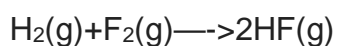


Using standard Heats of formation, calculate the standard Enthalpy change for the following reaction.



Solution:

$$1. \Delta H_{f,298^\circ}(\text{H}_2(\text{g})) = 436 \frac{\text{kJ}}{\text{mol}};$$

$$\Delta H_{f,298^\circ}(\text{F}_2(\text{g})) = 159 \frac{\text{kJ}}{\text{mol}};$$

$$\Delta H_{f,298^\circ}(\text{HF}(\text{g})) = 566 \frac{\text{kJ}}{\text{mol}}.$$

$$2. \Delta H_{\text{substance}} = \sum \Delta H_{\text{products}} - \sum \Delta H_{\text{reagents}}$$

$$3. \sum \Delta H_{\text{reagents}} = \Delta H(\text{H}_2(\text{g})) + \Delta H(\text{F}_2(\text{g}));$$

$$\sum \Delta H_{\text{reagents}} = 436 + 159 = 595 \frac{\text{kJ}}{\text{mol}}.$$

$$4. \sum \Delta H_{\text{products}} = 2 \times \Delta H(\text{HF});$$

$$\sum \Delta H_{\text{products}} = 2 \times 566 = 1132 \frac{\text{kJ}}{\text{mol}}.$$

$$5. \Delta H_{\text{substance}} = 1132 - 595 = 537 \frac{\text{kJ}}{\text{mol}}.$$

$$\text{Answer: } \Delta H_{\text{substance}} = +537 \frac{\text{kJ}}{\text{mol}}.$$