The following mechanism has been proposed for the gas phase reaction of H$_2$ with ICl:

\[ \text{step 1} \quad \text{H}_2(g) + \text{ICl}(g) \rightarrow \text{HI}(g) + \text{HCl}(g) \]

\[ \text{step 2} \quad \text{HI}(g) + \text{ICl}(g) \rightarrow \text{I}_2(g) + \text{HCl}(g) \]

a. Write the balanced equation for the overall reaction. Identify any intermediates in the reaction.

*Add elementary steps together and cancel anything that shows up on both sides –*

\[ \text{H}_2(g) + 2 \text{ICl}(g) \rightarrow \text{I}_2(g) + 2 \text{HCl}(g) \]

*Intermediate = item that is produced in first step and consumed in second = HI*

b. Write rate laws for each elementary step of the reaction.

\[ \text{rate}_1 = k_1[\text{H}_2][\text{ICl}] \]

\[ \text{rate}_2 = k_2[\text{HI}][\text{ICl}] \]

c. The first step of the reaction is the rate-determining step. Given this information, what is the rate law for the overall reaction?

*The rate law for the overall reaction is the rate law of the rate determining step – we can ignore the rate law for the faster elementary step(s), as they don’t contribute appreciably to the overall speed of the reaction. Therefore, the rate law for this particular reaction is*

\[ \text{rate} = k_1[\text{H}_2][\text{ICl}] \]