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AP Chemistry

Name BRACKEN KEY
Period ____

Kinetics Worksheet #4

1972 $2A + 2B \rightarrow C + D$

The following data about the reaction shown above were obtained from three experiments:

Trial	[A]	[B]	Rate of Formation of C (M min^{-1})
1	0.60	0.15	6.3×10^{-3}
2	0.20	0.60	2.8×10^{-3}
3	0.20	0.15	7.0×10^{-4}

(a) What is the rate law for this reaction?

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

(b) What is the numerical value of the rate constant k? What are the units for k?

$6.3 \times 10^{-3} = k[.60]^2[.15]$

$\boxed{k = .12}$

$\frac{\cancel{M}}{min} = k\, M^2 \cdot \cancel{M}$

$\boxed{k = .12\ M^{-2}\ min^{-1}}$

1981 B $A + 2B \rightarrow 3C + D$

For the reaction above, carried out in a solution at 30°C, the following kinetic data were obtained:

Trial	[A]	[B]	Rate of Reaction (M/hr)
1	0.240	0.480	8.00
2	0.240	0.120	2.00
3	0.360	0.240	9.00
4	0.120	0.120	0.500
5	0.240	0.0600	1.00
6	0.0140	1.35	?

(a) Write the rate law expression for this reaction.

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

(b) Calculate the value of the specific rate constant k at 30°C and specify its units.

$8.00 = k[.240]^2[.480]$

$\boxed{k = 289}$

$\frac{\cancel{M}}{hr} = k\, M^2\, \cancel{M}$

$\boxed{k = 289\ M^{-2}\ hr^{-1}}$

(c) Calculate the initial rate of this reaction at 30°C for Trial #6.

RATE $= 289[.0140]^2[1.35]$

$\boxed{\text{RATE} = .0765\ M\ hr^{-1}}$

1987B $2A + B \rightarrow 2C + 2D + E$

Trial	Initial [A]	Initial [B]	Initial Rate of Formation of C (M min^{-1})
1	0.0836	0.202	5.2×10^{-5}
2	0.0836	0.404	2.08×10^{-4}
3	0.0418	0.404	1.06×10^{-4}
4	0.0316	?	1.27×10^{-4}

(a) According to the data shown, what is the rate law for the reaction above?

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

(b) On the basis of the rate law from part (a), calculate the specific rate constant. Specify the units.

$5.2 \times 10^{-5} = k[.0836][.202]^2$

$\frac{M}{min} = k\ \cancel{M}\ M^2$

$\boxed{k = .0152}$

$\boxed{k = .0152\ M^{-2}\ min^{-1}}$

(c) What is the numerical value for the initial rate of disappearance of B for Trial #1 ?

TRICKY PROBLEM: FROM BALANCED EQUATION, WE SEE THAT <u>C</u> FORMS x2's AS FAST AS <u>B</u> DISAPPEARS.

$\boxed{2.6 \times 10^{-5}\ M\ min^{-1}}$

(d) Calculate the initial concentration of **B** for Trial #4.

$1.27 \times 10^{-4} = .0152[.0316][B]^2$

$.264 = [B]^2$

$\boxed{[B] = .514\ M}$

1991 B $2\ ClO_2 + F_2 \rightarrow 2\ ClO_2F$

The following results were obtained when the reaction represented above was studied at 25°C.

Trial	Initial $[ClO_2]$	Initial $[F_2]$	Initial Rate of Increase of ClO_2F (M min^{-1})
1	0.010	0.10	2.4×10^{-3}
2	0.010	0.40	9.6×10^{-3}
3	0.020	0.20	9.6×10^{-3}

(a) Write the rate law expression for the reaction above.

RATE = $k[ClO_2][F_2]$

* compare trials 1 + 2 to know 1st order w/ respect to F_2
* compare trials 2 + 3 to see "offsetting" changes to concentrations have no effect on rate

(b) Calculate the numerical value for the rate constant and specify the units.

$2.4 \times 10^{-3} = k[.010][.10]$

$\boxed{k = 2.4}$

$\frac{M}{min} = k\ M \cdot M$

$\boxed{k = 2.4\ M^{-1}\ min^{-1}}$

(c) In trial #2, what is the initial rate of decrease of $[F_2]$?

F_2 decreases at half the rate ClO_2F appears so

$\frac{9.6 \times 10^{-3}}{2} = \boxed{4.8 \times 10^{-3}\ M\ min^{-1}}$

(d) Which of the two proposed reaction mechanisms shown below is consistent with the rate law developed in part (a) ? <u>Justify your choice.</u>

RATE = $k[ClO_2F_2]$

Can't have intermediate in final answer

Mech #1:

$ClO_2 + F_2 \leftrightarrows ClO_2F_2$ *fast*

$ClO_2F_2 \rightarrow ClO_2F + F$ *slow*

$ClO_2 + F \rightarrow ClO_2F$ *fast*

$k_f[ClO_2][F_2] = k_{rev}[ClO_2F_2]$

$\boxed{RATE = k[ClO_2][F_2]}$

Mech #2:

$F_2 \rightarrow 2\ F$ *slow*

$2\ ClO_2 + 2\ F \rightarrow 2\ ClO_2F$ *fast*

$\boxed{RATE = k[F_2]}$

Mechanism #1, when all intermediates are removed, matches the rate law observed in the data shown above. Therefore, mechanism #1 is consistent.

1997 B $2A + B \rightarrow C + D$

The following results were obtained when the reaction represented above was studied at 25°C.

Experiment	Initial [A]	Initial [B]	Initial Rate of Formation of C (M min^{-1})
1	0.25	0.75	4.3×10^{-4}
2	0.75	0.75	1.3×10^{-3}
3	1.50	1.50	5.3×10^{-3}
4	1.75	?	8.0×10^{-3}

(a) Determine the order of the reaction with respect to A and to B.

RATE IS 1st ORDER WITH RESPECT TO A (compared trials 1 + 2)

RATE IS 1st ORDER WITH RESPECT TO B (compared trials 2 + 3)

(b) Write the rate law for the reaction. Calculate the value of the rate constant, specifying units.

RATE = $k[A][B]$

$4.3 \times 10^{-4} = k[.25][.75]$

$k = .0023$

$\frac{M}{min} = k[M][M]$

$\boxed{k = .0023\ M^{-1}\ min^{-1}}$

(c) Determine the initial rate of change of [A] in Experiment #3. (Interesting curveball says Bracken!)

BASED ON BALANCED EQUATION, A DISAPPEARS 2x's FASTER THAN C APPEARS.

$5.3 \times 10^{-3} \times 2 = \boxed{1.06 \times 10^{-2}\ M\ min^{-1}}$

(d) Determine the initial concentration of **B** in Experiment #4.

$8.0 \times 10^{-3} = .0023[1.75][B]$

$\boxed{B = 1.99\ M}$

(e) Identify which of the reaction mechanisms represented below is consistent with the rate law developed in part (b). Justify your choice.

Mech #1: $A + B \leftrightarrows C + M$ *Fast equilibrium*

$M + A \rightarrow D$ *Slow*

$k_f[A][B] = k_{rev}[C][M]$

$k_{rev}[C]$

RATE = $k[M][A]$

$\boxed{\text{RATE} = k\frac{[A]^2[B]}{[C]}}$

Mech #2: $B \leftrightarrows M$ *Fast equilibrium*

$M + A \rightarrow C + X$ *Slow*

$A + X \rightarrow D$ *Fast*

$k_f[B] = k_{rev}[M]$

RATE = $k[M][A]$

$\boxed{\text{RATE} = k[A][B]}$

consistent with data above

Mech #3: $A + B \leftrightarrows M$ *Fast equilibrium*

$M + A \rightarrow C + X$ *Slow*

$X \rightarrow D$ *Fast*

$k_f[A][B] = k_{rev}[M]$

RATE = $k[M][A]$

$\boxed{\text{RATE} = k[A]^2[B]}$