



AP Chemistry 1999 Sample Student Responses

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a. i. The standard entropy change is negative because we go from having 3 gas particles (high random state) to a solid and liquid particles, which are less random, thus entropy decreases and ΔS is negative.

ii. ΔH_{298}° provides the principal driving force for it is negative, which indicates that it is an exothermic reaction that gives off energy. ΔS_{298}° would not be the driving force, because according to the 2nd law of thermodynamics, there is a tendency to increase random ness (entropy) in the universe. If S is negative, that would prove to be a reaction that decreases entropy and thus entropy can't be the principal driving force. ΔH must be ^{the force} an exothermic reaction, which is the motivator for the rx.

iii.
$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

it is an

As temperature increases, ΔG° becomes more positive, as the coefficient $(-T\Delta S^\circ)$ becomes more positive with increasing temperatures for ΔS is negative.

b. i. At room temperature, some reactions may not have the activation energy necessary for their particles to collide and proceed to the transition state. Even though ΔG° may be negative, the reaction may be more spontaneous at higher temperatures as ΔG° does have temperature factored into its equation, $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$. The T included may be greater than room temperature, thus the reaction can still be considered spontaneous.

ii. A catalyst does not have an effect on ΔG° , for it merely makes the reaction

proceed more quickly, but it does not change ΔH° of AS_2 , thus ΔG° stays the same.



a) i) $\Delta S_{\text{net}}^{\circ}$ is negative because $\Delta S^{\circ} = \Delta S_{\text{products}} - \Delta S^{\circ}_{\text{reactants}}$. The products are in more ordered states (liquid and solid) than the reactants. So the value of $\Delta S^{\circ}_{\text{reactants}}$ is bigger than that of $\Delta S^{\circ}_{\text{products}}$, so the result is negative.

$$\text{ii) } \Delta G^{\circ} = \Delta H^{\circ} - T\Delta S^{\circ}$$

$$\Delta G^{\circ} = \Delta H - T(-S)$$

The ΔH , because ΔG must be negative for the reaction to be spontaneous. Since ΔS is negative, ΔH must be a larger negative than $\Delta S \cdot T$ so that $\Delta H - T(-\Delta S)$ is still negative.

iii) ΔG is increased with T because as T goes up, so does $\Delta H - T\Delta S$ because ΔS is negative. So, it becomes less spontaneous.

b) i) A negative ΔG does denote a spontaneous reaction but does not affect the rate of that reaction. So, the reaction does occur spontaneously, that is without prevention, but not necessarily at a measurably fast rate.

ii) The catalyst does not affect ΔG , it only changes rate. ΔG is not related to the rate.

a. i The reaction is spontaneous so ΔG° must be negative,

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

$$\Delta S^\circ = \frac{-\Delta G^\circ - \Delta H^\circ}{T} \quad \Delta S^\circ = \frac{-(-\Delta G^\circ) - (-\Delta H^\circ)}{+T}$$

ii ΔS° is negative

iii ΔH° spontaneous reactions tend to move toward more entropy. Since this reaction does not, the enthalpy change must compensate and be the driving force.

$$\text{iii } \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

ΔS° is negative so an increase in T causes a larger number to be added to ΔH° , ΔG° increases.

b. |

ii A catalyst lowers the activation energy so ΔH° is lower.

If ΔH° is lower the ΔG° is lower. $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$