Marks 4

- The table below gives the concentrations of C_2H_4O as a function of time at 690 K for the following reaction:

$$C_2H_4O(g) \rightarrow CH_4(g) + CO(g)$$

$[C_2H_4O]$ (M)	time (mins)	$\frac{1}{t}(\ln[A]_0 - \ln[A])$
0.0860	0	
0.0465	50	**0.012298**
0.0355	72	**0.012289**
0.0274	93	**0.012299**
0.0174	130	**0.012291**

The reaction is first order with respect to C_2H_4O.

Use the above data to determine the rate constant and the half-life of the reaction.

For a first order reaction,

$$\ln[A] = \ln[A]_0 - kt \quad \text{or} \quad k = \frac{1}{t}(\ln[A]_0 - \ln[A]) \quad \text{and} \quad t_{1/2} = \frac{\ln 2}{k}$$

A plot of ln[A] *vs*. t has gradient k. Alternatively, the value of $\frac{1}{t}(\ln[A]_0 - \ln[A])$ is included in the table above for each measurement. To 3 significant figures:

$$k = 0.0123 \text{ min}^{-1} \quad \text{and hence} \quad t_{1/2} = \frac{\ln 2}{0.0123} = 56.4 \text{ min}$$

k = **0.0123 min^{-1}**	$t_{½}$ = **56.4 min**

How long does it take for 75% of the C_2H_4O to react?

Using $\ln[A] = \ln[A]_0 - kt$,

$$\ln\left(\frac{[A]_0}{[A]}\right) = kt$$

When 75% has reacted, $[A] = 0.25 \times [A]_0$:

$$\ln(1/0.25) = (0.0123 \text{ min}^{-1})t \quad \text{so } t = 113 \text{ min.}$$

Answer: **113 min**