Answer on Question #63403, Chemistry / General Chemistry

Use average bond energies to estimate the energy change (in kJ/mol) for the reaction (all bonds are single bonds except as noted):

$$H_3C\text{-}CH_3 + 7F_2 \rightarrow 2CF_4 + 6HF$$

Solution:

Bond enthalpies (in kJ/mol): C−C (347); C−H (413); H−H (432); F-F (155); C-F (485); H-F (565).

Hess' Law for bond enthalpies is:

$$\Delta H = \Sigma\ E_{\text{reactant bonds broken}} - \Sigma\ E_{\text{product bonds broken}}$$

On the reactant side, we have these bonds broken:

$$\Sigma\ [\text{one C−C bond + six C−H bonds + seven F-F bonds}]$$

$$\Sigma\ [347 + (6 \times 413) + (7 \times 155)] = 3910\ \text{kJ/mol}$$

On the product side, we have these bonds broken:

$$\Sigma\ [\text{eight C−F bond + six H-F bonds}]$$

$$\Sigma\ [(8 \times 485) + (6 \times 565)] = 7270\ \text{kJ/mol}$$

Using Hess' Law, we have:

$$\Delta H = 3910 - 7270 = -3360\ \text{kJ/mol}$$

Answer: -3360 kJ/mol