

## 2<sup>nd</sup> order and zero-order reactions

1. At elevated temperatures, nitrogen dioxide decomposes to nitric oxide and molecular oxygen following 2<sup>nd</sup> order kinetics.



At 300°C, the rate constant for the reaction is  $0.54 \text{ M}^{-1} \text{ s}^{-1}$ .

- a) In a particular reaction, the initial concentration of  $\text{NO}_2$  is  $8.00 \times 10^{-3} \text{ M}$ . What is the **concentration** of  $\text{NO}_2$  at  $t = 20.0 \text{ min}$ .

$$\begin{aligned} \frac{1}{[\text{A}]_t} &= kt + \frac{1}{[\text{A}]_0} \\ \frac{1}{[\text{NO}_2]_t} &= \left(0.54 \frac{1}{\text{M} \cdot \text{s}}\right)(1200 \text{ s}) + \frac{1}{8.00 \times 10^{-3} \text{ M}} \\ \frac{1}{[\text{NO}_2]_t} &= 773 \\ [\text{NO}_2]_t &= \mathbf{1.3 \times 10^{-3} \text{ M}} \end{aligned}$$

- b) What is the half-life of the reaction when the initial concentration of  $\text{NO}_2$  is  $6.00 \times 10^{-3} \text{ M}$ ?

$$\begin{aligned} t_{\frac{1}{2}} &= \frac{1}{k[\text{A}]_0} \\ t_{\frac{1}{2}} &= \frac{1}{\left(0.54 \frac{1}{\text{M} \cdot \text{s}}\right)(6.00 \times 10^{-3} \text{ M})} = \mathbf{3.1 \times 10^2 \text{ s}} \end{aligned}$$

- c) What is the half-life of the reaction when the initial concentration of  $\text{NO}_2$  is  $3.00 \times 10^{-3} \text{ M}$ ?

$$\begin{aligned} t_{\frac{1}{2}} &= \frac{1}{k[\text{A}]_0} \\ t_{\frac{1}{2}} &= \frac{1}{\left(0.54 \frac{1}{\text{M} \cdot \text{s}}\right)(3.00 \times 10^{-3} \text{ M})} = \mathbf{6.2 \times 10^2 \text{ s}} \end{aligned}$$

2. Make a plot of rate of reaction vs. concentration for a zero-order reaction.

