

**This is BCE#33.**

I recommend you print out this page and bring it to class. [Click here](#) to show a set of five BCE33 student responses randomly selected from all of the student responses thus far in a new window.

John , here are [your responses](#) to the BCE and the [Expert's response](#).

1. A simple reaction follows first order kinetics. For first order reactions the half-life can be calculated using the expression  $t(1/2) = 0.693/k$ . The half-life is the time required for the concentration to fall to half of the initial value. Complete the following table;

[Reactant]	0.600 M	0.400 M	0.300 M	0.250 M	Q1
Time (sec)	0	Q2	137	Q3	274

← 0.150 M  
53%  
.225 28%

1. Determine the concentration of the reactant after 274 seconds.

Since we have some data we can determine the rate constant for the reaction since we know the initial concentration and the the time required for half the initial concentration to react (the half-life). Recall from class the half-life for a first order reaction is

$$t_{1/2} = 0.693/k$$

so

$$k = 0.693/ t_{1/2}$$

$$k = 0.693/ 137 \text{ s} = 0.00506 \text{ s}^{-1}$$

Now that we know the rate constant we can use the integrated rate law for a first order reaction to calculate concentration of the reactant at 274 s.

$$\ln ([\text{reactant}]_{t=274 \text{ s}}/[\text{reactant}]_0) = - kt$$

$$\ln ([\text{reactant}]_{t=274 \text{ s}}/0.600 \text{ M}) = - 0.00506 \text{ s}^{-1} * 274 \text{ s}$$

$$\ln ([\text{reactant}]_{t=274 \text{ s}}/0.600 \text{ M}) = - 1.386$$

$$\ln ([\text{reactant}]_{t=274 \text{ s}}) - \ln(0.600 \text{ M}) = - 1.386$$

$$\ln ([\text{reactant}]_{t=274 \text{ s}}) = - 1.386 + \ln(0.600)$$

$$\ln ([\text{reactant}]_{t=274 \text{ s}}) = - 1.386 + (-0.511)$$

$$\ln ([\text{reactant}]_{t=274 \text{ s}}) = - 1.897$$

$$e(\ln ([\text{reactant}]_{t=274 \text{ s}})) = e^{- 1.897}$$

$$[\text{reactant}]_{t=274 \text{ s}} = 0.150 \text{ M}$$

Also notice that 274 s is also two half-lives ( $2 * 137 \text{ s}$ ), so we could have figured out that the  $[\text{reactant}]$  would be half the  $[\text{reactant}]$  at 137 s, or  $0.5 * 0.300 \text{ M} = 0.150 \text{ M}$ .

**2. How long does it take for the concentration of the reactant to fall to 0.400 M?**

80 seconds 47% 68 min 43%

$$\ln ([\text{reactant}]_{t=? \text{ s}}/[\text{reactant}]_0) = - kt$$

$$\ln (0.400 \text{ M} / 0.600 \text{ M}) = - 0.00506 \text{ s}^{-1} * t$$

$$\ln (0.667) = - 0.00506 \text{ s}^{-1} * t$$

$$-0.405 = - 0.00506 \text{ s}^{-1} * t$$

$$-0.405 / -0.00506 \text{ s}^{-1} = t$$

$$t = 80.1 \text{ s}$$

**3. How long does it take for the concentration of the reactant to fall to 0.250 M?**

173 47%

$$\ln ([\text{reactant}]_{t=?} / [\text{reactant}]_0) = -kt$$

$$\ln (0.250 \text{ M} / 0.600 \text{ M}) = -0.00506 \text{ s}^{-1} * t$$

$$\ln (0.417) = -0.00506 \text{ s}^{-1} * t$$

$$-0.875 = -0.00506 \text{ s}^{-1} * t$$

$$-0.875 / -0.00506 \text{ s}^{-1} = t$$

$$t = 173 \text{ s}$$

When the first order integrate rate law is written in the following form;

$$\ln[A]_t = -kt + \ln[A]_0$$

The equation has the form of the equation for a straight line.

4. If the data in the table at the top of the page were plotted according to the equation,  $\ln[A]_t = -kt + \ln[A]_0$ , what would be plotted on the y-axis and what would be plotted on the x-axis?

the  $\ln [A]$  86%

On the y-axis we plot the  $\ln[A]$  and on the x-axis we must plot time.

5. What does the slope of the line tell us and what is the y-intercept?

-slope is equal to the rate constant. The y-intercept is the initial concentration of A 67%

The slope of the line is equal to  $-k$  (the rate constant for the reaction). The y-intercept is the  $\ln[A]_0$  (the initial concentration of A).

6. For a simple reaction that follows second order kinetics when the initial concentration is 0.0300 M the time required for the concentration to fall to 0.0150 M is 28 seconds. How long will it take for the concentration to fall to 0.0125 M? to 0.00625M?

39 seconds to for the reactant concentration to reach 0.0125 M and 106 seconds to for 53% 60%

**the reactant concentration to reach 0.00625 M.**

**This problem is for a reaction that follows second order kinetics. The integrated rate law for a simple, second order reaction is given as;**

$$1/[A]_t - 1/[A]_o = kt$$

**To solve the problem, to determine the time require for the concentration to fall from 0.0300 M to 0.0125 M we need to know the rate constant for the reaction. We are given the time required for the reaction to fall from 0.0300 M to 0.0150 M so we can substitute into the integrated rate laaw and solve for the rate constant:**

$$1/0.0150 \text{ M} - 1/0.0300 \text{ M} = k * 28 \text{ s}$$

$$66.7 \text{ M}^{-1} - 33.3 \text{ M}^{-1} = k * 28 \text{ s}$$

$$33.3 \text{ M}^{-1} = k * 28 \text{ s}$$

$$33.3 \text{ M}^{-1} / 28 \text{ s} = k$$

$$1.19 \text{ M}^{-1} \text{ s}^{-1} = k$$

**Now we can determine the time required for the [A] to fall from 0.0300 M to 0.0125 M, and also to 0.00625 M**

$$1/0.0125 \text{ M} - 1/0.0300 \text{ M} = 1.19 \text{ M}^{-1} \text{ s}^{-1} * t$$

$$80 \text{ M}^{-1} - 33.3 \text{ M}^{-1} = 1.19 \text{ M}^{-1} \text{ s}^{-1} * t$$

$$46.7 \text{ M}^{-1} = 1.19 \text{ M}^{-1} \text{ s}^{-1} * t$$

$$46.7 \text{ M}^{-1} / 1.19 \text{ M}^{-1} \text{ s}^{-1} = t$$

**39.2 s = t required for the [A] to fall from 0.0300 M to 0.0125 M**

$$1/0.00625 \text{ M} - 1/0.0300 \text{ M} = 1.19 \text{ M}^{-1} \text{ s}^{-1} * t$$

$$160 \text{ M}^{-1} - 33.3 \text{ M}^{-1} = 1.19 \text{ M}^{-1} \text{ s}^{-1} * t$$

$$126.7 \text{ M}^{-1} = 1.19 \text{ M}^{-1} \text{ s}^{-1} * t$$

$$126.7 \text{ M}^{-1} / 1.19 \text{ M}^{-1} \text{ s}^{-1} = t$$

**106.0 s = t required for the [A] to fall from 0.0300 M to 0.00625 M**

**7. Is there anything about the questions that you feel you do not understand? List your concerns/questions.**

**nothing**

**8. If there is one question you would like to have answered in lecture, what would that question be?**

**nothing**