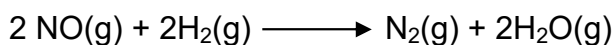


### Exercise n°1

Experiments were conducted to study the rate of the reaction represented by this equation.

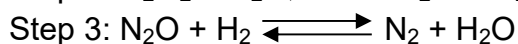
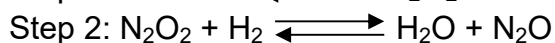


Initial concentrations and rates of reaction are given here.

Experiment	Initial Concentration [NO] (mol L <sup>-1</sup> )	Initial Concentration, [H <sub>2</sub> ] (mol L <sup>-1</sup> )	Initial Rate of Formation of N <sub>2</sub> (mol L <sup>-1</sup> min <sup>-1</sup> )
1	0.006	0.001	$1.8 \times 10^{-4}$
2	0.006	0.002	$3.6 \times 10^{-4}$
3	0.001	0.006	$0.30 \times 10^{-4}$
4	0.002	0.006	$1.2 \times 10^{-4}$

Consider the following questions:

- Determine the order for each of the reactants, NO and H<sub>2</sub>, from the data given and show your reasoning.
- Write the overall rate law for the reaction.
- Calculate the value of the rate constant, k, for the reaction. Include units.
- For experiment 2, calculate the concentration of H<sub>2</sub> remaining when exactly one-half of the original amount of NO had been consumed.
- The following sequence of elementary steps is a proposed mechanism for the reaction.



Based on the data presented, which of these is the rate determining step (all the steps are presented as equilibria: note that the rate determining step must be drawn as an irreversible reaction)? Show that the mechanism is consistent with the observed rate law for the reaction.

### Exercise n°2

Hydrogen iodide, HI, decomposes in the gas phase to produce H<sub>2</sub> and I<sub>2</sub>. The value of the rate constant, *k*, for the reaction was measured at several different temperatures and the data are shown here:

Temperature (K)	<i>k</i> (M <sup>-1</sup> s <sup>-1</sup> )
555	6.23 × 10 <sup>-7</sup>
575	2.42 × 10 <sup>-6</sup>
645	1.44 × 10 <sup>-4</sup>
700	2.01 × 10 <sup>-3</sup>

What is the value of the activation energy (in kJ mol<sup>-1</sup>) and of the pre-exponential factor (in M<sup>-1</sup> s<sup>-1</sup>) for this reaction? (Note: use a regression line method).