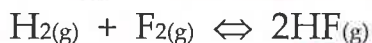


Writing the Equation the other way around

For the following reaction:



a) Write the K_{eq} equation. The $K_{\text{eq}} = 0.25$

$$K_{\text{eq}} = \frac{[\text{HF}]^2}{[\text{H}_2][\text{F}_2]} = 0.25 \quad \leftarrow \text{given}$$

b) Now, write the K_{eq} expression for this reaction, and find the K_{eq} constant value: $2\text{HF}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{F}_2(\text{g})$

$$K_{\text{eq}} = \frac{[\text{H}_2][\text{F}_2]}{[\text{HF}]^2}$$

This is the reciprocal of the equation above.
 \therefore The K_{eq} constant is the reciprocal of 0.25

$$\frac{1}{0.25} = K_{\text{eq}} \\ \therefore K_{\text{eq}} = 4.0$$

VI) Temperature and the K_{eq} Constant

↑ If temperature is increased for an equilibrium system, in what direction does the system shift? *endothermic direction is favored.*

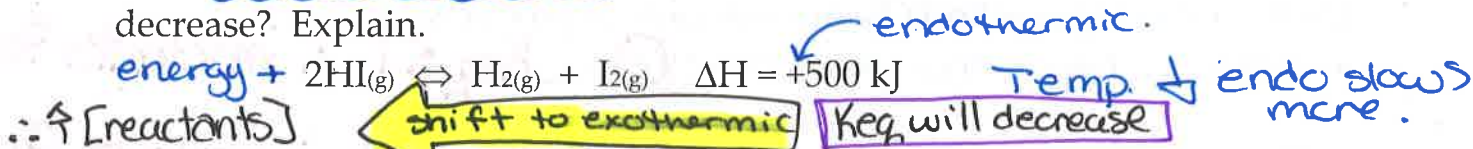
What if temperature is decreased? *exothermic is favored.*

If a shift right results due to temperature change, will the K_{eq} increase or decrease? \rightarrow right shift = \uparrow [products] $\therefore K_{\text{eq}}$ will increase

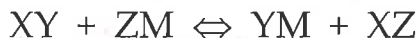
If a shift left results due to a temperature change, will K_{eq} increase or decrease? \leftarrow left shift = \uparrow [reactants] $\therefore K_{\text{eq}}$ will decrease

Example:

If the temperature is decreased in the following system, will K_{eq} increase or decrease? Explain.



Example: Given the following equation and data:



$$K_{\text{eq}} = 60.0 \text{ at } 300^\circ\text{C}$$

$$K_{\text{eq}} = 45.0 \text{ at } 500^\circ\text{C}$$

K_{eq} ↓ with a temp ↑



Is the forward reaction endothermic or exothermic? Explain.

- \uparrow temp increase favors a shift in the endothermic direction.
- For K_{eq} to decrease [reactants] > [products]
- Therefore a shift left is occurring.
- Therefore the rvs rxn is endothermic
(i.e. FWD rxn is exothermic.)