#82157 Chemistry, General Chemistry

A particular reactant decomposes with a half-life of 107 s when its initial concentration is 0.360 M. The same reactant decomposes with a half-life of 217 s when its initial concentration is 0.178 M. Calculate the rate constant (k) and reaction order?

Answer:
The reaction is not first-order because the half-life of a first-order reaction is independent of the initial concentration:

$t½ = (\ln(2))/k$

That leaves zero-order and second-order to test.

For a zero-order reaction, the half-life is given by the expression: $t½ = ([A]^0)/2k$

Calculate k for the two conditions given:

(a) 139 s when its initial concentration is 0.293 M
$t½ = ([A]^0)/2k$
139 s = (0.293 M)/2k
k = (0.293 M)/2)*(1/(139 s) = 1.05×10^{-3} mol/L·s

(b) 231 s when its initial concentration is 0.176 M
$t½ = ([A]^0)/2k$
231 s = (0.176 M)/2k
k = (0.176 M)/2)*(1/(231 s) = 3.81×10^{-4} mol/L·s
The values of k are different, so that rules out zero-order.

For a second-order reaction, the half-life is given by the expression: $t½ = 1/((k^*)[A]^0))$

Calculate k for the two conditions given:

(a) 139 s when its initial concentration is 0.293 M
$t½ = 1/((k^*)[A]^0))$
139 s = 1/(k*(0.293 M))
k = 1/((0.293 M)*(139 s)) = 2.46×10^{-2} L/mol·s

(b) 231 s when its initial concentration is 0.176 M
$t½ = 1/((k^*)[A]^0))$
231 s = 1/(k*(0.176 M))
k = 1/((0.176 M)*(231 s)) = $2.46.\times 10^{-2}$ L/mol·s

The values of k are the same, so the reaction is second-order, and k = $2.46.\times 10^{-2}$ L/mol·s.