**The half life for the reaction $N_2O_5 \rightleftharpoons 2NO_2 + \frac{1}{2}O_2$ in 24 hrs 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

Order = 1

$t_{1/2}$ = 24 hr =

96 → 4 × $t_{1/2}$

$$\left.\begin{array}{c} 10\text{ gm} \\ \downarrow 24\text{ hr} \\ 5\text{ gm} \\ \downarrow 24\text{ hr} \\ 2.5\text{ gm} \\ \downarrow 24\text{ hr} \\ 1.25\text{ gm} \\ \downarrow 24\text{ hr} \end{array}\right\} 96\text{ hr}$$

$\boxed{0.63\text{ gm}}$ ✓

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

(A) $3 \times 10^{-5}\ ms^{-1}$

(B) $3 \times 10^{-6}\ ms^{-1}$

(C) $3 \times 10^{-8}\ ms^{-1}$

(D) $3 \times 10^{-7}\ ms^{-1}$

$r = k[A]^1$

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

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

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 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

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

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

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

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

$t = \frac{2.303}{6} \log \frac{1/\cancel{2}}{1/\cancel{20}_{10}}$

$t = \frac{2.303}{6} \cdot \log 10 \Rightarrow t = \frac{2.303}{6} = 0.38\ sec$

**The data for the reaction A + B → 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 = $k[A]^2[B]^2$

$r = k[A]^x \cdot [B]^y$ — ①

$\frac{0.1}{0.8} = \left(\frac{\cancel{0.035}}{\cancel{0.070}_2}\right)^y$

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

$\frac{\cancel{0.1}}{\cancel{0.1}} = \left(\frac{\cancel{0.024}^2}{\cancel{0.012}}\right)^x$

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

$\boxed{x = 0}$

$\boxed{r = k[B]^3}$

In a first order reaction the concentration of ractant decreases from 800 mol/dm³ to 50 mol/dm³ is $2 \times 10^2$ sec. The rate constant of reaction in sec$^{-1}$ is :

(A) $2 \times 10^4$ (B) $3.45 \times 10^{-5}$

(C) $1.386 \times 10^{-2}$ (D) $2 \times 10^{-4}$.

$t = 2 \times 10^2 = 200$ sec

$[A]_0 = 800$ mol/dm3

$[A]_t = 50$ mol/dm3

$$k = \frac{2.303}{200} \log\left(\frac{\cancel{800}^{16}}{\cancel{50}}\right)$$

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

$$k = \frac{2.303}{\cancel{200}_{100}} \times \cancel{4}^{2} \log(2) = \frac{0.693 \times 2}{100}$$

$$= 1.386 \times 10^{-2}$$

**The iodination of acetone occurs by the mechanism, (I) $CH_3COCH_3 + H^+ \xrightarrow{k_1}$ intermediate (slow) (II) Intermediate + $I_2 \xrightarrow{k_2} CH_3COCH_2I + 2H^+ + I^-$ (fast). Which of the following statements is incorrect ?**

$r = k[CH_3COCH_3]^1 [H^+]^1 [I_2]^0$

acetone Acid

Order = 2

Bimolecular

(A) The rate-determining step is (I) and the rate of reaction is given by Rate $= k_1 [CH_3COCH_3] [H^+]$ — correct

(B) The reaction is bimolecular

(C) The rate of reaction will increase on increasing the concentration of acetone, acid and iodine × → wrong

(D) The reaction is first order each in acetone and $H^+$ and zero order in iodine.

**$NOBr_2$ is an intermediate, Which is formed during the slow step of a reaction mechanism. Why is $NOBr_2$ not ued in the rate law ?**

(A) Products are not listed in the rate law.

(B) The concentration of inter mediates 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.

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

**$RCl + H_2O \rightarrow ROH + HCl$**

(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

$\boxed{\begin{array}{l}\text{order} = 1 \\ \text{molecularity} = 2\end{array}}$

**For an elementary reaction, 2A + B → C + D the molecularity is:**

(A) Zero

(B) One

(C) Two

(D) ~~Three~~

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

Order = 3 = Molecularity

**Reaction : $2Br^- + H_2O_2 + 2H^+ \longrightarrow Br_2 + 2H_2O$, takes place in two steps :**

**(a) $Br^- + H^+ + H_2O_2 \xrightarrow{\text{slow}} HOBr + H_2O$**

**(b) $HOBr + Br^- + H^+ \xrightarrow{\text{fast}} H_2O + Br_2$**

**The order of the reaction is :**

(A) 3 (B) 6

(C) 5 (D) 0

$\boxed{R = k[Br^-]^1[H^+]^1[H_2O_2]^1}$ — ①

final rate law

Order = 1 + 1 + 1

= 3

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moderate

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