

**Chem 104**  
**Solution to Extra Problems for Chapter 14**

Chapter 14, Extra Problem 1. The radioactive isotope  $^{54}\text{V}$  decays by beta emission with a half-life of 55 s. (a) What fraction of a sample of  $^{54}\text{V}$  will remain after 220 s? (b) What fraction will remain after 75 s?

Solution.

(a) First determine how many half-lives have elapsed:

$$h = 220 \text{ s} / 55 \text{ s} = 4$$

From  $[A] = [A]_0(1/2)^h$ , the fraction  $[A]/[A]_0 = (1/2)^4 = 1/16 = 0.062_5 = 0.063$

(You should be able to do this kind of problem, in which a whole number of half-lives have elapsed, without a calculator.)

(b) Use the same approach for this, but use your calculator to find  $(1/2)^h$ .

$$h = 75 \text{ s} / 55 \text{ s} = 1.3_64$$

$$[A]/[A]_0 = (1/2)^{1.364} = 0.38_{86} = 0.39$$

Chapter 14, Extra Problem 2. Consider the hypothetical reaction  $\text{A}_2(\text{g}) + 2\text{B}(\text{g}) + 2\text{C}_2(\text{g}) \rightarrow 2\text{AC}(\text{g}) + 2\text{BC}(\text{g})$  for which the following kinetic data have been collected.

Exp.	$[\text{A}_2]$ , mol/L	$[\text{B}]$ , mol/L	$[\text{C}_2]$ , mol/L	Rate, mol/L·s
1	0.120	0.240	0.120	$3.62 \times 10^{-4}$
2	0.480	0.240	0.120	$7.24 \times 10^{-4}$
3	0.480	0.240	0.360	$7.24 \times 10^{-4}$
4	0.480	0.120	0.240	$3.62 \times 10^{-4}$

(a) Determine the rate law expression for the reaction. (b) Calculate the value of the rate constant,  $k$ , with the proper units.

Solution.

- (a) From experiments 1 and 2, multiplying  $[A_2]$  by 4 while keeping the other reactant concentrations the same causes the rate to increase by a factor of 2. Therefore, the order with respect to  $[A_2]$  is  $1/2$ , because  $(4)^{1/2} = 2$ . From experiments 2 and 3, multiplying  $[C]$  by three while keeping the other reactant concentrations the same causes no change in rate. Therefore, the order with respect to  $[C]$  is 0; i.e., rate does not depend on  $[C]$ . Because the rate is zero order in  $[C]$ , we can use either experiments 3 and 4 or 2 and 4 to see the effect of changing  $[B]$  on rate while  $[A_2]$  is held constant. By either comparison, diminishing  $[B]$  by half causes the rate to go to half. Therefore, the reaction is first order in  $[B]$ . The overall differential rate law for the reaction is  $Rate = k[A_2]^{1/2}[B]$ , which is  $3/2$  order overall.
- (b) Use data from any experiment and solve  $Rate = k[A_2]^{1/2}[B]$  for  $k$ .

$$k = 4.35 \times 10^{-3} (\text{mol/L})^{-1/2} \cdot \text{s}^{-1}$$

Chapter 14, Extra Problem 3: Consider the hypothetical reaction  $A_2(g) + 2B(g) + 2C_2(g) \rightarrow 2AC(g) + 2BC(g)$  for which the experimentally determined rate law has been found to be  $Rate = k[A_2]^{1/2}[B]$ . The following two mechanisms have been proposed for this reaction.

Mechanism I:



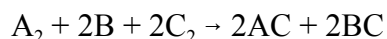
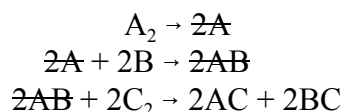
Mechanism II:



- (a) Show that both proposed mechanisms are consistent with the overall stoichiometry of the reaction,  $A_2(g) + 2B(g) + 2C_2(g) \rightarrow 2AC(g) + 2BC(g)$ .
- (b) What species are reaction intermediates in each mechanism?
- (c) Derive the rate law expression for each mechanism in terms of observable reactant species ( $A_2$ ,  $B$ , and  $C_2$ ). On the basis of your rate law expressions, which mechanism is more plausible?

Solution:

- (a) The equations are actually the same in both cases. In either, the second step equation and the third step equation need to be multiplied by 2 in order for all steps to add to the overall stoichiometry.



- (b) A and AB are reaction intermediates. Neither is present initially as a reactant or finally as a product. Both are produced and consumed in the course of the mechanism.

(c) Mechanism I:

From the slow rate-determining step (step 3), the overall rate is  $Rate = rate_3 = k_3[AB][C_2]$ . But AB is a reaction intermediate, so we need to derive an expression in terms of observable reactants for [AB]. From the step 2 equilibrium, we can write  $rate_2 = rate_{-2}$ ; i.e., the forward and reverse rates are equal. From the molecularity of the processes, we can then write

$$k_2[A][B] = k_{-2}[AB] \Rightarrow [AB] = (k_2/k_{-2})[A][B] = K_2[A][B]$$

But this expression for [AB] still involves an unobservable reaction intermediate, A. From the step 1 equilibrium, we can write  $rate_1 = rate_{-1}$ , and from the molecularity of the forward and reverse processes we can write

$$k_1[A_2] = k_{-1}[A]^2 \Rightarrow [A]^2 = k_1/k_{-1}[A_2] = K_1[A_2] \Rightarrow [A] = K_1^{1/2} [A_2]^{1/2}$$

Substituting this expression for [A] into the previous expression for [AB] gives

$$[AB] = K_2 \{K_1^{1/2} [A_2]^{1/2}\} [B] = K_2 K_1^{1/2} [A_2]^{1/2} [B]$$

Substituting this expression for [AB] into  $Rate = rate_3 = k_3[AB][C_2]$  gives

$$Rate = k_3 \{K_2 K_1^{1/2} [A_2]^{1/2} [B]\} [C_2] = k [A_2]^{1/2} [B] [C_2]$$

This does not match the experimentally observed  $Rate = k[A_2]^{1/2} [B]$ , so it is not plausible.

Mechanism II:

From the slow rate-determining step (step 2), the overall rate is  $Rate = k_2[A][B]$ . But A is a reaction intermediate, so we need to derive an expression in terms of observable reactants for

[A]. From the step 1 equilibrium, we can write  $rate_1 = rate_{-1}$ , and from the molecularity of the forward and reverse processes we can write

$$k_1[A_2] = k_{-1}[A]^2 \Rightarrow [A]^2 = k_1/k_{-1}[A_2] = K_1[A_2] \Rightarrow [A] = K_1^{1/2} [A_2]^{1/2}$$

Substituting this into  $Rate = k_2[A][B]$  gives

$$Rate = k_2 \{K_1^{1/2} [A_2]^{1/2}\} [B] = k[A_2]^{1/2} [B]$$

This matches the observed rate law, so Mechanism II is more plausible.