

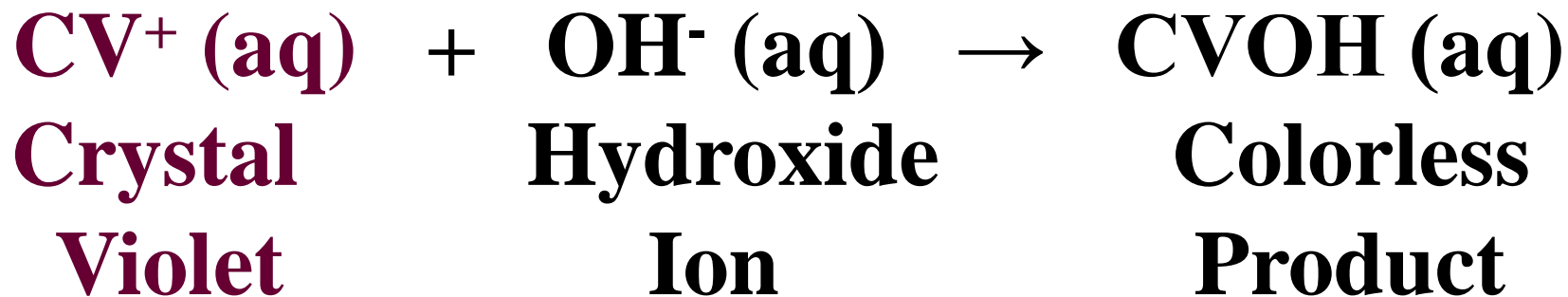
# Crystal Violet:

Goal of experiment is to study the kinetics and determine

1. the overall reaction order.

2. determine the activation energy( $E_a$ )

of the reaction between the dye crystal violet and hydroxide ion.



# Kinetics:

The study of the rate or speed at which a chemical reactions occur converting reactants into products.



$$\text{rate} = \frac{-\Delta[\text{reactant}]}{\Delta t} = \frac{\Delta[\text{product}]}{\Delta t}$$

$\Delta[\text{reactant}]$ : change in reactant concentration.

$\Delta[\text{product}]$ : change in product concentration.

$\Delta t$ : time period( $t_{\text{FINAL}} - t_{\text{INITIAL}}$ )

# Rate Laws and Reaction Order:

Reaction rates depend on reactant concentration.



Rate Law relates reactant concentrations to reaction rates.

$$\text{Rate} = k[A]^m[B]^n$$

**k**: rate constant

**m, n**: exponents

**[A], [B]**: concentration of reactants in molarity

# Reaction Order:

Reaction order refers to the values of the exponents in a rate law. In other words the value of m and n.

Example:



Rate 2nd order in A, 1st order in B, and 3rd order overall.

# First Order Reaction:



$$\mathbf{Rate} = \mathbf{k[R]}$$

## Integrated Rate Law:

$$\ln[\mathbf{R}]_t = \ln[\mathbf{R}]_o - \mathbf{kt}$$

$[\mathbf{R}]_o = [\mathbf{R}]$  at  $t = 0$  s.

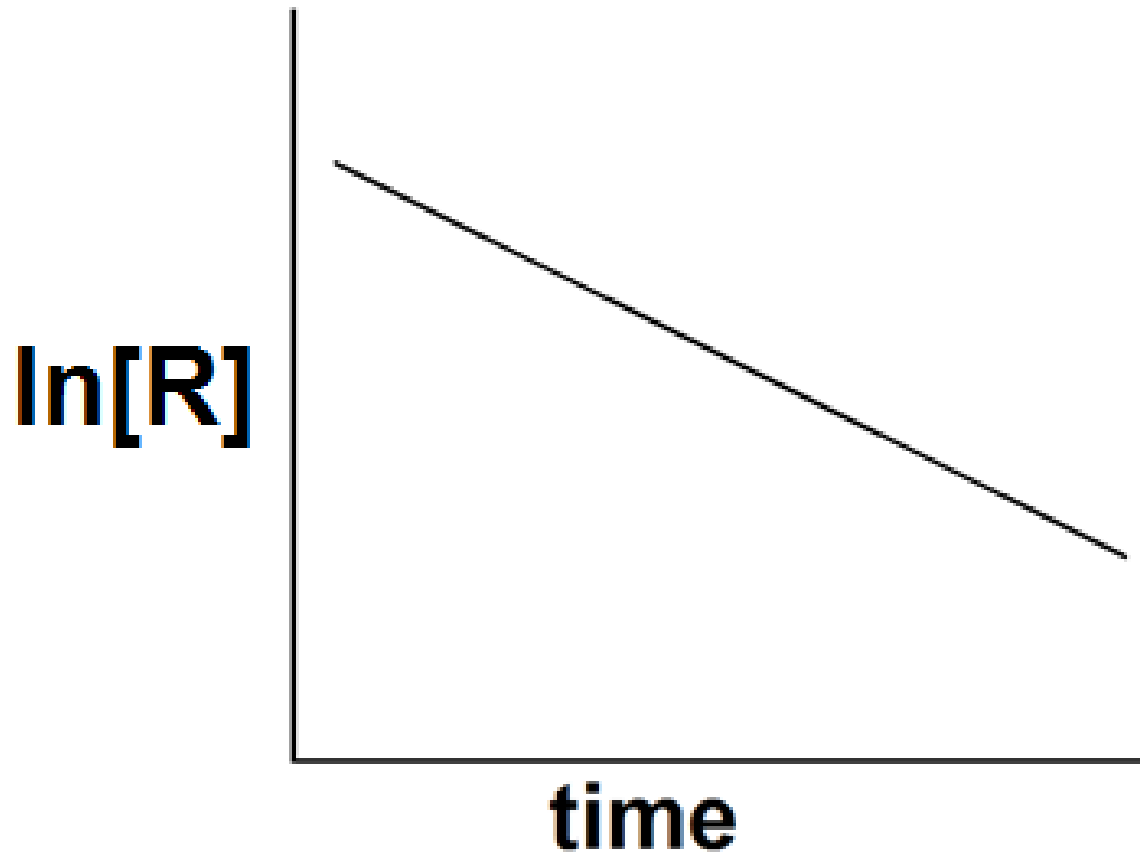
$[\mathbf{R}]_t = [\mathbf{R}]$  at time  $t$ .

$k =$  rate constant.

$t =$  time.

# First Order Reaction:

Plot of  $\ln[R]$  vs. time. If get a straight line proves first order.



# Second Order Reactions:

For a single reactant.



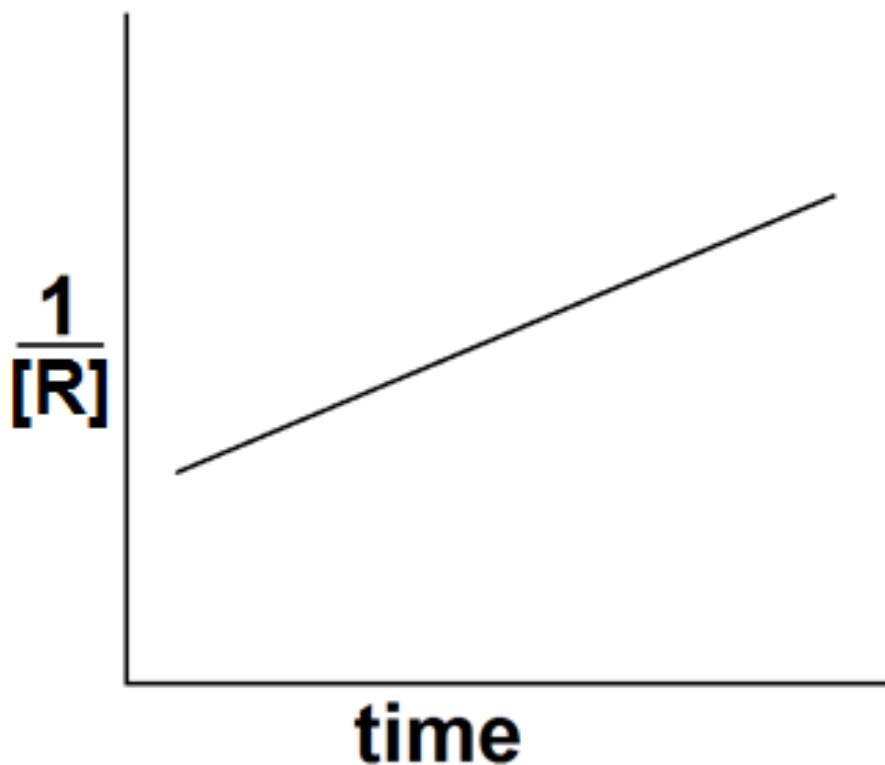
$$\mathbf{Rate} = \mathbf{k[R]^2}$$

Integrated Rate Law

$$\frac{1}{[R]_t} = kt + \frac{1}{[R]_o}$$

# Second Order Reaction:

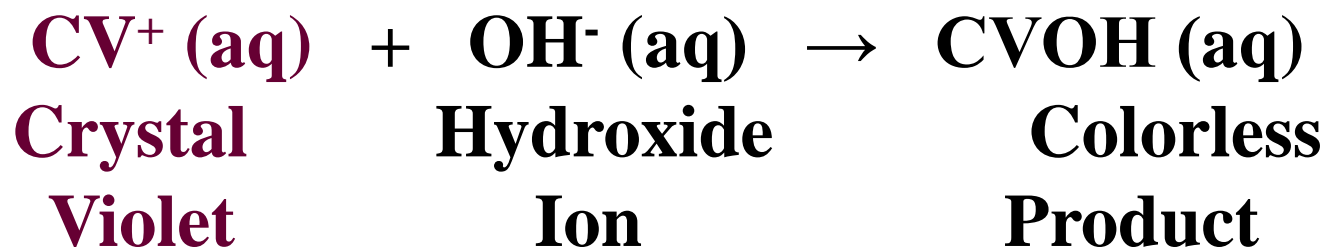
**Plot of  $1/[R]$  vs. time. If get a straight line proves second order.**





# Kinetics of Crystal Violet Reaction:

As reaction occurs the crystal violet color disappears.



The rate law for this reaction is:

$$\text{rate} = k[\text{CV}^+]^m[\text{OH}^-]^n$$

**k** is the rate constant, **m** is the order of the reaction with respect to **CV<sup>+</sup>**, and **n** is the order of the reaction with respect to **OH<sup>-</sup>**.

# **Part I: Determining Reaction Order:**

$$\text{rate} = k[\text{CV}^+]^m[\text{OH}^-]^n$$

**Determining m and n.**

**For m will monitor change in concentration as a function of time. Will measure Absorbance(A) of reaction mixture as a function of time.**

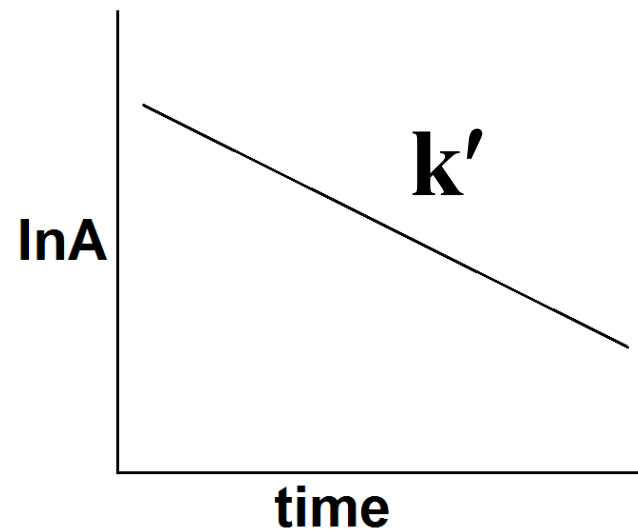
**“A” directly related to concentration of CV<sup>+</sup>.**

# Part I: Determining “m”:

**Absorbance(A) vs. time. Assume  $[\text{OH}^-]$  constant.**

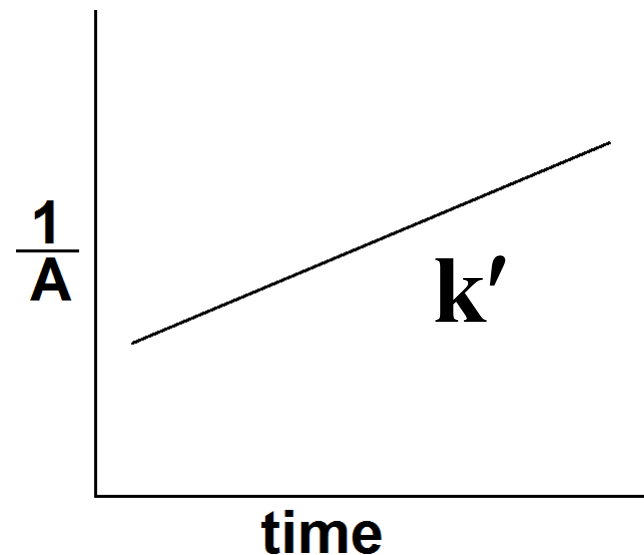
**Plot  $\ln A$  vs. time.**

**If linear,  $m = 1$ .**



**Plot  $1/A$  vs. time.**

**If linear,  $m = 2$ .**



## Part I: Determining “n”:

Will study reaction for two trials of  $[\text{OH}^-]$ .

For trial 1,  $[\text{OH}^-] = 0.0200 \text{ M}$

For trial 2,  $[\text{OH}^-] = 0.0100 \text{ M}$

Will determine rate constant for both trials,  $k'_1$  and  $k'_2$ . Determined from graph.

Determine  $n$ .

$$n = \frac{\log (k'_1/k'_2)}{\log ([\text{OH}^-]_1/ [\text{OH}^-]_2)}$$

## **Part II: Determining Activation**

### **Energy( $E_a$ ):**

**Activation Energy( $E_a$ ): The minimum energy required in a collision for a reaction to occur.**

**Will study reaction at different temperatures.**

# Arrhenius Equation:

**Relates reaction rate to temperature.**

$$\ln k = \frac{-E_A}{RT} + \ln A$$

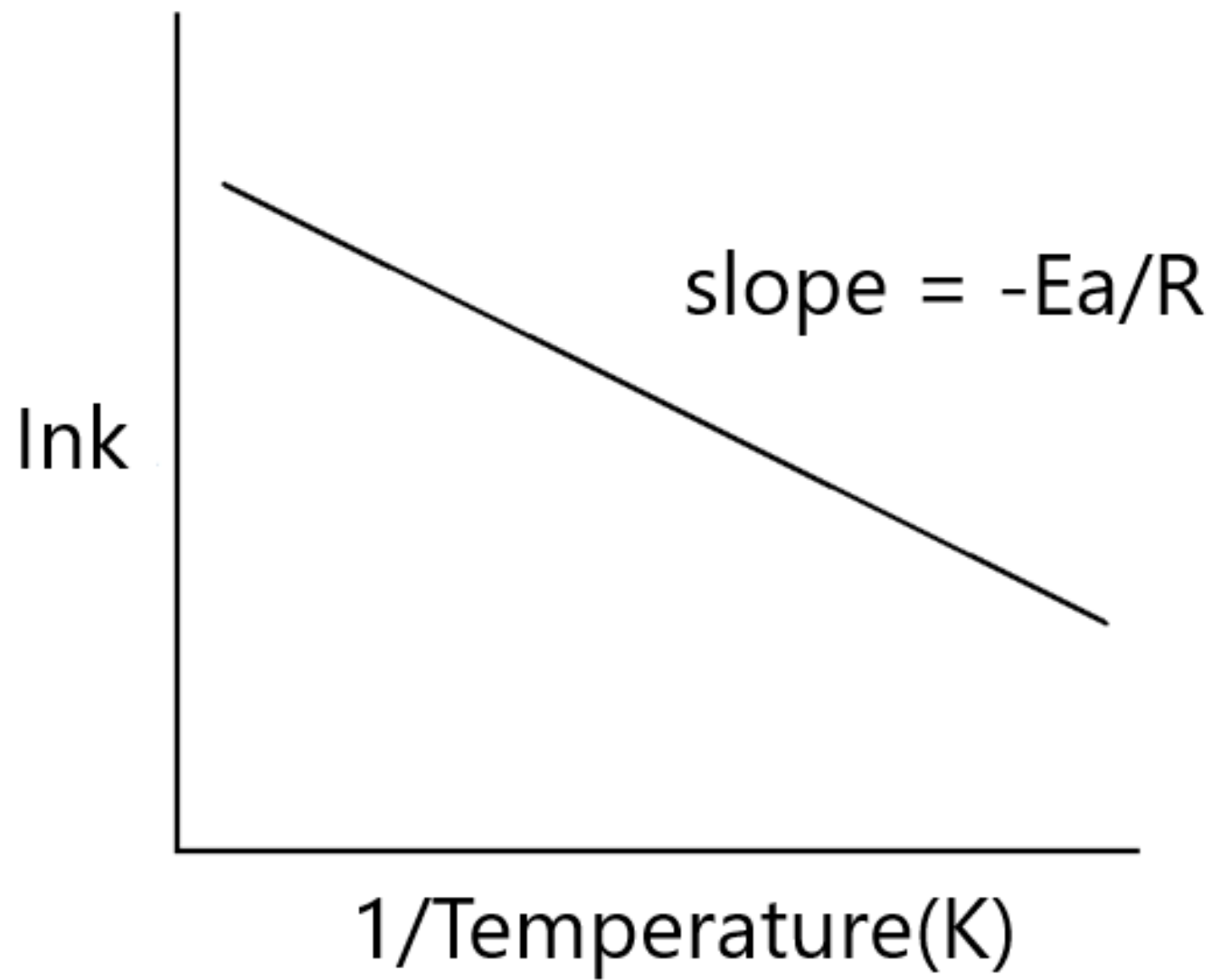
**$E_A$ : activation energy(J/mole)**

**k: rate constant**

**R: 8.314 J/K·mole**

**T: temperature in Kelvin**

**A: frequency factor**



**Plot  $\ln(A)$  versus time for trials 1-4. The pseudo rate constants,  $k'$ , can be calculated from the slopes of the plots.**

**Plot  $\ln(k')$  versus  $1/T$ . Find the slope of the best fit line equation. Remember that the slope equals  $-E_a/R$ , where  $R$  is the ideal gas constant ( $8.314 \text{ J/K}\cdot\text{mol}$ ).**