Answer on the Question #57954

6. Consider the hypothetical reaction : A + B -> C
This rapid reaction gives the following data:
Experiment [A-] (M) [B-](M) Initial Rate (M/s)
1 0.1 0.50 0.053
2 0.2 0.30 0.127
3 0.4 0.60 1.02
4 0.2 0.60 0.254
5 0.4 0.30 0.509

(i) Write the rate law for this reaction.
(ii) Calculate the average rate constant with proper units.
(iii) Calculate to the rate when [A] = 0.30M and [B] = 0.40M

(i)+(ii) $\boldsymbol{r = k[A]^x[B]^y}$

Trial 2: $r_2 = k[0.2]^x[0.3]^y = 0.127$
Trial 4: $r_4 = k[0.2]^x[0.6]^y = 0.254$

Divide r_4 by r_2:

$2 = 2^y;$

$\boldsymbol{y = 1.}$

Trial 3: $r_3 = k[0.4]^x[0.6]^y = 1.02$
Trial 4: $r_4 = k[0.2]^x[0.6]^y = 0.254$

Divide r_4 by r_3 (gives approximately 4):

$4 = 2^x;$

$\boldsymbol{x = 2.}$

Therefore:

$\boldsymbol{r = k[A]^2[B]}$

Calculate the constant from each trial:

$$\boldsymbol{k = \frac{r}{[A]^2[B]}\ [M]^{-2}s^{-1}}$$

$$k_1 = \frac{0.053\frac{M}{s}}{[0.1\ M]^2[0.5\ M]} = 10.600\ [M]^{-2}s^{-1};$$

$$k_2 = \frac{0.127\frac{M}{s}}{[0.2\ M]^2[0.3\ M]} = 10.583\ [M]^{-2}s^{-1};$$

$$k_3 = \frac{1.02\frac{M}{s}}{[0.4\ M]^2[0.6\ M]} = 10.625\ [M]^{-2}s^{-1}\ ;$$

$$k_4 = \frac{0.254\frac{M}{s}}{[0.2\ M]^2[0.6\ M]} = 10.583\ [M]^{-2}s^{-1}\ ;$$

$$k_5 = \frac{0.509\frac{M}{s}}{[0.4\ M]^2[0.3\ M]} = 10.604\ [M]^{-2}s^{-1}\ .$$

The average constant:

$$\boldsymbol{k_{avg}} = \frac{k_1 + k_2 + k_3 + k_4 + k_5}{5} = \frac{10.600 + 10.583 + 10.625 + 10.583 + 10.604}{5}$$
$$= 10.599[M]^{-2}s^{-1} \approx \mathbf{10.6}[\boldsymbol{M}]^{\mathbf{-2}}\boldsymbol{s}^{\mathbf{-1}}$$

The final rate law:

$$\boldsymbol{r = 10.6\ [A]^2[B]}$$

(iii) When [A] = 0.30M and [B] = 0.40M, rate equals to:

$$\boldsymbol{r = 10.6[M]^{-2}s^{-1} * [0.3M]^2 * [0.4M] = 0.3816\ M/s}$$