

Question 1:

Iodine 131 has a half life of 8 days and decays by a first order process (as is always the case for radioactive decay). How much of a 36 g sample of iodine 131 will remain after x days ?

According to the first order kinetics, rate equation of radioactive decay in differential form is:

$$-\frac{d[^{131}I]}{dt} = k[^{131}I]$$

$$\int_{I_0}^I -\frac{d[^{131}I]}{[^{131}I]} = \int_0^t k dt \quad \Rightarrow \quad \ln[^{131}I] - \ln[^{131}I]_0 = k(0 - t) = -kt \quad \Rightarrow \quad [^{131}I] = [^{131}I]_0 \cdot e^{-kt}$$

where $[^{131}I]$ is how much iodine 131 is left after some time t ($t=x$ days), $[^{131}I]_0$ is how much sample was there at the beginning of the radioactive decay (at time $t=0$)

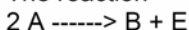
Rate constant k can be calculated from half life $k = \frac{\ln 2}{t_{1/2}}$ so the amount of iodine 131 that has left after x

days would be:

$$[^{131}I] = [^{131}I]_0 \cdot e^{\frac{-\ln 2}{t_{1/2}} t}$$

Question 2:

The reaction



is second order (in A) with the rate constant k (L/mol s). At infinite time $[A] = 0$. How long will it take a solution with $[A]_1$ (mol/L) to decompose such that $[A]_2$ (mol/L) ?

Rate equation in differential form can be expressed as:

$$\text{Rate} = -\frac{1}{2} \frac{dA}{dt} = kA^2$$

Integration of rate equation gives:

$$\int_{A_1}^{A_2} -\frac{1}{2} \frac{dA}{A^2} = \int_{t_1}^{t_2} k dt$$

$$\frac{1}{2} \left(\frac{1}{A_2} - \frac{1}{A_1} \right) = k(t_2 - t_1)$$

$$\Rightarrow t_2 - t_1 = \frac{1}{2k} \left(\frac{1}{A_2} - \frac{1}{A_1} \right)$$

where $t_2 - t_1$ is a time needed for solution with A to decompose from A_1 concentration to A_2

Question 3- The first order rate constant for the gas phase decomposition of dimethyl ether is $3.2 \times 10^{-4} \text{ s}^{-1}$ at 450 degrees Celsius. The reaction is carried out in a constant volume container. Initially only the reactant, dimethyl ether, is present and the pressure is P_0 atm. What is the pressure of the reactant after t seconds? Assume ideal-gas behaviour.

For a first order process: $[A] = [A]_0 e^{-kt}$

Given that you can assume ideal gas behavior, you can simply rewrite the equation above in terms of pressure, i.e. $P = P_0 \exp(-3.2 \times 10^{-4} s^{-1} \cdot t)$, and solve for P.

Question 4- (Note: for illustration purposes, the factor here is 15).

The rate of a certain chemical reaction is found to increase by the factor 15 between T_1 and T_2 degrees C. What is the activation energy (in J) for this reaction ?

Temperature dependence of rate of the reaction is determined by Arrhenius temperature dependence of rate constant :

$$Rate(T) = k(T)[A]^n[B]^m \quad ; \quad k(T) = A \cdot e^{-\frac{E_a}{RT}}$$

$$\Rightarrow Rate(T) = A \cdot e^{-\frac{E_a}{RT}} [A]^n [B]^m$$

Ratio of the rate of the reaction at different temperatures can be expressed as:

$$\frac{Rate(T_2)}{Rate(T_1)} = \frac{e^{-\frac{E_a}{RT_2}}}{e^{-\frac{E_a}{RT_1}}} = e^{-\frac{E_a}{R}(\frac{1}{T_2} - \frac{1}{T_1})} = 15$$

so the activation energy (in J) is

$$E_a = \frac{R \cdot \ln 15}{(\frac{1}{T_1} - \frac{1}{T_2})} \quad \text{where temperature is expressed in degrees K.}$$

Question 5 - The isomerization of a molecule in the gas phase occurs with a rate constant of {k} at {t} degrees C. Calculate the value of the standard molar Gibbs energy of activation (in J. mol⁻¹) for the reaction.

From transition state theory,

$$k(T) = \frac{k_B T}{h} e^{-\Delta G^{o\ddagger}/RT}$$

where $k_B = 1.3807 \times 10^{-23} \text{ JK}^{-1}$ and $h = 6.626 \times 10^{-34} \text{ J s}$. Solving for the standard molar Gibbs energy, by taking the ln of both sides of the equation above:

$$\ln\{k(T)\} = \ln\left\{\frac{k_B T}{h}\right\} - \frac{\Delta G^{o\ddagger}}{RT}$$

$$-\ln\{k(T)\} + \ln\left\{\frac{k_B T}{h}\right\} = \frac{\Delta G^{o\ddagger}}{RT}$$

yields

$$RT \left\{ -\ln\{k(T)\} + \ln\left\{\frac{k_B T}{h}\right\} \right\} = \Delta G^{o\ddagger}$$