

1) Which of the following processes have a $\Delta S > 0$?

- A) $\text{CH}_3\text{OH(l)} \rightarrow \text{CH}_3\text{OH(s)}$ $\Delta S < 0$ $l \rightarrow s$
 B) $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g})$ $4 \text{mol gas} \rightarrow 2 \text{mol gas}$ $\Delta S < 0$
 C) $\text{CH}_4(\text{g}) + \text{H}_2\text{O(g)} \rightarrow \text{CO(g)} + 3 \text{H}_2(\text{g})$ $2 \text{mol gas} \rightarrow 4 \text{mol gas}$ $\Delta S > 0$
 D) $\text{Na}_2\text{CO}_3(\text{s}) + \text{H}_2\text{O(g)} + \text{CO}_2(\text{g}) \rightarrow 2 \text{NaHCO}_3(\text{s})$ $2 \text{mol gas} \rightarrow 0 \text{mol gas}$ $\Delta S < 0$
 E) All of the above processes have a $\Delta S > 0$
 F) None of the above processes have a $\Delta S > 0$

2) Which of the following processes have a $\Delta S < 0$?

- A) Water freezes
 B) Isopropanol condenses
 C) Methanol (g, at 555 K) \rightarrow methanol (g, at 400 K)
 D) Carbon dioxide (g) \rightarrow carbon dioxide (s)
 E) All of the above processes have a $\Delta S < 0$.
 F) None of the above processes have a $\Delta S < 0$

3) Predict the entropy change (ΔS increases or decreases) if:

You heat a gas from 298K to 500K I

A reaction creates fewer moles of gas product than reactant D

An old refrigerator rusts in a junk yard I

Aluminum oxidizes I

A two dimensional solid is changed to a three dimensional form D

Minerals are deposited from aqueous phase to solid phase D

4a) Consider the following reaction $4\text{NH}_3(\text{g}) + 3 \text{O}_2(\text{g}) \rightarrow 2 \text{N}_2(\text{g}) + 6 \text{H}_2\text{O(g)}$

Given $\Delta H = -1267 \text{ kJ}$. Determine the value of $\Delta S_{\text{surroundings}}$ at 398K. Predict whether or not this reaction will be spontaneous at this temperature.

$$\Delta S_{\text{corr}} = -\Delta H_{\text{system}}/T = -(-1267 \text{ kJ})/398 \text{ K} = 3180 \text{ J/K}$$

$$\Delta G = \Delta H_{\text{sys}} - T \Delta S_{\text{sys}} = -1267 \text{ kJ} - 398 \left(.1788 \text{ kJ/K} \right) = -2047 \text{ kJ}$$

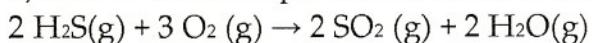
398 K **SPONT.**

4b) Given $\Delta S_{\text{reaction}} = 178.8 \text{ J/K}$, calculate $\Delta S_{\text{universe}}$

$$\Delta S_{\text{univ}} = \Delta S_{\text{rxn}} + \Delta S_{\text{surf}}$$

$$= 178.8 \text{ J/K} + 3180 \text{ J/K} = 3359 \text{ J/K}$$

5) Above what temperature does the following reaction become non-spontaneous?



$$\Delta H = -1036 \text{ kJ}; \Delta S = -153.2 \text{ J/K}$$

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

10 $O = -1036,000 \text{ J} - T(-153.2 \text{ J/K})$

$$T = \frac{1036000 \text{ J}}{153.2 \text{ J/K}} = 6762 \text{ K}$$

6a) Use the following information to demonstrate the temperature dependence of ΔG in 6ii and 6iii below: $\Delta S^\circ = -250 \text{ J/K}$; $\Delta H^\circ = -150 \text{ kJ}$; $T = 27^\circ\text{C}$ and $T = 402^\circ\text{C}$

$$\Delta S^\circ = 250 \text{ J/K}; \Delta H^\circ = 150 \text{ kJ}; T = 27^\circ\text{C} \text{ and } T = 402^\circ\text{C}$$

2 Sig Figs

$$\Delta G = \Delta H - T\Delta S = -150,000 \text{ J} - 300 \text{ K}(-250 \text{ J/K}) = -75 \text{ kJ}$$

$$= +150,000 \text{ J} - 300 \text{ K}(250 \text{ J/K}) = +75 \text{ kJ}$$

$$= -150,000 \text{ J} - 675 \text{ K}(-250 \text{ J/K}) = +19 \text{ kJ} \leftarrow$$

$$= 150,000 \text{ J} - 675 \text{ K}(250 \text{ J/K}) = -19 \text{ kJ} \leftarrow$$

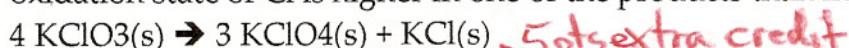
Increasing T for $\Delta H = -$, $\Delta S = -$ ΔG goes from $-$ to $+$
 " " T " " $= +$, " " $= +$ ΔG " " $+$ to $-$

6b) Given the relationship table below, use the information above to fill in the blanks:

	ΔH	ΔS	ΔG	Low Temp	High Temp
i)	-	+	-	S	spontaneous
ii)	-	-	temp dependent	S	NS
iii)	+	+	temp dependent	NS	S
iv)	+	-	+	non-spontaneous	non-spontaneous

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8) Potassium chlorate, a common oxidizing agent in explosives, fireworks, and match heads, undergoes a solid-state redox reaction when heated. In this reaction the oxidation state of Cl is higher in one of the products than in the other.



+5 +7 -1 \Rightarrow 5 pt extra credit

2 5 sig Figs

Cl oxidation	+5	+7	-1	units
S°	143	151	83	J/mol•K
ΔH°_f	-398	-433	-437	kJ/mol

3 5 sig Figs

a) Calculate the Gibbs Free Energy at 25°C for this reaction.

$$\begin{aligned}\Delta H &= \sum n \Delta H^\circ_{\text{prod}} - \sum n \Delta H^\circ_{\text{react}} \\ &= [3(-433 \text{ kJ}) + 1(-437 \text{ kJ})] - [4(-398 \text{ kJ})] \\ &= -1736 \text{ kJ} + 1592 \text{ kJ} \\ &= -144 \text{ kJ } (3 \text{ SIE Fis})\end{aligned}$$

12 $\Delta S = \sum n \Delta S^\circ_{\text{prod}} - \sum n \Delta S^\circ_{\text{react}}$

$$\begin{aligned}&= [3(151 \text{ J/K}) + 1(83 \text{ J/K})] - [4(143 \text{ J/K})] \\ &= (453 \text{ J/K} + 83 \text{ J/K}) - 572 \text{ J/K} \\ &= -36 \text{ J/K } (2 \text{ SIE Fis})\end{aligned}$$

$$\begin{aligned}\Delta G &= \Delta H - T \Delta S = -144000 \text{ J} - 298.15 \text{ K} (-36 \text{ J/K}) \\ &= -144000 \text{ J} + 10733 \text{ J} = -144000 \text{ J} + 11000 \text{ J} \\ &= -133 \text{ kJ}\end{aligned}$$

b) Tell whether the reaction is spontaneous or not and why.

$$\Delta G < 0 \text{ Spont.}$$

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EQUATIONS TO PONDER

$$\Delta H^\circ_{\text{rxn}} = \sum (n \Delta H^\circ_{\text{prod}}) - \sum (n \Delta H^\circ_{\text{react}})$$

$$\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}}$$

$$\Delta S^\circ_{\text{rxn}} = \sum (n \Delta S^\circ_{\text{prod}}) - \sum (n \Delta S^\circ_{\text{react}})$$

$$\Delta S_{\text{surroundings}} = -\Delta H_{\text{system}} / T$$

$$\Delta G^\circ_{\text{rxn}} = \sum (n \Delta G^\circ_{\text{prod}}) - \sum (n \Delta G^\circ_{\text{react}})$$

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

$$\Delta G = \Delta G^\circ + RT \ln Q \text{ (At equilibrium and } \Delta G = 0, \text{ Then } \Delta G^\circ = -RT \ln K)$$