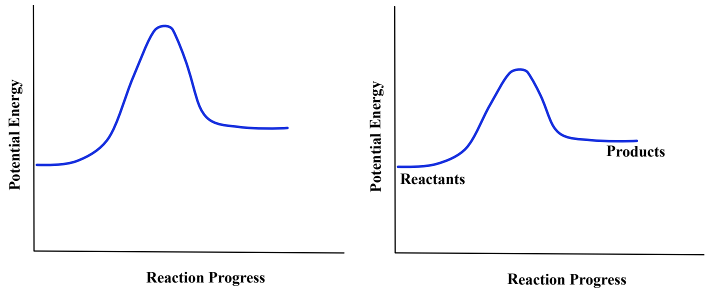
Activation Energy

1. From the two figures, A and B, which reaction is faster? Why?



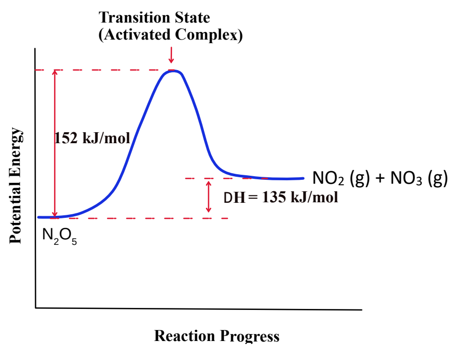
A B

Figure B shows the faster reaction due to the smaller activation energy.

2. Consider the following chemical equation:

$N_2O_5\ (g) \rightarrow NO_2\ (g) + NO_3\ (g) \qquad \Delta H = +135\ kJ/mol$

The activation energy, E_a, is 152 kJ/mol. Draw a labeled energy diagram for this reaction and calculate E_a for the reverse reaction. Does the forward or the reverse reaction have the largest rate constant, k? Is the reaction endothermic or exothermic in the forward direction?



$E_{a,reverse}$ = 152 kJ/mol – 135 kJ/mol = 17 kJ/mol
The reverse reaction has the largest rate constant
The reaction is endothermic

3. A certain first order reaction has a rate constant of $2.63 \times 10^{-2}\ s^{-1}$ at 22.0°C. What is the value of k at 75.0 °C if E_a = 76.9 kJ/mol?

$$ln\left(\frac{k_2}{k1}\right) = -\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

k_2 = ? T_2 = 348.15 K T_1 = 295.15 K $k_1 = 2.63 \times 10^{-2}\ s^{-1}$

$$ln\left(\frac{k_2}{2.63\times10^{-2}s^{-1}}\right) = -\frac{76900\ J/mol}{8.314\frac{J}{mol\cdot K}}\left(\frac{1}{348.15\ K} - \frac{1}{295.15\ K}\right)$$

$$ln\left(\frac{k_2}{2.63\times10^{-2}s^{-1}}\right) = 4.77$$ take antilog of both sides

$$\frac{k_2}{2.63\times10^{-2}s^{-1}} = e^{4.77}$$ **k_2 = 3.10**