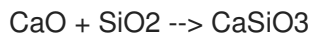


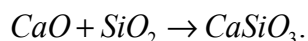
Answer on Question #60855 - Chemistry - Other

Task:

What is the minimum temperature for the following reaction to occur?



Solution:



The standard-state enthalpy of reaction is equal to the sum of the enthalpies of formation of the products minus the sum of the enthalpies of formation of the reactants:

$$\Delta H^\circ(\text{reaction}) = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants}).$$

$$\Delta H^\circ(\text{reaction}) = \Delta H_f^\circ(\text{CaSiO}_3) + \Delta H_f^\circ(\text{SiO}_2) - \Delta H_f^\circ(\text{CaO}).$$

$$\Delta H^\circ(\text{reaction}) = -1584,1 - (-859,3) - (-635,1) = -89,7(\text{kJ} \times \text{mol}^{-1}).$$

The standard entropy of reaction is equal to the sum of the entropies of the products minus the sum of the entropies of the reactants:

$$\Delta S(\text{reaction}) = \sum \Delta S(\text{products}) - \sum \Delta S(\text{reactants}).$$

$$\Delta S(\text{reaction}) = 82 - 42,1 - 39,7 = 0,2(\text{J} \times \text{mol}^{-1}).$$

For a favorable reaction ($\Delta G < 0$), $|\Delta H|$ must be smaller than $|T\Delta S|$, or ΔG must go through zero.

So the temperature where ΔG is zero is the temperature where the spontaneity changes.

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

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

$$\Delta H/\Delta S = T$$

$$T = 89700 \text{ J/mol} / (0,2 \text{ J/mol K}) = -448500 \text{ K}$$

So, T must be greater than -448500 K

