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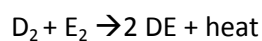
$$17. \quad 100.0 \text{ g} \left(\frac{\text{mol}}{12.01 + 4(1.0079) \text{ g}} \right) = 6.2334 \text{ mol}$$

$$\Delta H = 6.2334 \text{ mol} \left(\frac{-802.0 \text{ kJ}}{\text{mol}} \right) = -4999 \text{ kJ}$$

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15.

Bond	Energy
D-D	400 kJ/mol
E-E	500 kJ/mol



To break the two bonds requires $(1 \text{ mole} * (400 \text{ kJ/mole}) + 1 \text{ mole} * (500 \text{ kJ/mol})) = 900 \text{ kJ}$

But if the reaction is exothermic, then 2 DE releases *more than* 900 KJ. So DE's bond $> 900 \text{ kJ} / 2 \text{ moles} = 450 \text{ kJ/mol}$.

Answer (f)

Note that when calculating ΔH , we would use -450 kJ , but when expressed as a bond energy it is listed as a positive value.