Exercise n°1

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

$$2\,NO(g) + 2H_2(g) \longrightarrow N_2(g) + 2H_2O(g)$$

Initial concentrations and rates of reaction are given here.

Experiment	Initial Concentration [NO] (mol L^{-1})	Initial Concentration, [H_2] (mol L^{-1})	Initial Rate of Formation of N_2 (mol L^{-1} min^{-1})
1	0.006	0.001	1.8×10^{-4}
2	0.006	0.002	3.6×10^{-4}
3	0.001	0.006	0.30×10^{-4}
4	0.002	0.006	1.2×10^{-4}

Consider the following questions:

(a) Determine the order for each of the reactants, NO and H_2, from the data given and show your reasoning.
(b) Write the overall rate law for the reaction.
(c) Calculate the value of the rate constant, k, for the reaction. Include units.
(d) For experiment 2, calculate the concentration of H_2 remaining when exactly one-half of the original amount of NO had been consumed.
(e) The following sequence of elementary steps is a proposed mechanism for the reaction.

Step 1: $NO + NO \rightleftharpoons N_2O_2$
Step 2: $N_2O_2 + H_2 \rightleftharpoons H_2O + N_2O$
Step 3: $N_2O + H_2 \rightleftharpoons N_2 + H_2O$

Based on the data presented, which of these is the rate determining step (all the steps are presented as equilibriums: 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_2 and I_2. The value of the rate constant, k, for the reaction was measured at several different temperatures and the data are shown here:

Temperature (K)	***k*** **(M^{-1} s^{-1})**
555	6.23×10^{-7}
575	2.42×10^{-6}
645	1.44×10^{-4}
700	2.01×10^{-3}

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