})_2$ and the actual value. According to the measured data associated with the volume of H_3PO_4 and the stoichiometric coefficients within the chemical equation (three moles of $\text{Ca}(\text{OH})_2$ for every two moles of H_3PO_4), 6.2×10^{-5} mol. of $\text{Ca}(\text{OH})_2$ *should have been* used in the reaction. Yet, according to the empirical data, 1.1×10^{-4} mol. of $\text{Ca}(\text{OH})_2$ was *actually* used in the reaction (a -44% error).

Using the empirical results, a chemical reaction is formulated that is unable to be balanced:



This error between empirical and actual results may have been rooted from a multitude of occurrences. For example, when dropping the stir bar into the $\text{Ca}(\text{OH})_2$ solution for the sensitive trial, some of the solution splashed out which would have caused there to be a lower amount of calcium ions available for the reaction to occur, thus skewing the mole ratios. Furthermore, at the beginning of the lab, a “titration card” was not being used in order to view the volume measurements more accurately. However, within the middle of the sensitive trial, a card was put to use, which could have caused differences in the reported volume measurements and thus the precision of the measurement. Additionally, when performing the experiment, the water level in the beaker had to be increased in order to completely submerge the probes of the conductivity meter; in order to do this, water was added to the calcium hydroxide solution. Because the water used in lab is extracted from the tap and is not necessarily “deionized,” any ions present in the water could have also had an association with the phosphate ions present in the reaction, thus skewing the results.

Calculations:

Moles of PO_4^{3-} ions used to neutralize $\text{Ca}(\text{OH})_2$ solution:

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$0.0014 \text{ M} = \frac{x \text{ mol. of solute}}{.02913 \text{ L}}$$