Calculating Equilibrium Pressures (for an known mole amount, and an unknown K)

Example:

At a certain temperature a 1.00-L flask initially contained 0.298 mol $PCl_{3\,(g)}$ and 8.70 x 10^{-3} mol $PCl_{5\,(g)}$. After the system had reached equilibrium, 2.00 x 10^{-3} mol $Cl_{2\,(g)}$ was found in the flask. From the balanced equation below, calculate the equilibrium concentrations of ALL species and the value for K.

$$PCl_{5\,(g)} \leftrightarrow PCl_{3\,(g)} + Cl_{2\,(g)}$$

KNOWN MOLES & VOL.

STEP 1:

$$K = \frac{[PCl_3][Cl_2]}{[PCl_5]}$$

STEP 2: INITIAL CONCENTRATIONS

$$[Cl_2]_0 = \text{ZERO}, \quad [PCl_3] = \frac{0.298\text{ mol}}{L}, \quad [PCl_5] = \frac{8.70 \times 10^{-3}\text{ mol}}{L}$$

STEP 3: AT EQUILIBRIUM: RATIO OF REACTANTS / PRODUCTS

$$PCl_5 \rightleftharpoons PCl_3 + Cl_2$$

$$\frac{2.00 \times 10^{-3}\text{ mol}}{L} \rightleftharpoons \frac{2.00 \times 10^{-3}\text{ mol}}{L} + \frac{2.00 \times 10^{-3}\text{ mol}}{L}$$

AMOUNT THAT DECOMPOSED | AMOUNT FORMED

STEP 4: DETERMINE "ACTUAL" CONCENTRATION AMOUNTS!

PCl_5	$\rightleftharpoons$	PCl_3	+	Cl_2
8.70×10^{-3} mol	STARTED	0.298 mol		0 mol
$-\ 2.00 \times 10^{-3}$ mol	CHANGED	$+\ 2.00 \times 10^{-3}$ mol		$+\ 2.00 \times 10^{-3}$ mol
6.70×10^{-3} mol	FINISHED	0.300 mol		2.00×10^{-3} mol

STEP 5: CALCULATE K

$$K = \frac{\left(0.300\ \frac{mol}{L}\right)\left(2.00 \times 10^{-3}\ \frac{mol}{L}\right)}{\left(6.70 \times 10^{-3}\ \frac{mol}{L}\right)} = \boxed{0.0896\ \frac{mol}{L}}$$