

# Determination of Aspirin

## Background

- The intensity of the color of a solution is directly proportional to the concentration of the colored species in the solution.
- In a spectrophotometer, light shines through a sample solution and a detector measures the intensity of the light AFTER it passes through the sample. As the intensity the color increases, MORE light is ABSORBED by the solution, and *less light passes through* to the detector.

Commercially aspirin, acetylsalicylic acid, is available in 3 strengths, extra strength (500 mg), regular strength (325 mg), and low dosage (81 mg). In basic solution, aspirin is hydrolyzed to produce sodium salicylate and sodium acetate.



In water, the Fe<sup>+3</sup> cation forms the light yellow [Fe(H<sub>2</sub>O)<sub>6</sub>]<sup>+3</sup> ion. The [Fe(H<sub>2</sub>O)<sub>6</sub>]<sup>+3</sup>, reacts with the colorless salicylate ion to produce the purple [Fe(salicylate)(H<sub>2</sub>O)<sub>4</sub>]<sup>+</sup> ion.

- The intensity of the purple color of the solution is directly proportional to concentration of the salicylate ion present. Therefore, the more salicylate ion present in the solution (from the hydrolysis reaction of aspirin) the greater the absorption of light.
- Prepare a set of solutions with known sodium salicylate concentrations and add  $\text{FeCl}_3$  solution. Measure the absorbance of the solutions.

Plot Absorbance (y axis) vs. concentration (x axis) in a calibration plot. The slope of the best-fit straight line will be a constant, as seen in the equation (Beer's Law) below. The line MUST go through the origin (0,0).

$$\begin{aligned} \text{Absorbance} &= \text{constant} \times \text{Concentration} \\ (y) &= (\text{slope}) \times (x) \end{aligned}$$

Dissolve an unknown aspirin tablet in a basic solution to create sodium salicylate. Dilute an aliquot of this solution with  $\text{FeCl}_3$ . Measure the absorbance of this dilute aspirin solution.

Use Beer's Law and the Absorbance of the dilute aspirin solution to solve for the concentration of dilute aspirin solution.

Finally, determine the amount of aspirin in the original sample and the strength of the tablet.

## **Calculations**

**For calculating the M of a concentrated solution:**

- $M_{\text{conc}} = (\text{mass} / \text{Molar Mass}) / (\text{Volume})$

**For calculating the M of a dilute solution:**

- $M_{\text{conc}} \times \text{Vol}_{\text{conc}} = M_{\text{dilute}} \times \text{Vol}_{\text{dilute}}$

- $M_{\text{dilute}} = \left( \frac{M_{\text{conc}} \times \text{Vol}_{\text{conc}}}{\text{Vol}_{\text{dilute}}} \right)$

**For calculating the slope of a straight line:**

- $\text{slope} = \frac{\Delta y}{\Delta x} = \frac{(y_2 - y_1)}{(x_2 - x_1)}$

**For calculating concentration from Beer's Law:**

- $\text{Absorbance} = \text{slope} \times \text{Concentration}$

- $\text{Concentration} = \left( \frac{\text{Absorbance}}{\text{slope}} \right)$

**For calculating the molarity of a concentrated solution:**

- $M_{\text{conc}} = \left( \frac{M_{\text{dilute}} \times \text{Vol}_{\text{dilute}}}{\text{Vol}_{\text{conc}}} \right)$

**For calculating the mg of aspirin:**

$\text{mg} = \text{Molarity} \times \text{Volume} \times \text{Molar Mass} \times (1000 \text{ mg}/1 \text{ g})$