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Chemistry 30

29 max

Unit 1: Energy Changes in Chemical Reactions

Assignment 5: Sections 3-2 to 3-3

1. For each of the following pairs, circle the situation which favors a spontaneous reaction:

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a) endothermic reaction or [✓ exothermic reaction]

b) [✓ negative value of ΔH°] or positive value of ΔH°

c) negative value of ΔS° or [✓ positive value of ΔS°]

d) [✓ increasing entropy] or decreasing entropy

e) positive value of ΔG° or [✓ negative value of ΔG°]

2. Describe the circumstances where:

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a) $\Delta H_f^\circ = 0$ ***(TYPO with QUESTION) Pure elements***

b) $S = 0$ **pure crystals at absolute zero**

c) $\Delta G_f^\circ = 0$ **pure elements**

d) $\Delta G = 0$ **the reaction is at equilibrium**

3. Which one of the following shows and increase in entropy:

[✓ a)] dissolving sugar in a cup of hot tea.
b) arranging a pack of playing cards into suits.
c) building a sand castle on the beach.

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4. Using values of $\Delta G_f°$ from the table provided below, calculate $\Delta G°$ for the following reaction **and** tell whether or not the reaction will occur spontaneously. Show your work clearly. Use the formula $\Delta G = \Sigma\Delta G_{products} - \Sigma\Delta G_{reactants}$

$C_2H_{6(g)} + 2Cl_{2(g)} \rightarrow C_2H_4Cl_{2\ (g)} + 2\ HCl_{(g)}$

Substance	**$\Delta G_f°$ (kJ/mol)**
$C_2H_{6(g)}$	-32.9
$Cl_{2(g)}$	0.0
$C_2H_4Cl_{2(g)}$	-80.3
$HCl_{(g)}$	-95.2

Is the reaction spontaneous?

C_2H_6 (g)	+	2 Cl_2 (g)	→	$C_2H_4Cl_2$ (g)	+	2 HCl (g)
-32.9	+	3 (0)		-80.3	+	2 (-95.2)
	-32.0				-270.7	

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$$\mathbf{\Delta G = \Sigma\Delta G_{products} - \Sigma\Delta G_{reactants}}$$

ΔG = -270.7 – (-32.0) = <u>- 237.8 kJ</u>

1

Because G is negative we know the reaction will be spontaneous

5. Calculate $\Delta G°$ using the formula $\Delta G = \Delta H - T\Delta S$

Also, for each question, tell whether or not the reaction will be spontaneous.

Values for ΔH and ΔS are given. All reactions take place at 25°C (298 K). Remember to convert ΔS values to kJ.

a) $CH_3OH_{(\ell)} + 1½\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$

ΔH = -638.4 kJ ΔS = 156.9 J / K

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

= -638.4 – (298)(0.1569)
= -638.4 – 46.7562
= -685.2 kJ **The reaction will be spontaneous**

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b) $2\ NO_{2(g)} \rightarrow N_2O_{4(g)}$

$\Delta H = -\ 57.2$ kJ $\Delta S = -175.9$ J / K

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

$= -57.2 - (298)(-0.1759)$
$= -57.2+52.4182$
$= -4.78$ kJ **The reaction will be spontaneous**

6. Calculate ΔG° for the following reaction using values of $\Delta G_f°$ obtained from the Table of Thermochemical Data. Will the reaction be spontaneous?

   Use the formula $\Delta G = \Sigma\Delta G_{products} - \Sigma\Delta G_{reactants}$

   $3\ Fe_2O_{3(s)} \rightarrow 2\ Fe_3O_{4(s)} + ½\ O_{2(g)}$

$3\ Fe_2O_3$ (s)	→	$2\ Fe_3O_4$ (s)	+	½ $O_{2(g)}$
3 (-742.2)		2 (-1015.4)	+	½ (0)
-2226.6			-2030.8	

$\Delta G = \Sigma\Delta G_{products} - \Sigma\Delta G_{reactants}$

3

$\Delta G = -2030.8 - (-2226.6) = +195.8$ kJ

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Because G is positive we know the reaction will NOT be spontaneous

7. For a certain spontaneous reaction, the change in enthalpy (ΔH°) is -92.0 kJ and ΔG° = -50.2 kJ at 25ºC. Calculate ΔS.

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

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$-50.2 = -92.0 - (298)(\Delta S)$
$41.8 = -298\Delta S$
$\Delta S = -0.1403$ kJ/K·mol