c 2.7183
V . 8.314

43/50

Quiz 1 (Chapter 12)

Kelsey

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	Zero-Order	First-Order	Second-Order
rate law	rate = k	rate = $k[A]$	rate = $k[A]^2$
units of rate constant	$M\,s^{-1}$	s^{-1}	$M^{-1}\,s^{-1}$
integrated rate law	$[A] = -kt + [A]_0$	$\ln[A] = -kt + \ln[A]_0$	$\frac{1}{[A]} = kt + \left(\frac{1}{[A]_0}\right)$

Arrhenius Equation: $k = Ae^{-E_a/RT}$

1. Consider the following reaction in aqueous solution:

$$A + 2B \rightarrow C + D$$

a) Write the equation that relates the rate expressions for this reaction in terms of the disappearance of A and the disappearance of B.

$$-\frac{1}{1}\frac{\Delta[A]}{\Delta t} = -\frac{1}{2}\frac{\Delta[B]}{\Delta t} = \frac{1}{1}\frac{\Delta[C]}{\Delta t} = \frac{1}{1}\frac{\Delta[D]}{\Delta t}$$

Not part of the Q.

b) If the rate of disappearance of A at a particular moment during the reaction is 1.4×10^{-4} mol L^{-1} s^{-1}, what is the rate of disappearance of B at that moment?

$$\left(1.4\times10^{-4}\frac{mol}{L\cdot s}\right)\cdot -2 = 0.0002$$

unit.

$$-\frac{1}{1}\frac{\Delta[A]}{\Delta t} = -\frac{1}{2}\frac{\Delta[B]}{\Delta t}$$

x2 x2

$$-2\frac{\Delta[A]}{\Delta t} = 1\frac{\Delta[B]}{\Delta t}$$

2. The following data have been determined for the reaction:

$$[2NO]^m + [Br_2]^n \rightarrow 2NOBr_2$$ rate = k [NO] + [Br_2]

	[NO] initial (M)	[Br₂] initial (M)	Rate (mol L⁻¹ s⁻¹)
1	0.02 x2	0.02	9.6×10^{-2}
2	0.04	0.02 x2	3.8×10^{-1} x2
3	0.02	0.04	1.9×10^{-1}

Determine 1) the rate law and 2) the rate constant for this reaction.

[NO]

$$\frac{\text{rate 1}}{\text{rate 2}} = \frac{k[NO]^m + [Br_2]^n}{k[NO]^n + [Br_2]^n}$$

$$\frac{9.6\times10^{-2}}{3.8\times10^{-1}} = \left(\frac{0.02}{0.04}\right)^m \rightarrow \ln(0.2526) = m \ln(0.5) \rightarrow m = \ln\left(\frac{0.2526}{0.5}\right)$$

$m = -0.683$

(-1)

2.

[Br₂]

Between rate 2 & 3, the inital (M) and the rate are both doubled meaning it is the first order. $[Br_2]^1$

✓

1).

rate = k[2NO]¹ + [Br₂]¹

$3.8\times10^{-1} \frac{mol}{Ls} = k[0.04M]^1 + [0.02M]^1$

$3.8\times10^{-1} \frac{mol}{Ls} = k(0.06M)$

$k = 6.33 \frac{1}{Ls}$

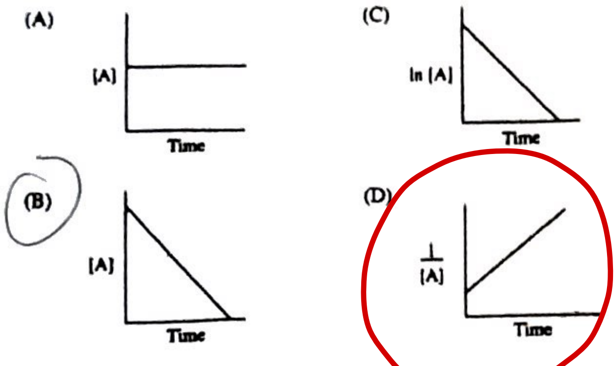
* Box your answer.

ok.

3. Which of the following graphs may have been created using the data gathered from the following reaction? Assume this is a single step reaction:

$$2A \rightarrow B + C$$

$A \rightarrow 1B + 1C$



(-2)

4. Dinitrogen pentoxide gas decomposes according to the equation: $2\ N_2O_{5(g)} \rightarrow 4\ NO_{2(g)} + O_{2(g)}$. The first-order reaction was allowed to proceed at 40.0 °C. The initial concentration of N_2O_5 was 0.400 M and after 20.0 minutes, the concentration changed to 0.289 M.

$0.400 - 0.289 = 0.111$

$20.0\ \text{min} \times \frac{60\ \text{sec}}{1\ \text{min}} = 1200$

(a) Calculate the rate constant for the reaction. $K = ?$

$Kt = \frac{\ln[A]_0}{\ln[A]} \rightarrow Kt = \ln\left(\frac{0.400\,M}{0.111\,M}\right) \rightarrow Kt = 1.2819$

$K = \frac{1.2819}{20.0\ \text{min}}$

$K = 0.0641\ \frac{1}{\text{min}} \times \frac{1\ \text{min}}{60\ \text{sec}} = 1.07 \times 10^{-3}\ \frac{1}{\text{sec}}$

(-1)

(b) After how many minutes will $[N_2O_5]$ be equal to 0.350 M?

$K = 0.0641\ \frac{1}{\text{min}}$

$t = ?$

$[A] = 0.350\,M$

$[A]_0 = 0.400\,M$

$Kt = \ln\left(\frac{0.400\,M}{0.050\,M}\right)$

$(0.0641\ \frac{1}{\text{min}})(t) = 2.079$

$t = \frac{2.079}{0.0641\ \frac{1}{\text{min}}}$

$t = 32.4\ \frac{1}{\text{min}}$

$0.400\,M - 0.350\,M = 0.05$

ok.

5. The rate constant at 550 °C for the decomposition reaction $2H_2O2 \rightarrow 2H_2O + O_2$ is $6.0 \times 10^{-7}\ s^{-1}$, and the frequency factor (A) is $1.2 \times 10^{12}\ s^{-1}$. Determine the activation energy for the reaction.

$R = 8.314 \frac{J}{mol \cdot K}$

$e = 2.7183$

$T = 550°C + 273.15 = 823.15K$

$K = 6.0 \times 10^{-7} \frac{1}{s}$

$A = 1.2 \times 10^{12} \frac{1}{s}$

$E_a = ?$

$K = Ae^{\frac{-E_a}{RT}}$

Correct

$6.0 \times 10^{-7} \frac{1}{s} = (1.2 \times 10^{12} \frac{1}{s})(2.7183)^{\wedge \frac{-E_a}{(8.314 \frac{J}{mol \cdot K})(823.15K)}}$

$5.0 \times 10^{-19} = 2.7183^{\frac{-E_a}{6843.67 \frac{J}{mol}}}$

$1.839 \times 10^{-19} = 1^{\frac{-E_a}{6843.67 \frac{J}{mol}}}$

x

(-1)

6. At 600 K, compound A decomposes to form compounds B and C via a first-order reaction. Discuss the effect of each of the following conditions on the half-life of A:

$A \rightarrow B + C$ $t_{\frac{1}{2}} = \frac{\ln 2}{K}$ $Kt = \frac{100}{50}$

(a) Increasing the initial concentration of A

increasing the concentration of A would overall increase more reacting molecules and would then reduce the half life of A

X (-1)

(b) Increasing the temperature at which the reaction occurs

increasing the temeperature would cause the collisions to happen more rapidly making the half life of A even less.

v

7. Consider the following:

1) $2NO_{(g)} \leftrightarrow N_2O_{2(g)}$ (fast, k_1 represents the forward rate constant, k_{-1} the reverse rate constant)

2) $N_2O_{2(g)} + H_{2(g)} \rightarrow N_2O_{(g)} + H_2O_{(g)}$ (slow, k_2 the rate constant)

3) $N_2O_{(g)} + H_{2(g)} \rightarrow N_{2(g)} + H_2O_{(g)}$ (fast, k_3 the rate constant)

(a) Write the overall reaction.

$2NO + \cancel{N_2O_2} + H_2 + \cancel{N_2O} + H_2 \rightarrow \cancel{N_2O_2} + \cancel{N_2O} + H_2O + N_2 + H_2O$

$\boxed{2NO + 2H_2 \rightarrow 2H_2O + N_2}$ ✓

(b) Identify all intermediates.

N_2O_2
N_2O ✓

(c) Write the overall rate law.

1) $rate_1 = K_1[2NO]^2 = K_{-1}[N_2O_2]$

$N_2O_2 = \frac{K_1 2NO}{K_{-1}}$

2) $rate_2 = K_2[N_2O_2] + [H_2]$

$K_2\left(\frac{K_1[2NO]}{K_{-1}}\right) + [H_2]$

$K_2\left(\frac{K_1}{K_{-1}}[2NO]\right) + [H_2]$

(−1)

rate = $\boxed{K[\cancel{2}NO]^2 + [H_2]}$