Writing the Equation the other way around

For the following reaction:

$$H_{2(g)} + F_{2(g)} \Leftrightarrow 2HF_{(g)}$$

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

$K_{eq} = \frac{[HF]^2}{[H_2][F_2]} = 0.25$ ← given

b) Now, write the K_{eq} expression for this reaction, and find the K_{eq} constant value: $2HF_{(g)} \Leftrightarrow H_{2(g)} + F_{2(g)}$

$K_{eq} = \frac{[H_2][F_2]}{[HF]^2}$ This is the reciprocal of the equation above. ∴ the K_{eq} constant is the reciprocal of 0.25 → $\frac{1}{0.25} = K_{eq}$ ∴ $K_{eq} = 4.0$

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.

(recall, endo is affected more) ↓ What if temperature is decreased? exothermic is favored.

If a shift right results due to temperature change, will the K_{eq} increase or decrease? → right shift = ↑[products] ∴ K_{eq} will increase

If a shift left results due to a temperature change, will K_{eq} increase or decrease? ← left shift = ↑[reactants] ∴ K_{eq} will decrease

Example:

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

endothermic.

energy + $2HI_{(g)} \Leftrightarrow H_{2(g)} + I_{2(g)}$ $\Delta H = +500$ kJ Temp ↓ endo slows more.

∴ ↑[reactants] ← shift to exothermic $\boxed{K_{eq} \text{ will decrease}}$

Example: Given the following equation and data: FWD → (endo) RVS ← (exo)

$$XY + ZM \Leftrightarrow YM + XZ$$

$K_{eq} = 60.0$ at 300°C

$K_{eq} = 45.0$ at 500°C K_{eq} ↓ with a temp ↑

Is the forward reaction endothermic or exothermic? Explain.

- ↑temp increase favors a shift in the endothermic direction.
- For K_{eq} to decrease [reactants] > [products]
- Therefore a shift left is occuring.
- therefore the rvs rxn is endothermic
  ie: FWD rxn is exothermic.