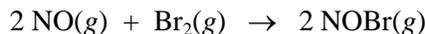


Kinetic Problem

From Raymond F. X. Williams

1. According to the equation,



In a certain reaction mixture the rate of formation of NOBr(g) was found to be $4.50 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$. What is the rate of consumption of Br₂(g), also in $\text{mol L}^{-1} \text{ s}^{-1}$?

a. $4.50 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$ b. $2.25 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$ c. $9.00 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$ d. $2.12 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$ e. $2.03 \times 10^{-3} \text{ mol L}^{-1} \text{ s}^{-1}$	$\frac{1}{2}(-r_{\text{NO}}) = (-r_{\text{Br}_2}) = \frac{1}{2} (r_{\text{NOBr}})$ $= (-r_{\text{Br}_2}) = \frac{1}{2} (r_{\text{NOBr}}) = \frac{1}{2} (4.50 \times 10^{-4}) = 2.25 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$
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2. The units in which the rate of a chemical reaction in solution is measured are (could be);

a. $\text{L}^2 \text{ mol}^{-1} \text{ s}^{-1}$ b. $\text{mol L}^{-1} \text{ s}^{-1}$ c. s^{-2} d. mol s L^{-1} e. $\text{sec L}^{-1} \text{ mol}^{-1}$	$\text{rate} = \frac{\Delta[A]}{\Delta t} = \frac{M}{\text{sec}} = \frac{\text{mol}}{\text{L} \cdot \text{sec}} = \text{mol} \cdot \text{L}^{-1} \cdot \text{sec}^{-1}$
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3. If a reaction involving a single reactant is first order with a rate constant of $4.50 \times 10^{-2} \text{ s}^{-1}$, how much time is required for 75.0% of the initial quantity of reactant to be used up?

a. 16.7 seconds b. 30.8 seconds c. 23.1 seconds d. 25.3 seconds e. 11.6 seconds	$\ln \frac{[A]_o}{[A]_t} = kt$ $\ln \frac{100}{100 - 75} = 4.50 \times 10^{-2} \cdot t$ $100 \div 25 = \ln, \text{ans}, = \div (4.50 \times 10^{-2}) = t = 30.8 \text{ sec}$
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4. A reaction has the rate law, $\text{rate} = k[\text{A}][\text{B}]^2$. Which one of the following will cause the greatest increase in the reaction rate?

a. decreasing the temperature without changing the concentrations b. doubling the concentration of B c. quadrupling the concentration of A d. tripling the concentration of B e. doubling the concentration of A	a. $T \downarrow, \text{rate} \downarrow$ b. $\text{Rate} \approx [2\text{B}]^2 = 4\text{rate}$ c. $\text{Rate} \approx [4\text{A}] = 4\text{rate}$ d. $\text{Rate} \approx [3\text{B}]^2 = 9\text{rate}$ e. $\text{Rate} \approx [2\text{A}] = 2\text{rate}$
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5. The reaction, $2 \text{NO}(g) + \text{O}_2(g) \rightarrow 2 \text{NO}_2(g)$, was found to be first order in each of the two reactants and second order overall. The rate law should therefore be written as;

a. $\text{rate} = k[\text{NO}]^2$ b. $\text{rate} = k([\text{NO}][\text{O}_2])$ c. $\text{rate} = k[\text{NO}_2]^2[\text{NO}]^2[\text{O}_2]^{-1/2}$ d. $\text{rate} = k[\text{NO}]^2[\text{O}_2]^2$ e. $\text{rate} = k([\text{NO}][\text{O}_2])^2$	$\text{Rate} = k [\text{NO}] [\text{O}_2]$
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6. A reaction has the rate law, $\text{rate} = k[A][B]^2$. What is the overall order of the reaction?

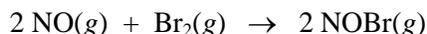
<p>a. 2 b. 4 c. 1 d. 3 e. 0</p>	<p>$\text{Rate} = k[A][B]^2$ $n = 1 + 2 = 3$</p>
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7. For the reaction, $2 \text{XO} + \text{O}_2 \rightarrow 2 \text{XO}_2$, The rate law is therefore;

run #	[XO]	[O ₂]	rate, mmol L ⁻¹ s ⁻¹
1	0.010	0.010	2.5
2	0.010	0.020	5.0
3	0.030	0.020	45.0

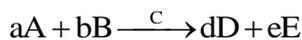
<p>a. $\text{rate} = k[\text{XO}]^2 [\text{O}_2]$ b. $\text{rate} = k[\text{XO}][\text{O}_2]^2$ c. $\text{rate} = k[\text{XO}][\text{O}_2]$ d. $\text{rate} = k[\text{XO}]^2 [\text{O}_2]^2$ e. $\text{rate} = k[\text{XO}]^2 / [\text{O}_2]^2$</p>	<p>$\frac{r_2}{r_1} = \left(\frac{[\text{O}_2]_2}{[\text{O}_2]_1}\right)^x, \quad \frac{5}{2.5} = \left(\frac{0.02}{0.01}\right)^x \quad 2 = (2)^x \quad x = 1$ $\frac{r_3}{r_2} = \left(\frac{[\text{XO}]_3}{[\text{XO}]_2}\right)^y, \quad \frac{45}{5} = \left(\frac{0.03}{0.01}\right)^y \quad 9 = (3)^y \quad y = 2$ $\text{rate} = k[\text{XO}]^2 [\text{O}_2]$</p>
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8. The experimental rate law is $\text{rate} = k[\text{NO}][\text{Br}_2]$. In a certain reaction mixture the rate of formation of NOBr(g) was found to be $4.50 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$. Which unit below is the correct unit for the rate constant in this case?



<p>a. mol L⁻¹ s⁻¹ b. s⁻¹ c. mol² L⁻² s⁻¹ d. mol⁻¹ L s⁻¹ e. mol⁻² L² s⁻¹</p>	<p>$n=1+1=2$ $k = \frac{\text{rate}}{[\text{A}]^2} = \frac{M}{\text{sec} \cdot M^2} = \frac{1}{M \cdot \text{sec}} = \frac{L}{\text{mol} \cdot \text{sec}} = \text{mol}^{-1} \cdot \text{L} \cdot \text{sec}^{-1}$</p>
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9. Given the reaction,



where C is a catalyst and $\text{rate} = k[\text{A}]^q[\text{B}]^r[\text{C}]^s$, which one of the statements below is **false**?

- The exponents q, r, and s are often integers.
- The exponent s must be determined experimentally.
- The exponents q and r are equal to the coefficients a and b, respectively.**
- The overall order of the reaction is q + r + s.
- The symbol k represents the rate constant.

10. The half-life of a chemical reaction was found to be **independent of the quantity** of material which the researcher employed. The reaction is therefore;

- possibly first order
- definitely first order**
- zero order
- possibly second order
- definitely second order

11. In a first order reaction with only one reagent, the reaction was started with a concentration of reactant equal to 0.0800 molar. After exactly two hours, the concentration had fallen to 0.0400 molar. What is the molarity after exactly three hours?

<p>a. 0.0300 M b. 0.0267 M c. 0.0340 M d. 0.0283 M e. 0.0200 M</p>	$\ln \frac{[A]_o}{[A]_t} = kt = \frac{0.693}{t_{1/2}} t$ $\ln \frac{0.08}{[A]_t} = \frac{0.693 \times 3}{2}$ $0.693 \times 3 \div 2 = \text{shift, ln, ans, } =, 0.08 \div \text{ans} = [A]_t = 0.0283M$
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12. A first order reaction $A \rightarrow B$ with a rate constant of $1.20 \times 10^{-3} \text{ min}^{-1}$. If $[A] = 0.0500$ molar, what will the concentration be 150 minutes later?

<p>a. 0.00900 M b. 0.0418 M c. 0.00926 M d. 0.00499 M e. 0.000333 M</p>	$\ln \frac{[A]_o}{[A]_t} = kt$ $\ln \frac{0.05}{[A]_t} = 1.20 \times 10^{-3} \times 150$ $1.20 \times 10^{-3} \times 150 = \text{shift, ln, ans} = 0.05 \div \text{ans} = [A]_t = 0.0418M$
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13. The rate constant for a first order decomposition reaction is 0.0111 min^{-1} . What is $t_{1/2}$?

<p>a. 111 min b. 62.4 min c. 5000 sec d. 31.25 min e. 27.1 min</p>	$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{0.0111} = 62.4 \text{ min}$
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14. Given a reaction, $2A + B \rightarrow P$, for which the observed rate law is $\text{rate} = k[A]$. Which one of the following is true?

- a. $[A] = 1/kt$
- b. $\ln[A] = k/t$
- c. $1/[A] = kt$
- d. the half-life is $0.693/k$**
- e. $e^{[A]} = -kt$

15. In a first order reaction, what fraction of the material will remain after 4 half-lives?

<p>a. 1/16 b. 1/8 c. 1/9 d. 1/4 e. 1/3</p>	$t_{1/2}(4) = \frac{[A]_o}{2^n} = \frac{1}{2^4} = \frac{1}{16}$
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16. The initial concentration of a reactant in a first order reaction is 0.620 molar. What will be its concentration after 3 half-lives?

<p>a. 0.0865 M b. 0.310 M c. 0.0775 M d. 0.103 M e. 0.207 M</p>	$t_{1/2}(3) = \frac{[A]_o}{2^n} = \frac{0.620}{2^3} = \frac{0.620}{8} = 0.0775M$
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17. For the reaction, $A \rightarrow B + C$, the rate law is $k[A]$. If it takes 80.0 seconds for 70.0% of a 10.0 gram sample of A to be transformed into products, what is the value of the rate constant?

<p>a. 0.00450 s^{-1} b. 0.0290 s^{-1} c. 0.00530 s^{-1} d. 0.0150 s^{-1} e. 5.40 s^{-1}</p>	$\ln \frac{[A]_o}{[A]_t} = kt$ $\ln \frac{100}{100 - 70} = k \times 80$ $100 \div 30 = \ln, \text{ans}, =, \div 80 = k = 0.0150 \text{ s}^{-1}$
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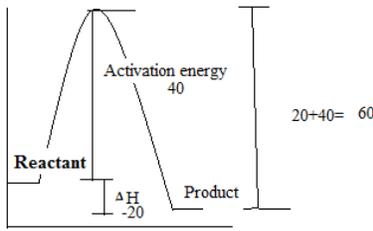
18. A reaction is first order overall. For a given sample, if its initial rate is $0.0200 \text{ mol L}^{-1} \text{ s}^{-1}$ and 25.0 days later its rate dropped to $6.25 \times 10^{-4} \text{ mol L}^{-1} \text{ s}^{-1}$, what is its half-life?

<p>a. 25.0 days b. 50.0 days c. 12.5 days d. 5.0 days e. 37.5 days</p>	$\ln \frac{[A]_o}{[A]_t} = kt = \frac{0.693}{t_{1/2}} t$ $\ln \frac{0.02}{6.26 \times 10^{-4}} = \frac{0.693}{t_{1/2}} \times 25$ $0.02 \div 6.26 \times 10^{-4} = \ln, \text{ans}, 0.693 \times 25 \div \text{ans} = t_{1/2} = 5 \text{ days}$
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19. For a one step reaction, the activation energy for the forward reaction is 40.0 kJ mol^{-1} , and the enthalpy of reaction is $-20.0 \text{ kJ mol}^{-1}$. Which statement below is true?

- The activation energy of the forward reaction would be affected to a greater extent than the activation energy of the reverse reaction by addition of a catalyst.
- The value for the enthalpy of reaction would be decreased by addition of a catalyst.
- The reaction is endothermic.
- The reverse reaction is slower than the forward reaction (smaller rate constant).**
- The reaction rate would be decreased by an increase in temperature.

20. For a one step reaction, the activation energy for the forward reaction is 40.0 kJ mol^{-1} , and the enthalpy of reaction is $-20.0 \text{ kJ mol}^{-1}$. Calculate the activation energy for the reverse reaction.

<p>a. $+60.0 \text{ kJ mol}^{-1}$ b. $-20.0 \text{ kJ mol}^{-1}$ c. $-1200 \text{ kJ mol}^{-1}$ d. $+20.0 \text{ kJ mol}^{-1}$ e. $+1200 \text{ kJ mol}^{-1}$</p>	 <p style="text-align: right;">$20 + 40 = 60$</p>
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21. The activation energy for a reaction can be found by finding the slope of a plot of $\ln(k)$ vs T^{-1} and

<p>a. adding this slope to $-R$</p> <p>b. multiplying this slope by 2.303</p> <p>c. dividing this slope by $-R$</p> <p>d. multiplying this slope by 2.303R</p> <p>e. multiplying this slope by $-R$</p>	$\ln k = \ln A - \frac{E_a}{R} \left(\frac{1}{T} \right)$ $y = b + mx$ $y = \ln k, \quad x = \frac{1}{T}, \quad m = \frac{-E_a}{R}$ $mx(-R) = E_a$
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22. For a chemical reaction, the rate constant at 250.0 °C is 0.00383 s⁻¹, and the activation energy is 22.40 kilojoules. Calculate the value of the rate constant at 335.0 °C.

<p>a. 0.00513 s⁻¹</p> <p>b. 0.00946 s⁻¹</p> <p>c. 0.00787 s⁻¹</p> <p>d. 0.0224 s⁻¹</p> <p>e. 0.000640 s⁻¹</p>	$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{T_1 - T_2}{T_1 \times T_2} \right)$ $\ln \frac{0.00383}{k_2} = \frac{22.40}{8.314 \times 10^{-3}} \left(\frac{523 - 608}{523 \times 608} \right)$ $22.4 \times (-85) \div (8.314 \times 10^{-3}) \div 523 \div 608 =, \text{shift, ln, ans} = 0.00383 \div \text{ans} = k_2 = 0.00787 \text{ sec}^{-1}$
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23. The rate constant for a certain chemical reaction is 0.00250 L mol⁻¹ s⁻¹ at 25.0 °C and 0.0125 L mol⁻¹ s⁻¹ at 50.0 °C. What is the activation energy for the reaction, expressed in kJ?

<p>a. 25.1 kJ</p> <p>b. 51.6 kJ</p> <p>c. 37.6 kJ</p> <p>d. 45.3 kJ</p> <p>e. 60.3 kJ</p>	$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{T_1 - T_2}{T_1 \times T_2} \right)$ $\ln \frac{0.0025}{0.0125} = \frac{E_a}{8.314 \times 10^{-3}} \left(\frac{298 - 323}{298 \times 323} \right)$ $0.0025 \div 0.0125 =, \text{ln, ans} = x(8.314 \times 10^{-3}) \times 298 \times 323 \div (-25) = E_a = 51.6 \text{ kJ}$
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25. Which one of the following statements concerning the rate of a chemical reaction is *false*?

- a. It will be very rapid if the activation energy is large.**
- b. It will be slow if one or more of the steps is slow.
- c. It may be inhibited sometimes by certain catalytic agents.
- d. It is dependent on temperature.
- e. It often increases when the concentrations of one of the reactants is increased.

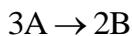
26. A variable which has **no** effect on the rate of a chemical reaction under any circumstances is;

- a. energy of activation
- b. catalyst
- c. concentration of the reactants
- d. temperature
- e. **standard enthalpy of reaction for the system**

27. A catalyst alters the rate of a chemical reaction by

- a. providing an alternate pathway which has a different activation energy
- b. changing the products formed in the reaction
- c. changing the frequency of collisions between molecules
- d. always providing a surface on which molecules react
- e. changing the enthalpy of reaction for the reaction

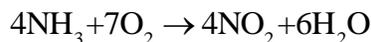
28. Consider the following reaction:



The average rate of appearance of B is given by $\Delta[B]/\Delta t$. Comparing the rate of appearance of B and the rate of disappearance of A, we get $\Delta[B]/\Delta t = \underline{\hspace{2cm}} \times (-\Delta[A]/\Delta t)$

<p>A) -2/3</p> <p>B) +2/3</p> <p>C) -3/2</p> <p>D) +1</p>	$\frac{1}{3}(-r_{NH_3}) = \frac{1}{2}(r_B)$ $(r_B) = \frac{2}{3}(-r_A)$ $\frac{2}{3}$
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29. Which substance in the reaction below either appears or disappears the fastest?



<p>A) NH_3</p> <p>B) O_2</p> <p>C) NO_2</p> <p>D) H_2O</p>	$\frac{1}{4}(-r_{NH_3}) = \frac{1}{7}(-r_{O_2}) = \frac{1}{4}(r_{NO_2}) = \frac{1}{6}(r_{H_2O})$ $(r_{O_2}) = \frac{7}{4}(r_{NH_3}) = \frac{7}{4}(r_{NO_2}) = \frac{7}{6}(r_{H_2O})$ $(r_{O_2}) = \textit{fastest}$
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30. The overall order of a reaction is 2. The units of the rate constant for the reaction are _____.

<p>A) M/s</p> <p>B) $M^{-1}s^{-1}$</p> <p>C) 1/s</p> <p>D) 1/M</p>	$k = \frac{\textit{rate}}{[A]^2} = \frac{M}{\textit{sec} \cdot M^2} = \frac{1}{M \cdot \textit{sec}} = M^{-1} \cdot \textit{sec}^{-1}$
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A flask is charged with 0.124 mol of A and allowed to react to form B according to the reaction $A(g) \rightarrow B(g)$. The following data are obtained for [A] as the reaction proceeds in 1L flask: (31-33)

Time (s)	0.00	10.0	20.0	30.0	40.0
Moles of A	0.124	0.110	0.088	0.073	0.054

31. The average rate of disappearance of A between 10 s and 20 s is _____ mol/s.

<p>A) 2.2×10^{-3} B) 1.1×10^{-3} C) 4.4×10^{-3} D) 454</p>	$\text{rate} = -\frac{\Delta[A]}{\Delta t} = -\frac{0.088 - 0.110}{20 - 10} = 2.2 \times 10^{-3}$
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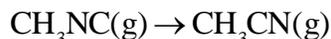
32. The average rate of appearance of B between 20 s and 30 s is _____ mol/s.

<p>A) $+1.5 \times 10^{-3}$ B) $+5.0 \times 10^{-4}$ C) -1.5×10^{-3} D) $+7.3 \times 10^{-3}$</p>	$\text{rate} = \frac{\Delta[B]}{\Delta t} = \frac{0.088 - 0.0073}{30 - 20} = 1.5 \times 10^{-3}$
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33. How many moles of B are present at 10 s?

<p>A) 0.011 B) 0.220 C) 0.110 D) 0.014</p>	$\text{rate} = \frac{\Delta[B]}{\Delta t} = \frac{[B]_{10} - [B]_0}{t - 0} = \frac{0.124 - 0.110}{10 - 0} = 0.014$
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34. At the start of an experiment, there are 0.200 mol of reactant and 0 mol of product in the reaction vessel (1L). After 25 min, 0.108 mol of reactant (CH_3NC) remain. There are _____ mol of product (CH_3CN) in the reaction vessel.



<p>A) 0.022 B) 0.540 C) 0.200 D) 0.308 E) 0.092</p>	$\Delta[\text{CH}_3\text{CN}] = 0.200 - 0.108 = 0.092M$
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35. If the rate law for the reaction;



is **first order in A** and **second order in B**, then the rate law is rate = _____.

<p>A) $k[A][B]$ B) $k[A]^2[B]^3$ C) $k[A][B]^2$ D) $k[A]^2[B]$</p>	<div style="background-color: yellow; display: inline-block; padding: 5px;">$rate = k[A][B]^2$</div>
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36. The kinetics of the reaction below were studied and it was determined that the reaction **rate** increased by a factor of **9** when the **concentration of B** was **tripled**. The reaction is _____ order in B.



$$\begin{aligned} \text{Rate} &= k[B]^x \\ 9\text{rate} &= [3B]^x \\ X &= 2 \end{aligned}$$

- A) zero B) first C) **second** D) third

37. A reaction was found to **be third order** in A. Increasing the concentration of A by a factor of **3** will cause the reaction rate to _____.

$$\begin{aligned} \text{Rate} &= k[B]^3 \\ \text{rate} &= [3B]^3 \\ \text{rate} &= 3^3 \text{ times} = 27 \end{aligned}$$

- A) remain constant
B) increase by a factor of 27
 C) increase by a factor of 9
 D) triple

The data in the table below were obtained for the reaction: (38-40)



Experiment Number	[A] (M)	[B] (M)	Initial Rate (M/s)
1	0.273	0.763	2.83
2	0.273	1.526	2.83
3	0.819	0.763	25.47

38. The order of the reaction in A is _____.

<p>A) 1 B) 2 C) 3 D) 4</p>	<div style="background-color: yellow; display: inline-block; padding: 5px;"> $\frac{r_3}{r_1} = \left(\frac{[A]_3}{[A]_1}\right)^x, \quad \frac{25.47}{2.83} = \left(\frac{0.819}{0.273}\right)^x \quad 9 = (3)^x \quad x = 2$ </div>
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39. The order of the reaction in B is _____.

<p>A) 1 B) 2 C) 3 D) 4 E) 0</p>	$\frac{r_2}{r_1} = \left(\frac{[B]_2}{[B]_1}\right)^y, \quad \frac{2.83}{2.83} = \left(\frac{1.526}{0.763}\right)^y \quad 1 = (2)^x \quad y = 0$
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40. The overall order of the reaction is _____.

$$n = 2 + 0 = 2$$

- A) 1
B) 2
C) 3
D) 4

41. For a first-order reaction, a plot of _____ versus _____ is linear.

<p>A) $\ln [A]_t, \frac{1}{t}$ B) $\ln [A]_t, t$ C) $\frac{1}{[A]_t}, t$ D) $[A]_t, t$</p>	$\ln \frac{[A]_o}{[A]_t} = kt$ $\ln [A]_o - \ln [A]_t = kt$ $\ln [A]_t = \ln [A]_o - kt$ $y = b + mx$ $y = \ln [A]_t \quad x = t \quad m = -k$
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42. The reaction below is first order in $[H_2O_2]$:



A solution originally at 0.600 M H_2O_2 is found to be 0.075 M after 54 min. The half-life for this reaction is _____ min.

<p>A) 6.8 B) 18 C) 14 D) 28</p>	$\ln \frac{[A]_o}{[A]_t} = kt = \frac{0.693}{t_{1/2}} t$ $\ln \frac{0.6}{0.075} = \frac{0.693}{t_{1/2}} \times 54$ $0.6 \div 0.075 = \ln, \text{ans}, 0.693 \times 54 \div \text{ans} = t_{1/2} = 18 \text{ days}$
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43. Of the following, unit for a reaction rate is _____.

<p>A) mol/L B) M/s C) L/mol s D) s/mol.L</p>	$\text{rate} = \frac{\Delta[A]}{\Delta t} = \frac{M}{\text{sec}} = \frac{\text{mol}}{L \cdot \text{sec}} = \text{mol} \cdot L^{-1} \cdot \text{sec}^{-1}$
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44. The data in the table below were obtained for the reaction
 $A + B \rightarrow P$

Experiment Number	[A] (M)	[B] (M)	Initial Rate (M/s)
1	0.273	0.763	2.83
2	0.273	1.526	2.83
3	0.819	0.763	25.47

The rate law for this reaction is rate = _____.

A) $k[A][B]$	$\frac{r_2}{r_1} = \left(\frac{[B]_2}{[B]_1}\right)^y, \quad \frac{2.83}{2.83} = \left(\frac{1.526}{0.763}\right)^y \quad 1 = (2)^y \quad y = 0$
B) $k[P]$	
C) $k[A]^2[B]$	$\frac{r_3}{r_1} = \left(\frac{[A]_3}{[A]_1}\right)^x, \quad \frac{25.47}{2.83} = \left(\frac{0.819}{0.273}\right)^x \quad 9 = (3)^x \quad x = 2$ $\text{rate} = k[A]^2$
D) $k[A]^2[B]^2$	
E) $k[A]^2$	

45. A compound decomposes by a first-order process. If 25.0 % of the compound decomposes in 60.0 minutes, the half-life of the compound is _____.

A) 65 minutes	$\ln \frac{[A]_o}{[A]_t} = kt = \frac{0.693}{t_{1/2}} t$ $\ln \frac{100}{100 - 25} = \frac{0.693}{t_{1/2}} \times 60$ $100 \div 75 =, \ln, \text{ans}, 0.693 \times 60 \div \text{ans} = t_{1/2} = 145 \text{ minutes}$
B) 120 minutes	
C) 145 minutes	
D) 180 minutes	

46. Of the following, _____ will lower the activation energy for a reaction.
- A) increasing the concentrations of reactants
 B) raising the temperature of the reaction
 C) adding a catalyst for the reaction
 D) removing products as the reaction proceeds

47. A particular first-order reaction has a rate constant of $1.35 \times 10^2 \text{ s}^{-1}$ at 25.0 °C. What is the magnitude of k at 95.0 °C if $E_a = 55.5 \text{ kJ/mol}$?

A) 9.56×10^3	$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{T_1 - T_2}{T_1 T_2} \right)$ $\ln \frac{1.35 \times 10^2}{k_2} = \frac{55.5}{8.314 \times 10^{-3}} \left(\frac{298 - 368}{298 \times 368} \right)$ $55.5 \times (-70) \div (8.314 \times 10^{-3} \div 298 \div 368) =, \text{shift}, \ln, \text{ans} = 1.35 \times 10^2 \div \text{ans} = k_2 = 9.56 \times 10^3 \text{ sec}^{-1}$
B) 2.85×10^4	
C) 576	
D) 4.33×10^{87}	