

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  $\Delta H^\circ$  or positive value of  $\Delta H^\circ$
- c) negative value of  $\Delta S^\circ$  or positive value of  $\Delta S^\circ$
- d) increasing entropy or decreasing entropy
- e) positive value of  $\Delta G^\circ$  or negative value of  $\Delta G^\circ$

2. Describe the circumstances where:

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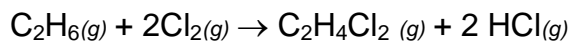
- a)  $\Delta H_f^\circ = 0$  **(TYPO with QUESTION) Pure elements**
- b)  $S = 0$  **pure crystals at absolute zero**
- c)  $\Delta G_f^\circ = 0$  **pure elements**
- d)  $\Delta G = 0$  **the reaction is at equilibrium**

3. Which one of the following shows an 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.

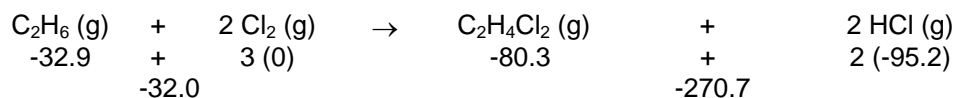
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4. Using values of  $\Delta G_f^\circ$  from the table provided below, calculate  $\Delta G^\circ$  for the following reaction **and** tell whether or not the reaction will occur spontaneously. Show your work clearly. Use the formula  $\Delta G = \Sigma\Delta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



Substance	$\Delta G_f^\circ$ (kJ/mol)
$\text{C}_2\text{H}_6(g)$	-32.9
$\text{Cl}_2(g)$	0.0
$\text{C}_2\text{H}_4\text{Cl}_2(g)$	-80.3
$\text{HCl}(g)$	-95.2

Is the reaction spontaneous?



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

$$\Delta G = -270.7 - (-32.0) = \underline{-237.8 \text{ kJ}}$$

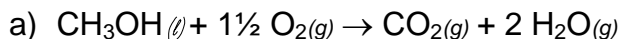
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**Because G is negative we know the reaction will be spontaneous**

5. Calculate  $\Delta G^\circ$  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  $\Delta H$  and  $\Delta S$  are given. All reactions take place at  $25^\circ\text{C}$  (298 K). Remember to convert  $\Delta S$  values to kJ.



$$\Delta H = -638.4 \text{ kJ}$$

$$\Delta S = 156.9 \text{ J / K}$$

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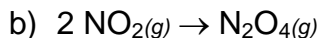
$$\Delta G = \Delta H - T\Delta S$$

$$= -638.4 - (298)(0.1569)$$

$$= -638.4 - 46.7562$$

$$= -685.2 \text{ kJ}$$

**The reaction will be spontaneous**



$$\Delta H = -57.2 \text{ kJ}$$

$$\Delta S = -175.9 \text{ J / K}$$

4

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

$$= -57.2 - (298)(-0.1759)$$

$$= -57.2 + 52.4182$$

$$= -4.78 \text{ kJ}$$

**The reaction will be spontaneous**

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

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



$3 \text{Fe}_2\text{O}_3 (s)$	$\rightarrow$	$2 \text{Fe}_3\text{O}_4 (s)$	$+$	$\frac{1}{2} \text{O}_2(g)$
$3 (-742.2)$		$2 (-1015.4)$	$+$	$\frac{1}{2} (0)$
$-2226.6$			$-2030.8$	

$$\Delta G = \sum \Delta G_{\text{products}} - \sum \Delta G_{\text{reactants}}$$

3

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

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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 ( $\Delta H^\circ$ ) is  $-92.0 \text{ kJ}$  and  $\Delta G^\circ = -50.2 \text{ kJ}$  at  $25^\circ\text{C}$ . Calculate  $\Delta S$ .

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

3

$$-50.2 = -92.0 - (298)(\Delta S)$$

$$41.8 = -298\Delta S$$

$$\Delta S = -0.1403 \text{ kJ/K}\cdot\text{mol}$$