



THERMODYNAMICS

ENTROPY AND SPONTANEITY

CHEMISTRY 165 // SPRING 2020

PRACTICE PROBLEM 1

Which of the following combinations of entropy changes is not possible?

— *answer* —

	ΔS_{sys}	ΔS_{surr}	ΔS_{univ}	Possible?
(i)	<0	>0	>0	
(ii)	<0	<0	>0	
(iii)	>0	<0	>0	

PRACTICE PROBLEM 1

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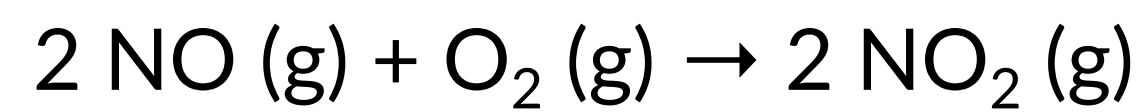
For all three choices, the $\Delta S_{\text{univ}} > 0$, which means that all are spontaneous processes—this is the second law of thermodynamics.

Now, we can work through each combination to see if the following is true:

$$\begin{aligned}\Delta S_{\text{sys}} + \Delta S_{\text{surr}} &= \Delta S_{\text{univ}} \\ \Delta S_{\text{sys}} + \Delta S_{\text{surr}} &> 0\end{aligned}$$

	ΔS_{sys}	ΔS_{surr}	ΔS_{univ}	Possible?
(i)	<0	>0	>0	Yes, if $ \Delta S_{\text{surr}} > \Delta S_{\text{sys}} $.
(ii)	<0	<0	>0	No.
(iii)	>0	<0	>0	Yes, if $ \Delta S_{\text{surr}} < \Delta S_{\text{sys}} $.

PRACTICE PROBLEM 2



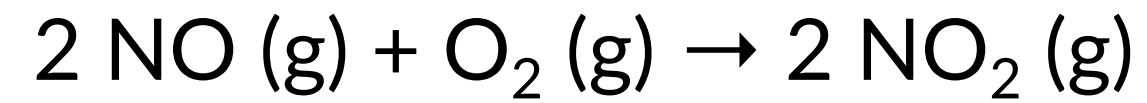
(a) Predict the sign of ΔS° .

(b) Calculate the ΔS° for the reaction using the following standard molar entropies (S°).

	NO (g)	O ₂ (g)	NO ₂ (g)
$S^\circ \left(\frac{\text{J}}{\text{mol} \cdot \text{K}} \right)$	210.7	205.0	240.0

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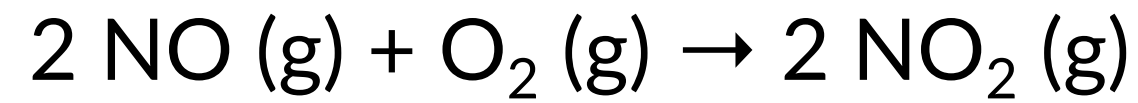
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— answer —

(a) First, recognize that entropy is a measure of the disorder of a system, so the greater the disorder, the greater the entropy.

Because we are going from 3 moles of gaseous species in the reactants to only 2 moles of gaseous species in the products, the entropy (disorder) of the system decreases ($\Delta S < 0$).

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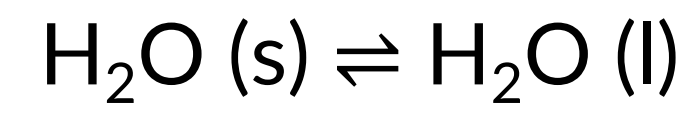
$$(b) \Delta S_{\text{rxn}}^\circ = \sum n_{\text{prod}} S_{\text{prod}}^\circ - \sum n_{\text{react}} S_{\text{react}}^\circ$$

$$= (2 \text{ mol NO}_2) \times \left(240.0 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) - \left[(2 \text{ mol NO}) \times \left(210.5 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) + (1 \text{ mol O}_2) \times \left(205.0 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) \right]$$

$$\Delta S_{\text{rxn}}^\circ = -146.0 \frac{\text{J}}{\text{K}}$$

PRACTICE PROBLEM 3

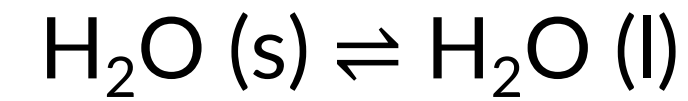
The entropy change for a block of ice melting is 22.1 J/K and the surroundings transfer 6.00 kJ of heat to the system. Is this process spontaneous at 10.0 °C?



— *answer* —

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— answer —

First, realize that for a process to be spontaneous, the entropy change of the universe must be positive:

$$\Delta S_{\text{sys}} + \Delta S_{\text{surr}} = \Delta S_{\text{univ}} > 0$$

We are already given the entropy change of the system: $\Delta S_{\text{sys}} = 22.1 \frac{\text{J}}{\text{K}}$.

Now we can find the entropy change of the surroundings, which can be related to the heat transfer.

$$\begin{aligned}\Delta S_{\text{surr}} &= \frac{q_{\text{surr}}}{T} \\ &= \frac{-6.00 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}}}{(10.0 + 273.15)\text{K}} \\ \Delta S_{\text{surr}} &= -21.19 \frac{\text{J}}{\text{K}}\end{aligned}$$

And finally we can find the entropy change of the universe to determine the process is **spontaneous**.

$$\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} = 22.1 \frac{\text{J}}{\text{K}} + \left(-21.19 \frac{\text{J}}{\text{K}}\right) = 0.9 \frac{\text{J}}{\text{K}} > 0$$