

## Kinetics Key

1. Calculate the **average** speed of reaction if at 10.0 seconds the reactant concentration is 1.25M but at 55.0 seconds the concentration is 1.02M.

$$\text{Ave. rate} = \Delta M / \Delta t = (M_f - M_i) / (t_f - t_i) = (1.02 - 1.25 \text{ M}) / (55.0 - 10.0 \text{ s}) = -5.1 \times 10^{-3} \text{ M/s}$$

2. If the rate of disappearance of H<sub>2</sub> is 3.00 M/s, calculate the rate of NH<sub>3</sub> appearance. Circle your final answer. Given: N<sub>2</sub> + 3 H<sub>2</sub> → 2 NH<sub>3</sub>

$$\Delta \text{NH}_3 / 2 \Delta t = \Delta \text{H}_2 / 3 \Delta t \quad \text{so rearrange } \Delta \text{NH}_3 / \Delta t = (2/3) \Delta \text{H}_2 / \Delta t = (2/3) 3.00 \text{ M/s} = 2.00 \text{ M/s}$$

3. If a hypothetical rate law is rate = k[A]<sup>2</sup>[B]<sup>2</sup> what is the overall reaction order? 4th

4. The rate law for the following reaction is found to be rate = k[H<sub>2</sub>][I<sub>2</sub>]. H<sub>2</sub> (g) + I<sub>2</sub> (g) → 2 HI(g)  
If the rate = 4.24 x 10<sup>-3</sup> M/s when [H<sub>2</sub>] = 0.150 M and [I<sub>2</sub>] = 0.250 M, calculate the rate when [H<sub>2</sub>] = 0.300 M and [I<sub>2</sub>] = 0.200 M. (Hint: first find k)

$$4.24 \times 10^{-3} \text{ M/s} = k (0.150 \text{ M}) (0.250 \text{ M}) \quad \text{and now solve for } k = 0.1131 \text{ M}^{-1} \text{ s}^{-1}$$

$$\text{then rate} = 0.1131 \text{ M}^{-1} \text{ s}^{-1} (0.300 \text{ M})(0.200 \text{ M}) = 0.00678 \text{ M/s}$$

5. If the rate of a reaction increases by a factor of 9 when the concentration of a reactant is tripled, then the reaction is 2nd order with respect to that reactant.

6. If 25% of a sample, which obeys first order kinetics, remains after 45 seconds, calculate the rate constant.

$$\ln (25 / 100) = -k (45\text{s}) \quad \text{solve} \quad k = 0.031 \text{ s}^{-1}$$

7. If plutonium-246 decays with a first order half-life of 200.0 years, what percentage of a sample would remain after 600.0 years?

600 years would be 3 half lives, so 100% to 50% in one half life, 50 to 25% in 2 half lives, 25 to 12.5% in 3 half lives. So the answer is **12.5%**. Or you could do the calculation the long way:

$$200 \text{ y} = 0.693 / k \quad \text{solve for } k = 0.003465 \text{ y}^{-1}$$

$$\text{then } \ln A_t = (-0.003465 \text{ y}^{-1})(600 \text{ y}) + \ln 100 \quad \text{and solve for } A_t = 12.5 \%$$

8. The half-life for a radioactive element that follows first-order decay is 534 years. If an 8.5 M solution is buried at a waste site, how long will it take for the solution to reach a concentration of 0.50 M?

$$534 \text{ y} = 0.693 / k \quad \text{thus } k = 0.001298 \text{ y}^{-1} \quad \text{then } \ln (0.50 / 8.5) = (-0.001298 \text{ y}^{-1}) t \quad \text{and } t = 2180 \text{ years}$$

9. Given the following reaction: CO (g) + Cl<sub>2</sub> (g) → COCl<sub>2</sub> (g) Determine the orders for CO and Cl<sub>2</sub> and then write the rate law for this reaction. Then calculate the rate constant, k, and include the proper units.

[CO] (M)	[Cl <sub>2</sub> ] (M)	Rate (M/s)
1.00x10 <sup>-2</sup>	2.00x10 <sup>-2</sup>	2.64 x 10 <sup>4</sup>
3.00x10 <sup>-2</sup>	2.00x10 <sup>-2</sup>	7.92 x 10 <sup>4</sup>
1.00x10 <sup>-2</sup>	4.00x10 <sup>-2</sup>	1.06 x 10 <sup>5</sup>

$$\text{CO order} = 1 \quad \text{Cl}_2 \text{ order} = 2 \quad \text{Rate} = k[\text{CO}][\text{Cl}_2]^2$$

$$k = \frac{\text{Rate}}{[\text{CO}][\text{Cl}_2]^2} = \frac{2.64 \times 10^4 \text{ M/s}}{(1.00 \times 10^{-2} \text{ M})(2.00 \times 10^{-2} \text{ M})^2} = 6.60 \times 10^9 \text{ M}^{-2} \text{ s}^{-1}$$