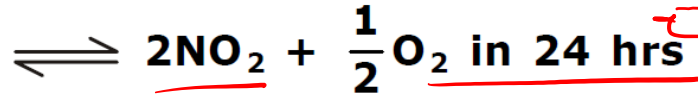


The half life for the reaction N_2O_5



at 30°C . Starting with 10g of N_2O_5

how many grams of N_2O_5 will remain after a period of 96 hours:

- (A) 1.25 g ~~(B) 0.63 g~~
 (C) 1.77 g (D) 0.5 g

$t_{1/2} = 24 \text{ hrs} =$
 $96 \rightarrow 4 \times t_{1/2}$

$0.5^n = 1$ 10 gm

↓ 24 hrs

5 gm

↓ 24 hrs

2.5 gm

↓ 24 hrs

1.25 gm

↓ 24 hrs

0.63 gm ✓

} 96 hrs

The rate constant of a first order reaction is 3×10^{-6} per second. If the initial concentration is 0.10 m, the initial rate of reaction is:

- (A) $3 \times 10^{-5} \text{ ms}^{-1}$ (B) $3 \times 10^{-6} \text{ ms}^{-1}$
(C) $3 \times 10^{-8} \text{ ms}^{-1}$ ~~(D) $3 \times 10^{-7} \text{ ms}^{-1}$~~

$$r = k[A]^1$$

$$r = 3 \times 10^{-6} \times 10^{-1}$$

$$r = 3 \times 10^{-7} \frac{\text{mol}}{\text{litre sec}}$$

A first order reaction with respect to the reactant A has a rate constant of 6 sec^{-1} . If we start with $[A] = 0.5 \text{ mol/litre}$, then in what time the concentration of A becomes 0.05 mol/litre :

(A) 0.384 sec

(B) 0.214 sec

(C) 3.84 sec

(D) 0.402 sec

$$k = 6 \text{ sec}^{-1}$$

$$[A]_{t=0} = 0.5 \frac{\text{mol}}{\text{litre}} = \frac{1}{2} \text{ mol/litre}$$

$$[A]_{t=t} = 0.05 = \frac{0.5}{10} = \frac{1}{10 \times 2} = \frac{1}{20} \text{ mol/litre}$$

$$k = \frac{2.303}{t} \log \left(\frac{[A]_0}{[A]_t} \right)$$

$$t = \frac{2.303}{6} \log \frac{1/2}{1/20}$$

$$t = \frac{2.303}{6} \log 10 \Rightarrow t = \frac{2.303}{6} = 0.384 \text{ sec}$$

The data for the reaction $A + B \rightarrow C$ is

Exp.	[A] ₀	[B] ₀	Initial rate
(1)	0.012	0.035	0.10
(2)	0.024	0.070	0.80
(3)	0.024	0.035	0.10
(4)	0.012	0.070	0.80

The rate law corresponds to the above data is :

(A) Rate = $k[B]^3$

(B) Rate = $k[B]^4$

(C) Rate = $k[A][B]^3$ (D) Rate = $k[A]^2[B]^2$

$$r = k[A]^x[B]^y \quad \text{--- (1)}$$

$$\frac{0.1}{0.8} = \left(\frac{0.035}{0.070} \right)^y$$

$$\left(\frac{1}{2} \right)^3 = \left(\frac{1}{2} \right)^y \Rightarrow y = 3$$

$$\frac{0.1}{0.1} = \left(\frac{0.024}{0.012} \right)^x$$

$$1 = (2)^x \Rightarrow 2^0 = 2^x \Rightarrow x = 0$$

$$r = k[B]^3$$

In a first order reaction the concentration of reactant decreases from 800 mol/dm³ to 50 mol/dm³ is 2 × 10² sec. The rate constant of reaction in sec⁻¹ is :

- (A) 2 × 10⁴ (B) 3.45 × 10⁻⁵
 (C) 1.386 × 10⁻² (D) 2 × 10⁻⁴.

$$t = 2 \times 10^2 = 200 \text{ sec}$$

$$[A]_0 = 800 \text{ mol/dm}^3$$

$$[A]_t = 50 \text{ mol/dm}^3$$

$$k = \frac{2.303}{200} \log \left(\frac{800}{50} \right)$$

$$k = \frac{2.303}{200} \cdot \log (2)^4$$

$$k = \frac{2.303}{\frac{200}{100}} \times \frac{4}{4} \log (2) = \frac{0.693 \times 2}{100} = \underline{\underline{1.386 \times 10^{-2}}}$$

The iodination of acetone occurs by the mechanism, (I) $\text{CH}_3\text{COCH}_3 + \text{H}^+ \xrightarrow{k_1}$ intermediate (slow) (II) Intermediate + $\text{I}_2 \xrightarrow{k_2}$ $\text{CH}_3\text{COCH}_2\text{I} + 2\text{H}^+ + \text{I}^-$ (fast). Which of the following statements is incorrect?

- (A) The rate-determining step is (I) and the rate of reaction is given by Rate = $k_1 [\text{CH}_3\text{COCH}_3] [\text{H}^+]$ — correct
- (B) The reaction is bimolecular →
- (C) ~~The rate of reaction will increase on increasing the concentration of acetone, acid and iodine.~~ x → wrong
- (D) The reaction is first order each in acetone and H^+ and zero order in iodine.

$$r = k [\text{CH}_3\text{COCH}_3]^1 [\text{H}^+]^1 [\text{I}_2]^0$$

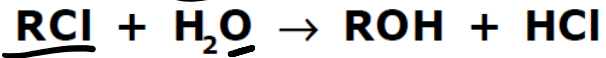
\swarrow acetone \swarrow Acid

Order = 2
Bimolecular

NOBr_2 is an intermediate, which is formed during the slow step of a reaction mechanism. Why is NOBr_2 not used in the rate law ?

- (A) Products are not listed in the rate law.
- (B) The concentration of intermediates can not be measured ✓
- (C) The rate law is determined by the fast step, not the slow step.
- (D) Not enough of NOBr_2 is formed to affect the rate of the reaction.

In a reaction involving hydrolysis of an organic chloride in presence of large excess of water



- (A) Molecularity is 2, order of reaction is also 2 ✓
- ~~(B) Molecularity is 2, order of reaction is 1~~
- ~~(C) Molecularity is 1, order of reaction is 2~~
- ~~(D) Molecularity is 1, order of reaction is also 2~~

order = 1
molecularity = 2

For an elementary reaction, $2A + B \rightarrow C + D$ the molecularity is:

(A) Zero

(B) One

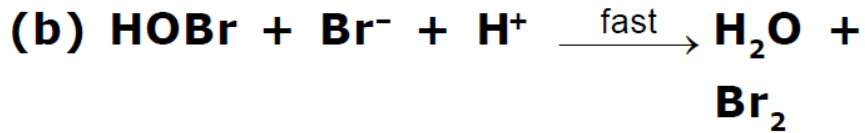
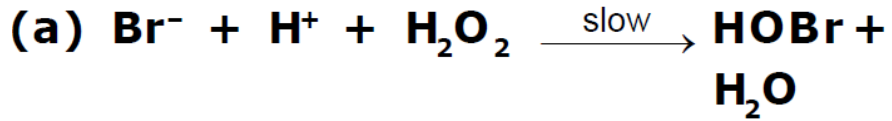
(C) Two

~~(D) Three~~

$$r = k [A]^2 [B]^1$$

Order = 3 = molecularity

Reaction : $2\text{Br}^- + \text{H}_2\text{O}_2 + 2\text{H}^+ \longrightarrow \text{Br}_2 + 2\text{H}_2\text{O}$, takes place in two steps :



The order of the reaction is :

- (A) 3 (B) 6
(C) 5 (D) 0

$$R = k [\text{Br}^-] [\text{H}^+] [\text{H}_2\text{O}_2] \quad \text{--- (1)}$$

final rate law

$$\text{Order} = 1 + 1 + 1 = 3$$

moderate

1	2	3	4	5	6	7	8	9	10
B	D	A	A	C	C	B	B	D	A