Answer on Question #55556 – Chemistry – General chemistry

Question:

The reaction between phosphorous,P4(s), and chlorine, Cl2(g), is exothermic and leads to either PCl3(g) or PCl5(g) depending on the stoichiometric amount of Cl2(g) used. Given the following two chemical equations (EQ 1 and EQ 2) and their associated enthalpy changes for the formation of PCl3(g) and PCl5(g);

EQ 1: P4(s) + 6Cl2(g) \rightarrow 4PCl3(g) = -2439 kJ/mol-rxn

EQ 2: P4(s) + $10Cl2(g) \rightarrow 4PCl5(g) = -3438 \text{ kJ/mol-rxn}$

EQ 3: $PCI5(g) \rightarrow PCI3(g) + CI2(g) = ?$

- a. Calculate the expected enthalpy change for the decomposition of one mole of PCI5(g) shown in equation 3 (EQ 3).
- b. Calculate the expected enthalpy change for the decomposition of 10.50 g of PCI5(g) shown in equation 3 (EQ 3).

Solution:

$$P_{4 (s)} + 6 Cl_{2 (g)} \rightarrow 4 PCl_{3 (g)} \Delta H_{rxn} = -2439 \text{ kJ mol}^{-1}$$
 $P_{4 (s)} + 10 Cl_{2 (g)} \rightarrow 4 PCl_{5 (g)} \Delta H_{rxn} = -3438 \text{ kJ mol}^{-1}$
 $PCl_{5 (g)} \rightarrow PCl_{3 (g)} + Cl_{2 (g)} \Delta H_{rxn} = ?$

$$4 PCl_{5 (g)} \rightarrow P_{4 (s)} + 10 Cl_{2 (g)} \Delta H_{rxn} = 3438 \text{ kJ mol}^{-1} + P_{4 (s)} + 6 Cl_{2 (g)} \rightarrow 4 PCl_{3 (g)} \Delta H_{rxn} = -2439 \text{ kJ mol}^{-1}$$

$$= 4 PCl_{5 (g)} \rightarrow 4 PCl_{3 (g)} + 4 Cl_{2 (g)} \Delta H_{rxn} = 999 \text{ kJ mol}^{-1}$$
 $PCl_{5 (g)} \rightarrow PCl_{3 (g)} + Cl_{2 (g)} \Delta H_{rxn} = 249.75 \text{ kJ mol}^{-1}$

 $Q = -\Delta H_{rxn} n = -\Delta H_{rxn} m/M = -249.75 \text{ kJ mol}^{-1} * 10.50 \text{ g} / 208.22 \text{ g mol}^{-1} = -12.59 \text{ kJ}$

Answer: -12.59 kJ (reaction is endothermic)