Kinetics

1.

The rate of the following reaction is 0.720 M/s. What is the relative rate of change of each species in the reaction?

$$A + 4B \longrightarrow 2C$$

$\frac{\Delta[A]}{\Delta t} =$ Number M / s

$\frac{\Delta[B]}{\Delta t} =$ Number M / s

$\frac{\Delta[C]}{\Delta t} =$ Number M / s

2.

The "iodine clock reaction" is a popular chemical demonstration. As part of that demonstration, the I_3^- ion is generated in the reaction

$$S_2O_8^{2-}(aq) + 3I^-(aq) \longrightarrow 2SO_4^{2-}(aq) + I_3^-(aq)$$

In one trial, the unique rate of reaction was 8.69 $\mu mol \cdot L^{-1} \cdot s^{-1}$.

a. What was the rate of reaction of iodide ions?

Number $\mu mol \cdot L^{-1} \cdot s^{-1}$

b. What was the rate of formation of sulfate ions?

Number $\mu mol \cdot L^{-1} \cdot s^{-1}$

3.

The decomposition of N_2O_5 can be described by the equation

$$2N_2O_5(soln) \longrightarrow 4NO_2(soln) + O_2(g)$$

Given these data for the reaction at 45°C in carbon tetrachloride solution, calculate the average rate of reaction for each successive time interval.

t (s)	$[N_2O_5]$ (M)
0	1.76
195	1.56
536	1.26
825	1.05

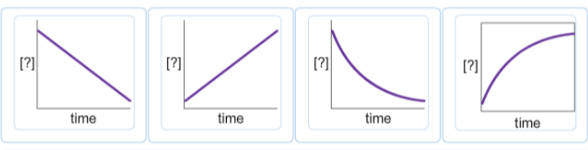
Interval:	0 s to 195 s	195 s to 536 s	536 s to 825 s
Reaction rate:	Number M / s	Number M / s	Number M / s

11.

For the reaction

$$X \longrightarrow Y$$

identify what the graphs of [X] versus time and [Y] versus time would look like for various orders.



13.

Using the given data, calculate the rate constant of this reaction.

$$A + B \longrightarrow C + D$$

Trial	[A] (M)	[B] (M)	Rate (M/s)
1	0.290	0.350	0.0131
2	0.290	0.875	0.0819
3	0.464	0.350	0.0210

$k =$ Number

Units

15.

For the reaction

Map

$$2A(g) + 2B(g) + C(g) \longrightarrow 3G(g) + 4F(g)$$

the following initial rate data was collected, where $[A]_0$, $[B]_0$, and $[C]_0$ are the initial concentrations of A, B, and C, respectively.

Experiment	$[A]_0$ (mmol· L^{-1})	$[B]_0$ (mmol· L^{-1})	$[C]_0$ (mmol· L^{-1})	Initial rate (mmol· L^{-1}· s^{-1})
1	13.0	100.0	300.0	2.60
2	26.0	100.0	225.0	5.20
3	26.0	200.0	75.0	20.8
4	13.0	100.0	150.0	2.60

Identify the order of each reactant.

Reactant A

zero-order reactant
first-order reactant
second-order reactant

Reactant B

zero-order reactant
first-order reactant
second-order reactant

Reactant C

zero-order reactant
first-order reactant
second-order reactant

What is the overall order of the reaction?

Number

Continued belo

Write the rate law for the reaction where *k* is the rate constant.

rate =

Con

Calculate the rate constant, *k*, and identify its units.

$k =$ Number

s^{-1}
$L \cdot mmol^{-1} \cdot s^{-1}$
$L^2 \cdot mmol^{-2} \cdot s^{-1}$

Con

Determine the initial rate of the reaction when $[A]_0$ = 4.27 mmol· L^{-1}, $[B]_0$ = 0.130 mmol· L^{-1}, and $[C]_0$ = 12.0 mmol· L^{-1}.

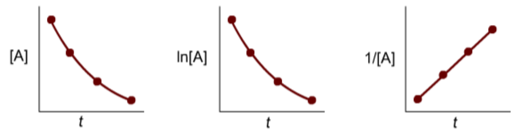
Number $\text{mmol} \cdot L^{-1} \cdot s^{-1}$

Correct.

21.

For A $\longrightarrow$ **products**, time and concentration data were collected and plotted as shown here.

[A] (M)	*t* (s)
0.900	0.0
0.474	30.0
0.321	60.0
0.243	90.0



Determine the reaction order, the rate constant, and the units of the rate constant.

order = Number

k = Number   Units

22.

The following data were collected at 780 K for the reaction $H_2(g) + I_2(g) \longrightarrow 2HI(g)$

Time (s)	$[I_2]$ (mmol·L^{-1})
0	1.00
1.0	0.43
2.0	0.27
3.0	0.20
4.0	0.16

a. Using a graphing calculator or software, plot the data in an appropriate fashion to determine the order of the reaction.

0
1
2

b. From the graph, determine the rate constant for the consumption of I_2.

k = Number

mol·L^{-1}·s^{-1}
s^{-1}
L·mol^{-1}·s^{-1}

23.

After 79.0 min, 39.0% of a compound has decomposed. What is the half-life of this reaction assuming first-order kinetics?

Number **min**

24.

For each of the following cases, identify the order with respect to the reactant, A.

Case (A $\longrightarrow$ **products**)	**Order**
The half-life of A decreases as the initial concentration of A decreases.	Number
A twofold increase in the initial concentration of A leads to a fourfold increase in the initial rate.	Number
A twofold increase in the initial concentration of A leads to a 1.41-fold increase in the initial rate.	Number
The time required for [A] to decrease from $[A]_0$ to $[A]_0/2$ is equal to the time required for [A] to decrease from $[A]_0/2$ to $[A]_0/4$.	Number
The rate of decrease of [A] is a constant.	Number

25.

The rate constant for the reaction is 0.640 $M^{-1} \cdot s^{-1}$ at 200 °C.

$$\text{A} \longrightarrow \text{products}$$

If the initial concentration of A is 0.00960 M, what will be the concentration after 815 s?

Number
M

26.

A particular reactant decomposes with a half-life of 143 s when its initial concentration is 0.296 M.
The same reactant decomposes with a half-life of 233 s when its initial concentration is 0.182 M.

Determine the reaction order.

0
1
2

What is the value and unit of the rate constant for this reaction?

Number

$k =$

27.

Consider the following elementary reaction equation.

$$K(g) + HCl(g) \longrightarrow KCl(g) + H(g)$$

What is the order with respect to K?

Number

What is the overall order of the reaction?

Number

Classifv the reaction as unimolecular, bimolecular, or termolecular.

termolecular
bimolecular
unimolecular

31.

Consider the following mechanism.

step 1:	$A + B \rightleftharpoons C$	equilibrium
step 2:	$C + A \longrightarrow D$	slow
overall:	$2A + B \longrightarrow D$	

Determine the rate law for the overall reaction (where the overall rate constant is represented as *k*).

rate =

32.

The decomposition of nitramide, O_2NNH_2, in water has the following chemical equation and rate law.

$$O_2NNH_2(aq) \longrightarrow N_2O(g) + H_2O(l) \qquad \text{rate} = k\frac{[O_2NNH_2]}{[H^+]}$$

A proposed mechanism for this reaction is

$$(1)\quad O_2NNH_2(aq) \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} O_2NNH^-(aq) + H^+(aq) \qquad \text{(fast equilibrium)}$$

$$(2)\quad O_2NNH^-(aq) \xrightarrow{k_2} N_2O(g) + OH^-(aq) \qquad \text{(slow)}$$

$$(3)\quad H^+(aq) + OH^-(aq) \xrightarrow{k_3} H_2O(l) \qquad \text{(fast)}$$

What is the relationship between the observed value of *k* and the rate constants for the individual steps of the mechanism?

$k =$

Drag the symbols into the numerator or denominator as needed.

Correct.

The mechanism for the reaction described by the equation

$$2N_2O_5(g) \longrightarrow 4NO_2(g) + O_2(g)$$

is suggested to be

$$(1)\quad N_2O_5(g) \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} NO_2(g) + NO_3(g)$$

$$(2)\quad NO_2(g) + NO_3(g) \xrightarrow{k_2} NO_2(g) + O_2(g) + NO(g)$$

$$(3)\quad NO(g) + N_2O_5(g) \xrightarrow{k_3} 3NO_2(g)$$

Assuming that $[NO_3]$ is governed by steady-state conditions, derive the rate law for the production of $O_2(g)$ and enter it in the space below.

rate of reaction= $\dfrac{\Delta[O_2]}{\Delta t} =$

40.

The rate constant for the conversion of cyclopropane into propene was determined at several temperatures.

T (K)	*k* (s^{-1})
750.	1.8×10^{-4}
800.	2.7×10^{-3}
850.	3.0×10^{-2}
900.	0.26

a. Use a graphing calculator or standard graphing software to make an Arrhenius plot and calculate the activation energy for the reaction.

Number

$E_a =$ [ ] $kJ \cdot mol^{-1}$

b. Calculate the rate constant at 564.0 °C.

Number

$k =$ [ ] s^{-1}

41.

For the reversible, one-step reaction shown below,

$$A + A \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} B + C$$

the rate constant for the forward reaction, k_1, is 269 $L \cdot mol^{-1} \cdot min^{-1}$ and the rate constant for the reverse reaction, k_{-1}, is 373 $L \cdot mol^{-1} \cdot min^{-1}$ at a given temperature. The activation energy for the forward reaction is 44.2 $kJ \cdot mol^{-1}$, while the activation energy for the reverse reaction is 20.0 $kJ \cdot mol^{-1}$.

Determine the equilibrium constant, K, of this reaction.

Number

$K =$

Determine whether this reaction is endothermic or exothermic.

- endothermic
- exothermic

What effect will raising the temperature of the reaction have on the rate constants and the equilibrium constant?

- Raising the temperature will increase the reverse rate constant, k_{-1}, more than it will increase the forward rate constant, k_1, resulting in a decrease in the equilibrium constant, K.
- Raising the temperature will increase the forward rate constant, k_1, and reverse rate constant, k_{-1}, equally, leaving the equilibrium constant, K, unchanged.
- Raising the temperature will increase the forward rate constant, k_1, more than it will increase the reverse rate constant, k_{-1}, resulting in an increase in the equilibrium constant, K.
- Raising the temperature will not affect either rate constant or the equilibrium constant, K.

48.

The presence of a catalyst provides a reaction pathway in which the activation every of a reaction is reduced by 43.00 $kJ \cdot mol^{-1}$.

Uncatalyzed: $A \longrightarrow B \quad E_a = 115.00 \ kJ \cdot mol^{-1}$

Catalyzed: $A \longrightarrow B \quad E_a = 72.00 \ kJ \cdot mol^{-1}$

Determine the factor by which the catalyzed reaction is faster than the uncatalyzed reaction at 293.0 K if all other factors are equal.

Number

times faster

Determine the factor by which the catalyzed reaction is faster than the uncatalyzed reaction at 339.0 K if all other factors are equal.

Number

times faster