Chapter 18 -- Review Problem

Consider the gas-phase reaction: $N_2O_5 + H_2O \longrightarrow 2\ HNO_3$ and the following thermodynamic data.

Compd	$\Delta H°_f$ (kJ/mole)	S° (J/mole-K)
$N_2O_{5(g)}$	11.0	356
$H_2O_{(g)}$	- 242	189
$HNO_{3(g)}$	- 174	156

(a) Decide whether or not the above reaction is spontaneous at 25°C by calculating the value of the *appropriate* thermodynamic quantity.

(b) Calculate the *temperature* (in °C) at which the above reaction should have an equilibrium constant (K_p) equal to 1.00.

(a) $\Delta H° = -117$ kJ

$\Delta S° = -233$ J/K $= -0.233$ kJ/K

$\Delta G° = \Delta H° - T\Delta S° = -117\text{ kJ} - (298\text{ K})(-0.233\text{ kJ/K}) = -47.6\text{ kJ}$

The negative $\Delta G°$ value confirms that the reaction is spontaneous (at least under standard conditions).

(b) The temp where lnK = 1.00 is the temp at which $\Delta G = 0$.

$$\Delta G = -RT\ln K = -RT\ln(1.00) = 0$$

Because ΔH and ΔS are relatively independent of temperature, we can estimate this temp by setting $\Delta G = 0$ as follows.

$\Delta G = \Delta H - T\Delta S = 0$

$\therefore\ T = \Delta H / \Delta S = (-117\text{ kJ}) / (-0.233\text{ kJ/K}) = 502\text{ K}$

$T = 502 - 273 = 229$ °C