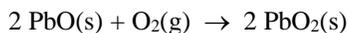


3) When a metal reacts with oxygen a metal oxide forms. For many metals there are several different oxides that can be produced, each with its own formula. Adding oxygen to one metal oxide will often produce a different metal oxide.

Consider the following reaction



substance	ΔH°_f (kJ/mol)	ΔG°_f (kJ/mol)	S° (J/mol·K)
$\text{O}_2(g)$	0.0	0.0	205.0
$\text{PbO}(s)$	- 217.3	- 187.9	68.7
$\text{PbO}_2(s)$	- 277.4	- 217.4	68.6

Thermochemical data for the reactants and products are given above at $T = 298.0 \text{ K}$

a) Using the data in the table, find the values for $\Delta H^\circ_{\text{rxn}}$, $\Delta G^\circ_{\text{rxn}}$, and $\Delta S^\circ_{\text{rxn}}$ for the above reaction at $T = 298.0 \text{ K}$.

$$\begin{aligned} \Delta H^\circ_{\text{rxn}} &= [2 (\Delta H^\circ_f(\text{PbO}_2(s)))] - [2 \Delta H^\circ_f(\text{PbO}(s)) + \Delta H^\circ_f(\text{O}_2(g))] \\ &= [2 (- 277.4)] - [2 (- 217.3) + (0.0)] = - 120.2 \text{ kJ/mol} \end{aligned}$$

$$\begin{aligned} \Delta G^\circ_{\text{rxn}} &= [2 (\Delta G^\circ_f(\text{PbO}_2(s)))] - [2 \Delta G^\circ_f(\text{PbO}(s)) + \Delta G^\circ_f(\text{O}_2(g))] \\ &= [2 (- 217.4)] - [2 (- 187.9) + (0.0)] = - 59.0 \text{ kJ/mol} \end{aligned}$$

$$\begin{aligned} \Delta S^\circ_{\text{rxn}} &= [2 (S^\circ(\text{PbO}_2(s)))] - [2 S^\circ(\text{PbO}(s)) + S^\circ(\text{O}_2(g))] \\ &= [2 (68.6)] - [2 (68.7) + (205.0)] = - 205.2 \text{ J/mol}\cdot\text{K} \end{aligned}$$

b) As discussed in class, a second way for finding $\Delta G^\circ_{\text{rxn}}$ for a chemical reaction is to use the relationship

$$\Delta G^\circ_{\text{rxn}} = \Delta H^\circ_{\text{rxn}} - T\Delta S^\circ_{\text{rxn}}$$

Use this equation to find $\Delta G^\circ_{\text{rxn}}$ from the values for $\Delta H^\circ_{\text{rxn}}$ and $\Delta S^\circ_{\text{rxn}}$ calculated in part a, and check to see if you get the same value for $\Delta G^\circ_{\text{rxn}}$ found directly in part a of the problem.

$$\Delta G^\circ_{\text{rxn}} = \Delta H^\circ_{\text{rxn}} - T\Delta S^\circ_{\text{rxn}} = (- 120.2 \text{ kJ/mol}) - (298.0 \text{ K}) (- 0.2052 \text{ kJ/mol}\cdot\text{K}) = - 59.1 \text{ kJ/mol}$$

That is close to the value for $\Delta G^\circ_{\text{rxn}}$ calculated directly.

c) Is the above reaction spontaneous for standard conditions (yes/no/cannot tell)? Briefly justify your answer.

Since $\Delta G^\circ_{\text{rxn}} < 0$, the reaction is spontaneous for standard conditions. Note that if you calculated the value for $\Delta S^\circ_{\text{univ}}$ you would get a positive result, consistent with the reaction being spontaneous for standard conditions and at $T = 298. \text{ K}$.