

For answers, send email to: [admin@tutor-homework.com](mailto:admin@tutor-homework.com).

**Include file name:** Chemistry\_Worksheet\_0009

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1. Chlorine dioxide,  $\text{ClO}_2$ , dissolves in basic solution to give  $\text{ClO}_2^-$  and  $\text{ClO}_3^-$  ions, according to the following equation (no catalyst is needed):  $2\text{ClO}_2(\text{aq}) + 2\text{OH}^-(\text{aq})$



Use the following data and determine the rate law for this reaction using the Method of Initial Rates. Concentrations are in moles per liter, and rate is in mole per liter per second (M/s).

Trial	$[\text{ClO}_2]$	$[\text{OH}^-]$	Rate
1	0.060	0.030	0.0248
2	0.020	0.030	0.00276
3	0.020	0.090	0.00828

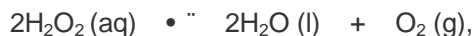
Show the equations you used to obtain the exponents.

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2. What is the rate constant,  $k$ , for the reaction in Question #8? Use the data and the rate law you obtained in answering that question. Be sure to include units.

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3. Hydrogen peroxide ( $\text{H}_2\text{O}_2$ ) decomposes to water and oxygen in a first-order process:



governed by the integrated rate law:  $\ln[\text{H}_2\text{O}_2]_t = -kt + \ln[\text{H}_2\text{O}_2]_0$ .

Commercial stock solutions of hydrogen peroxide are 30.0 % by mass  $\text{H}_2\text{O}_2$  in water. If it takes 295 seconds for the initial concentration to decrease to 3.00 M, what was the molarity of the original 30.0 % solution? The rate constant,  $k$ , for this reaction is  $0.00366 \text{ s}^{-1}$  at 21 degrees C.

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4. Given that the density of the 30.0 % solution in Question #10 is 1.00 g/mL, convert 30.0 % by mass  $\text{H}_2\text{O}_2$  to molarity. Assuming this value to be the accepted value, calculate the % error of the experimental value in Question #10 from this one.

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5. What is the half-life for the decomposition in question #10?

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6. Often called the "fake blood" reaction, iron (III) ion reacts with thiocyanate ion ( $\text{SCN}^-$ ) to form a complex ion that looks like blood:



In excess thiocyanate, the process is first-order in  $\text{Fe}^{3+}$  and has a rate constant,  $k$ , equal to  $1.27 \text{ s}^{-1}$ . Starting with an initial  $\text{Fe}^{3+}$  concentration of  $0.880 \text{ M}$ , what will the concentration of  $\text{Fe}^{3+}$  be after  $2.00$  seconds have passed?

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7. Butadiene reacts to form its dimer (two molecules joined together) in a second-order process:



If an initial concentration of  $0.0100 \text{ M}$   $\text{C}_4\text{H}_8$  decreases to  $0.00241 \text{ M}$  in  $5204$  seconds, what is the half-life for this reaction?

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8. The reaction:  $2\text{NOCl}(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{Cl}_2(\text{g})$

has rate constants of  $9.32 \times 10^{-6} \text{ s}^{-1}$  at  $77.0 \text{ deg C}$  and  $6.94 \times 10^{-4} \text{ s}^{-1}$  at  $128 \text{ deg C}$ .

Use the two-point form of the Arrhenius equation (Eqn. 12.11 on p. 587) to determine the activation energy for this reaction.

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9. What is the rate constant for the reaction in question #15 at  $460 \text{ K}$ ?

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10. The radioactive isotope  $^{32}\text{P}$  decays by first-order kinetics and has a half-life of  $14.3$  days. How long does it take for  $98.5\%$  of a sample of  $^{32}\text{P}$  to decay?

\_\_\_\_\_ days

[View a Summary of Kinetics Equations](#)

### 11. Temperature Dependence of Rate Constants and the Collision Model. (Question 12.56)

A first-order reaction has rate constants of  $4.6 \times 10^{-2} \text{ s}^{-1}$  and  $8.1 \times 10^{-2} \text{ s}^{-1}$  at  $0 \text{ }^\circ\text{C}$  and  $20. \text{ }^\circ\text{C}$ , respectively. What is the value of the activation energy?

$E_a =$  \_\_\_\_\_  $\text{kJ/mol}$

## 12. Temperature Dependence of Rate Constants and the Collision Model. (Question 12.58)

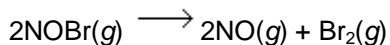
Chemists commonly use a rule of thumb that an increase of 10 K in temperature doubles the rate of a reaction. What must the activation energy be for this statement to be true for a temperature increase from 32 to 42 °C?

$$E_a = \text{_____} \text{ kJ/mol}$$

[View the Equation for Calculating  \$E\_a\$  from Experiments at two Different Temperatures](#)

## 13. Integrated Rate Laws. (Question 12.42)

The rate law for the reaction



at some temperature is  $\text{Rate} = -\text{Change in } [\text{NOBr}]/\text{change in } t = k[\text{NOBr}]^2$

a. If the half-life for this reaction is 2.50 s when  $[\text{NOBr}]_0 = 0.900 \text{ M}$ , calculate the value of  $k$  for this reaction.

$$k = \text{_____} \text{ L/(mol}\cdot\text{s)}$$

b. How much time is required for the concentration of NOBr to decrease to 0.100 M?

$$\text{_____} \text{ s}$$