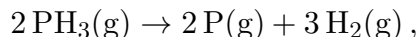


This print-out should have 25 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001 10.0 points

If the average rate of formation of $\text{H}_2(\text{g})$ is $3.90 \text{ (mol H}_2\text{)/L/s}$ for the reaction



what is the unique average reaction rate?

- 1.30 mol/L/s **correct**
- 2.60 mol/L/s
- 3.90 mol/L/s
- 7.80 mol/L/s
- 11.7 mol/L/s

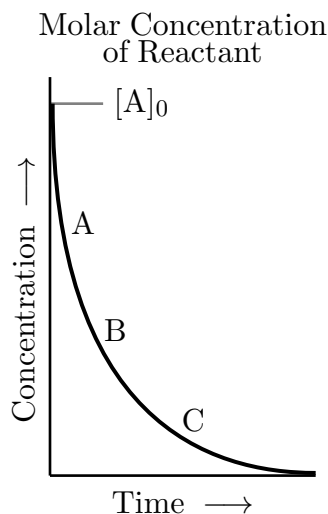
Explanation:

The unique average reaction rate equals the average rate of formation of H_2 divided by the coefficient of H_2 in the reaction equation:

$$\frac{3.90 \text{ (mol H}_2\text{)/L/min}}{3} = 1.30 \text{ mol/L/min}$$

002 10.0 points

Consider the concentration-time dependence graph for a first-order reaction.



At which point on the curve is the reaction fastest?

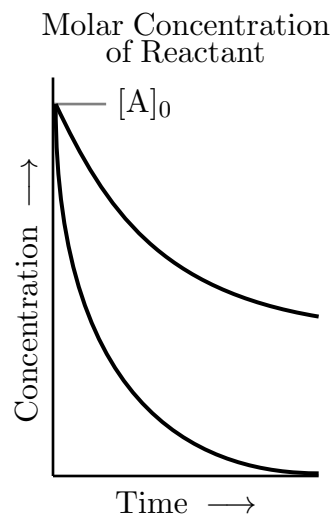
- A **correct**
- The rates are the same at all points.
- $A + t_{1/2}$
- C
- B

Explanation:

Steeper portions of the graph define faster reactions.

003 10.0 points

Consider the concentration-time dependence as shown below for two first order reactions.



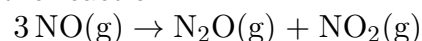
Which reaction has the greatest $t_{1/2}$?

- the reaction represented by the lower curve
- Unable to determine
- the reaction represented by the upper curve **correct**

Explanation:

004 10.0 points

When the reaction



is proceeding under conditions such that 0.015 mol/L of N_2O is being formed each second,

the rate of the overall reaction is ? and the rate of change for NO is ?.

1. $0.015 \text{ M}\cdot\text{s}^{-1}$; $-0.005 \text{ M}\cdot\text{s}^{-1}$
2. $0.015 \text{ M}\cdot\text{s}^{-1}$; $-0.045 \text{ M}\cdot\text{s}^{-1}$ **correct**
3. none of the other answers is correct
4. $0.030 \text{ M}\cdot\text{s}^{-1}$; $-0.005 \text{ M}\cdot\text{s}^{-1}$
5. $0.015 \text{ M}\cdot\text{s}^{-1}$; $+0.045 \text{ M}\cdot\text{s}^{-1}$

Explanation:

Notice that three moles of NO are required to make one mole of N_2O in this reaction, so the rate of disappearance of NO (which has a negative sign) is *three* times the rate of appearance of N_2O .

005 10.0 points

What is the rate law for the reaction



if the following data were collected?

Exp	[A] ₀	[B] ₀	[C] ₀	Initial Rate
1	0.4	1.2	0.7	2.32×10^{-3}
2	1.3	1.2	0.9	7.54×10^{-3}
3	0.4	4.1	0.8	9.25×10^{-2}
4	1.3	1.2	0.2	7.54×10^{-3}

1. $\text{rate} = 3.36 \times 10^{-3} [\text{A}]^1 [\text{B}]^3 [\text{C}]^0$ **correct**
2. $\text{rate} = 5.37 \times 10^{-3} [\text{A}]^1 [\text{B}]^3 [\text{C}]^0$
3. $\text{rate} = 1.49 \times 10^{-3} [\text{A}]^0 [\text{B}]^3 [\text{C}]^1$
4. $\text{rate} = 4.48 \times 10^{-3} [\text{A}]^1 [\text{B}]^2 [\text{C}]^1$
5. $\text{rate} = 1.79 \times 10^{-3} [\text{A}]^0 [\text{B}]^2 [\text{C}]^1$

Explanation:

006 10.0 points

Consider the reaction



What is the overall order of the reaction and the order with respect to $[\text{O}_3]$?

1. 0 and 1
2. 3 and 2
3. -1 and 3
4. 2 and 2
5. 1 and 2 **correct**

Explanation:

The order with respect to O_3 is the power of $[\text{O}_3]$ in the rate law, which is 2. The overall order of reaction is the sum of the powers in the rate law, which is $2 + (-1) = 1$.

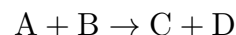
007 10.0 points

For a given reaction, the rate-law is

1. an equation in which reaction rate is equal to a mathematical expression involving concentrations of reactants. **correct**
2. a constant of proportionality between reaction rate and the concentrations of reactants.
3. the sum of the powers to which reactant concentrations appear.

Explanation:

The rate law for a reaction



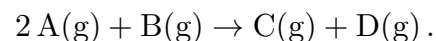
is set up as

$$\text{Rate} = k [\text{A}]^x [\text{B}]^y$$

This is an expression involving the concentrations. The constant of proportionality option is the rate constant k , while the sum of powers option is the order of reaction.

008 10.0 points

Consider the reaction



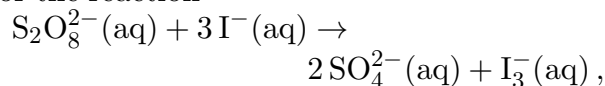
When $[A] = [B] = 0.10 \text{ M}$, the rate is 2.0 M/s ;
 for $[A] = [B] = 0.20 \text{ M}$, the rate is 8.0 M/s ;
 and for $[A] = 0.10 \text{ M}$, $[B] = 0.20 \text{ M}$, the rate
 is 2.0 M/s . What is the rate law?

1. rate = $k [A] [B]$
2. rate = $k [A]^2$ **correct**
3. rate = $k [A]$
4. rate = $k [A] [B]^0$
5. rate = $k [B]^2$

Explanation:

009 10.0 points

For the reaction



rate = $k [\text{S}_2\text{O}_8^{2-}][\text{I}^-]$. When the reaction is followed under pseudo-first-order conditions with $[\text{S}_2\text{O}_8^{2-}] = 200 \text{ mM}$ and $[\text{I}^-] = 1.5 \text{ mM}$, the rate constant was 1.82 s^{-1} . What is the second order rate constant k for the reaction?

1. $6.07 \times 10^3 \text{ M}^{-1}\text{s}^{-1}$
2. $1.37 \times 10^{-2} \text{ M}^{-1}\text{s}^{-1}$
3. $9.10 \text{ M}^{-1}\text{s}^{-1}$ **correct**
4. $1.21 \times 10^3 \text{ M}^{-1}\text{s}^{-1}$
5. $1.82 \text{ M}^{-1}\text{s}^{-1}$

Explanation:

pseudo-first order means that all but one concentration are so high they remain constant and are a "part" of the observed rate constant. In this case, the observed rate constant of 1.82 is really $k[\text{S}_2\text{O}_8^{2-}]$. Just solve for k .

$$k = 1.82/0.200 = 9.10$$

010 (part 1 of 3) 10.0 points

A chemical reaction is expressed by the balanced chemical equation



Consider the data

	Initial [A] M	Initial [B] M	Initial rate M/min
1	0.16	0.16	0.00146227
2	0.16	0.32	0.00584909
3	0.32	0.32	0.0116982

Find the rate law for the reaction.

1. $R = k[A] [B]$
2. $R = k[A] [B]^2$ **correct**
3. $R = k[A]^2 [B]$

Explanation:



rate law = ?

$$R = k [A] [B]^2$$

011 (part 2 of 3) 10.0 points

Calculate the value of the specific rate constant.

1. 0.352
2. 0.234
3. 0.25
4. 0.427
5. 0.336
6. 0.208
7. 0.225
8. 0.379
9. 0.371
10. 0.357

Correct answer: 0.357.

Explanation:

value of k (specific rate constant) = ?

$$R = k [A] [B]^2$$

$$k = \frac{R}{[A] [B]^2}$$

$$\begin{aligned}
 &= \frac{0.00146227 \text{ M/min}}{(0.16 \text{ M}) (0.16 \text{ M})^2} \\
 &= 0.357 \text{ min}^{-1} \cdot \text{M}^{-2}
 \end{aligned}$$

012 (part 3 of 3) 10.0 points

If the initial concentrations of both A and B are 0.29 M, at what initial rate is C formed?

1. 0.0279865
2. 0.0256851
3. 0.0249589
4. 0.00870687
5. 0.0190222
6. 0.0200586
7. 0.0336315
8. 0.016192
9. 0.0118528
10. 0.0135992

Correct answer: 0.00870687 M/min.

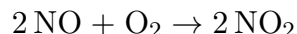
Explanation:

$[A] = 0.29 \text{ M}$ $[B] = 0.29 \text{ M}$
 initial rate that C is formed = ?

$$\begin{aligned}
 R &= k [A] [B]^2 \\
 &= (0.357 \text{ min}^{-1} \cdot \text{M}^{-2}) (0.29 \text{ M}) (0.29 \text{ M})^2 \\
 &= 0.00870687 \text{ M/min}
 \end{aligned}$$

013 10.0 points

We know that the rate expression for the reaction



at a certain temperature is $\text{Rate} = k [\text{NO}]^2 [\text{O}_2]$. We carry out two experiments involving this reaction at the same temperature, but in the second experiment the initial concentration of NO is doubled while the initial concentration of O_2 is halved. The initial rate in the second experiment will be how many times that of the first?

1. 8
2. 1
3. 4
4. 2 correct

Explanation:

$$[\text{NO}]_2 = 2 [\text{NO}]_1; [\text{O}_2]_2 = \frac{1}{2} [\text{O}_2]_1$$

$$\begin{aligned}
 \text{Rate} &= k [\text{NO}]^2 [\text{O}_2] \\
 \frac{\text{Rate}_2}{\text{Rate}_1} &= \frac{k_2 [\text{NO}]_2^2 [\text{O}_2]_2}{k_1 [\text{NO}]_1^2 [\text{O}_2]_1} \\
 &= \frac{(2 [\text{NO}]_1)^2 \frac{1}{2} [\text{O}_2]_1}{[\text{NO}]_1^2 [\text{O}_2]_1} = \frac{4}{2} = 2
 \end{aligned}$$

014 10.0 points

Consider the data collected for a chemical reaction between compounds A and B that is first order in A and first order in B:

	[A]	[B]	rate
	M	M	M/s
1	0.2	0.05	0.1
2	?	0.05	0.4
3	0.4	?	0.8

From the information above for 3 experiments, determine the missing concentrations of A and B. Answers should be in the order of [A] then [B].

1. 0.80 M; 0.20 M **correct**
2. 0.40 M; 0.20 M
3. 0.40 M; 0.10 M
4. 0.80 M; 0.10 M
5. 0.20 M; 0.80 M
6. 1.60 M; 0.40 M

Explanation:

As we know the reaction is first order in A and first order in B, the rate law has the form

$$\text{Rate} = k [\text{A}] [\text{B}]$$

To determine the rate constant k , consider trial 1:

$$\begin{aligned}\text{Rate} &= k[A][B] \\ 0.1 \text{ M/s} &= k(0.2 \text{ M})(0.05 \text{ M}) \\ k &= 10 \text{ M}^{-1} \cdot \text{s}^{-1}\end{aligned}$$

So, in the second trial,

$$\begin{aligned}\text{Rate} &= k[A][B] \\ 0.4 \text{ M/s} &= 10 \text{ M}^{-1} \cdot \text{s}^{-1}[A](0.05 \text{ M}) \\ [A] &= 0.80 \text{ M}\end{aligned}$$

And in the third trial,

$$\begin{aligned}\text{Rate} &= k[A][B] \\ 0.8 \text{ M/s} &= 10 \text{ M}^{-1} \cdot \text{s}^{-1}(0.4 \text{ M})[B] \\ [B] &= 0.20 \text{ M}\end{aligned}$$

015 10.0 points

For a reaction that is zero-order overall,

1. the rate constant is zero.
2. the activation energy is zero.
3. the reactant concentration does not change with time.
4. the rate does not change during the reaction. **correct**

Explanation:

Remember that for a zero order reaction, the reactant is still used up; its concentration will decrease as time passes, but the rate will not be affected. If you double or triple the concentration of that reactant, the rate will not change for a zero order reaction.

016 10.0 points

Consider the reaction: $A + B \rightarrow C$

If it is 1st order in A and 0th order in B, a plot of $\ln[A]$ vs time will have a slope that is

1. slowly decreasing
2. constant **correct**

3. decreasing exponentially

4. slowly increasing

5. increasing exponentially

Explanation:

The first order integrated rate equation shows a linear relationship between the natural log of concentration and time.

017 10.0 points

Consider the reaction:



If it has a half-life of 1.6 s, how long will it take a system with an initial $[\text{H}_2\text{CO}_3]$ of 2 M to reach a $[\text{H}_2\text{CO}_3]$ of 125 mM?

1. 6.4 s **correct**
2. 0.8 s
3. 3.2 s
4. Not enough information is given.
5. 0.4 s

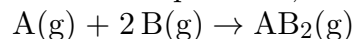
Explanation:

$$\begin{aligned}2 \cdot 0.5^x &= 0.125 \text{ M} \\ x &= 4 \text{ half lives}\end{aligned}$$

$$4 \cdot 1.6 \text{ s} = 6.4 \text{ s}$$

018 10.0 points

At a certain fixed temperature, the reaction



is found to be first order in the concentration of A and zero order in the concentration of B. The reaction rate constant is 0.05 s^{-1} . If 2.00 moles of A and 4.00 moles of B are placed in a 1.00 liter container, how many seconds will elapse before the concentration of A has fallen to 0.30 mole/liter?

1. 34.06

2. 0.05

3. 56.67

4. 37.94 **correct**

5. Not enough information to answer.

Explanation:

019 10.0 points

A reaction



is observed to obey first-order kinetics. Which of the following plots should give a straight line?

1. $\ln[A]$ vs t **correct**2. $[A]$ vs t 3. $[A]$ vs $\frac{1}{t}$ 4. $\ln[A]$ vs $\frac{1}{t}$ 5. $\ln[A]$ vs $\frac{1}{k}$ 6. $[A]$ vs k 7. $[A]$ vs $\frac{1}{k}$ 8. $\ln[A]$ vs k **Explanation:**

The first order integrated rate equation is

$$\ln\left(\frac{[A]_0}{[A]}\right) = a k t$$

$$\ln[A]_0 - \ln[A] = a k t$$

$$\ln[A] = -a k t + \ln[A]_0$$

Since a , k , and $\ln[A]_0$ are constants, this equation is in the form $y = mx + b$, where $y = \ln[A]$ and $x = t$. A straight line is produced by $y = mx + b$.

020 10.0 points

The order of a reaction

Z1) is the product of the powers to which the reactant concentrations are raised in the rate law.

Z2) can't be greater than two since a rate-determining step involving the collision of three or more reactants would practically never occur.

Z3) is the sum of the powers to which the reactant concentrations are raised in the rate law.

Z4) is not related to events on an atomic or molecular scale.

1. Z3 only **correct**

2. Z4 only

3. Z1 only

4. Z3 and Z4 only

5. Z2 only

Explanation:

The overall order of the reaction rate is the sum of the orders of the reactants.

021 10.0 points

A compound decomposes with a half-life of 8.0 s and the half-life is independent of the concentration. How long does it take for the concentration to decrease to one-ninth of its initial value?

1. 32 s

2. 3.6 s

3. 25 s **correct**

4. 72 s

5. 64 s

Explanation:

independent of conc means first order

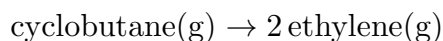
$$k = \ln 2 / t_{1/2} = 0.0866$$

$$\ln(9) / 0.0866 = t$$

$$25.4 = t$$

022 10.0 points

For the reaction



at 800 K, a plot of $\ln[\text{cyclobutane}]$ vs t gives a straight line with a slope of -1.6 s^{-1} . Calculate the time needed for the concentration of cyclobutane to fall to $\frac{1}{16}$ of its initial value.

1. 1.3 s
2. 0.63 s
3. 1.6 s
4. 2.3 s
5. 1.7 s **correct**

Explanation:

023 10.0 points

The initial concentration of the reactant A in a first-order reaction is 1.2 M. After 69.3 s, the concentration has fallen to 0.3 M. What is the rate constant k ? ($\ln 2 = 0.693$)

1. 0.2 s^{-1}
2. not enough information
3. 0.02 s^{-1} **correct**
4. 0.01 s^{-1}

Explanation:

$$t_{1/2} = \frac{69.3 \text{ s}}{2} = 34.7 \text{ s}$$

$$k = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{34.7 \text{ s}} = 0.02 \text{ s}^{-1}$$

024 10.0 points

A reaction is found to be first order with respect to one of the reactant species,

A. When might a plot of $\ln[A]$ vs time not yield a straight line?

1. if the reaction has any significant backward rate
2. all of the above **CORRECT**
3. if the reaction comes to equilibrium
4. when the rate also depends on the concentration of another reactant

Explanation:

If a reaction is first order with respect to species A, a plot of $\ln[A]$ vs time will yield a straight line only if the reaction is zeroth order with respect to all other species (first order overall), or if the concentrations of the other species is effectively constant (pseudo-first order).

Additional, the reaction should not come to equilibrium as this will mean that the forward rate and backward rate are equal. At this point the plot will level off to a constant value. This is the same as having a significant backward rate.

025 10.0 points

Calculate the time required for the activity of a 9.0 mCi cobalt-60 source to decay to 8.5

mCi. The half-life of cobalt-60 is 5.26 years.

1. 4.6 months
2. 2.3 months
3. 0.090 months
4. 10 months
5. 5.2 months **correct**

Explanation: