

CH302 TA Practice Quiz 6 In principle, this can be done without a calculator if you use the hints provided and a piece of scratch paper.

Note:  $\ln 2 = .693$

1. Consider the reaction below. Which of the following is the correct rate expression?



- a.  $-\Delta [A]/(2\Delta t) = -\Delta [B]/(\Delta t) = \Delta [C]/(\Delta t)$
- b.  $-\Delta [A]/(\Delta t) = -\Delta [B]/(2\Delta t) = -\Delta [C]/(\Delta t)$
- c.  $\Delta [A]/(2\Delta t) = -\Delta [B]/(\Delta t) = \Delta [C]/(2\Delta t)$
- d.  $-\Delta [A]/(2\Delta t) = -\Delta [B]/(\Delta t) = \Delta [C]/(2\Delta t)$

The phrase "rate expression" refers to the ratio of the change in amount per change in time (in other words, the slope.) Because you can write a rate expression for every compound in the reaction, you need to be able to equate them. This is done by using the stoichiometric coefficients in the denominator and the sign (negative for the stuff on the left and positive for the stuff on the right) to correct for whether the compound is forming or reacting.

2. Consider the data below:

Experiment	Initial [Xe] (M)	Initial [F <sub>2</sub> ] (M)	Initial Rate (M/sec)
1	0.5	0.25	0.00156
2	1.5	1.0	0.05625
3	0.75	0.25	0.0032
4	1.50	0.25	0.01406
5	0.5	1.0	0.00625

Which of the following is a correct rate law for the reaction between Xe and F<sub>2</sub>?

- a.  $r = k[Xe][F_2]$
- b.  $r = k$
- c.  $r = k[Xe][F_2]^2$

$$r = k[Xe]^2[F_2]$$

The rate of the reaction =  $k[Xe]^x[F_2]^y$

Solve for x (when [F<sub>2</sub>] doesn't change):

$$r_4/r_3 = (0.01406 \text{ M/sec}) / (0.0032 \text{ M/sec}) = (k \cdot 1.5^x \cdot 0.25^y) / (k \cdot 0.75^x \cdot 0.25^y)$$

[F<sub>2</sub>] remains the same, so they cancel out

$$\text{So, } 4 = 2^x \text{ and } x = 2$$

Solve for y: (when [Xe] doesn't change):

$$r_5/r_1 = (0.00625 \text{ M/sec}) / (0.00156 \text{ M/sec}) = (k \cdot 0.5^x \cdot 1.0^y) / (k \cdot 0.5^x \cdot 0.25^y)$$

$$4 = 4^y \text{ and } y = 1$$

$$r = k[Xe]^2[F_2]$$

3. What is the order of the reaction if  $k = 5 \times 10^7 \text{ M}^{-2}\text{s}^{-1}$ ?

- a. 1st order
- b. 2nd order
- c. 3rd order
- d. 0th order

Quick and easy way to determine order:

Order of reaction = 1 - (-2) = 3

To determine the order of the reaction just subtract the exponent of the molarity from 1.

4. Consider a reaction with the activation energy of 1.6 kJ/mol. At room temperature  $k = 2.3 \times 10^3 \text{ s}^{-1}$ . Calculate the temperature when the rate constant is  $2.2 \times 10^3 \text{ s}^{-1}$ . (Hint:

$\ln(0.9565) = -0.045$ )

- a. 279K
- b. 300K
- c. 298 K
- d. 295 K

$\ln(k_2/k_1) = (E_a/R)(1/T_1 - 1/T_2)$

$T_1 = 1/(\ln(k_2/k_1)R/E_a + 1/T_2) = 1/(\ln(2.2 \times 10^3 / 2.3 \times 10^3) * 8.314 / 1600 \text{ J} + 1/298)$

=278.81 K

5. You sneak 500 Cheetos into a movie theatre where the movie playing is 90 minutes long. If the rate constant for your Cheetos-eating is 5 Cheetos per minute, how many Cheetos do you have in your possession when a theatre-employee throws you out of the movie for texting after 40 minutes?

- a. None, he forcefully removed them from you along with your cell phone.
- b. 200
- c. 300
- d. 50

First of all, it is illegal to take away your cell phone. I don't know about the Cheetos, though.

This seemingly real life example is actually subject to the teachings of your kinetics lectures, even if you decided to do it using common sense rather than chemistry concepts (this might lead you to believe that perhaps common sense is also used in teaching chemistry concepts, but you don't have to.)

First step is to see that the units of the rate constant are [ ]/t which is zeroth order. So now it is just plug and chug.

$[\text{cheetos}] = [\text{initial cheetos}] - akt$

(we will assume you eat the cheetos one at a time so  $a = 1$ )

Therefore  $[\text{cheetos}] = 500 - (1)(5)(40) = 300$

The 90 minute length of the movie is in there just to be annoying.

6. If the half life of a compound is 13.86 seconds and it decays in a first order reaction, what is the reaction constant,  $k$ ?  $[C_0] = 2 \text{ M}$

- a. .05 M/s
- b. .072 M/s

c.  $.05 \text{ s}^{-1}$

d.  $.072 \text{ s}^{-1}$

$$t_{1/2} = .693/k$$

$$k = .693/(13.86 \text{ s}) = .05 \text{ s}^{-1}$$

If you didn't know that equation you could still use the first order integrated rate law:

$$\ln [1/2] = \ln [1] - kt, \text{ this occurs at } t = t_{1/2}$$

$$\ln 1 - \ln 2 = \ln[1] - kt_{1/2}, \ln 1 = 0$$

$$.693 = kt_{1/2}$$

$$k = .693/(13.86 \text{ s}) = .05 \text{ s}^{-1}$$

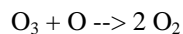
The concentration was included to confuse you.

7. Which of the following statements does not explain why a thermodynamically favorable reaction, a reaction where  $\Delta G \ll 1$ , may not occur?

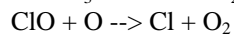
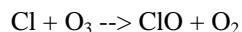
- a. There was not enough free energy to overcome the activation energy of the reaction.
- b. There is only enough energy to form the transition state complex.
- c. The reactants did not collide in the proper orientation.
- d. The temperature is low, and the reaction is endothermic.

According to the transition state theory, a thermodynamically favorable reaction will not occur if there is insufficient energy to create a transition state complex.

8. Consider the following reaction:



If the mechanism is:



Which compound is the catalyst and which compound is an intermediate?

- a. ClO, ClO
- b. ClO, Cl
- c. O, ClO
- d. Cl, ClO

Cl enters and leaves the reaction unchanged, so it is a catalyst. ClO is formed during the reaction but is not present in the products so it is an intermediate.