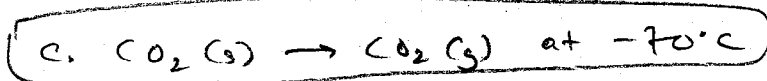


#1 Entropy is the amount of disorder. The probability of figure "a" occurring is slight and therefore has the largest amount of order. Figure "b" has the greatest likelihood of occurring and therefore has the largest amount of disorder (entropy). "c" is somewhere between these two and therefore

$$(a) < (c) < (b)$$

2) Look first for phase changes. In the order  $g \rightarrow l \rightarrow s$ , the molecules are becoming more ordered and there is a decrease in entropy. Going from  $s \rightarrow l \rightarrow g$ , there is less order and an increase in entropy. When the phases on both sides are the same, look for which side has more gaseous molecules when identifying the higher entropy.



3) Going from 2 gas molecules to 1 gas molecule is a decrease in disorder and therefore a decrease in entropy ( $\Delta S = \ominus$ ).  $\Delta H = \text{Bonds broken} - \text{bonds formed}$ .  
Because the bonds were broken, but a bond was formed,  $\Delta H = \ominus$  ( $\Delta H = 0 - 1 \text{ bond formed}$ )

4) For a reaction to be spontaneous at all temperatures,  $\Delta G$  must always be negative:  $\Delta G = \Delta H - T\Delta S$

This means we want  $\Delta H$  to be negative and  $\Delta S$  to be positive because subtracting a positive from a negative will give a negative number, no matter what the temperature.

5)  $A(g) + B(g) \rightarrow C(l)$  always has  $\Delta S = \ominus$

$$\Delta G = \Delta H - T\Delta S = \Delta H + T(\ominus)$$

to get a  $-\Delta G$ ,  $\Delta H$  must be negative. Therefore, C is correct.

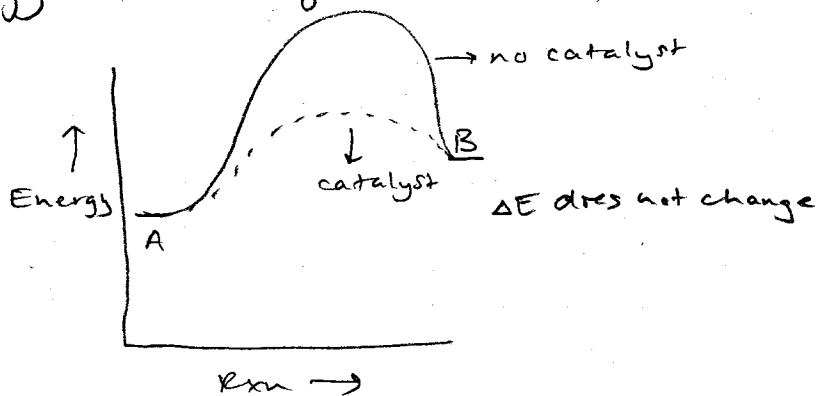
\* A: False. From the equation, entropy is decreasing.

B: Catalysts, which affects kinetics, do not affect thermodynamics

D: High temperatures will only make the " $-T\Delta S$ " (where  $\Delta S$  is  $\ominus$  in this problem) more positive

E: False. Shown it can be above by manipulating  $\Delta H$ .

16) B is correct because, whether a reaction proceeds in 5 seconds or 5 years, it can still be spontaneous. A catalyst simply lowers the activation energy and rate of the reaction, it does not affect the thermodynamics.



D is a common wrong answer. Just because a claim is made that all spontaneous rxns are exothermic, the converse does not need to be true. That is, not all exothermic rxns need be spontaneous.

17)  $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$        $\Delta H^\circ = -1854 \text{ kJ}$        $\Delta S^\circ = -236 \text{ J}\cdot\text{K}^{-1}$   
 $= -236 \times 10^{-3} \text{ kJ}\cdot\text{K}^{-1}$

Remember, we need to use Kelvin and not Celsius

$25^\circ\text{C} = 298^\circ\text{K}$

$\Delta G^\circ = -1854 \text{ kJ} - (298 \text{ K} \cdot -236 \times 10^{-3} \text{ kJ}\cdot\text{K}^{-1})$   
 $= -1784 \text{ kJ}$       A

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

$0 = 136.5 \text{ kJ} - T(0.2875 \text{ kJ/K})$

$-136.5 = -0.2875 T$

$T = 474.78$

$T = 475 \text{ K}$

19) any  $\Delta G > 0$  is nonspontaneous

$\Delta G = \Delta G^\circ + RT \ln Q$

$Q = \frac{1}{(0.2)(0.88)}$  \* do not include liquid in Q expression

$\Delta G = -90.5 + 0.008314(298) \ln \left( \frac{1}{(0.2)(0.88)} \right) = -86.2 \text{ kJ}$   
 ↑  
 R must be in kJ

$\Delta G = -86.2 \text{ kJ}$

#10

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

$$\Delta G = \Delta G^\circ + RT \ln K \rightarrow \text{at equilibrium, } \Delta G = 0$$

so this equation becomes...

$$0 = \Delta G^\circ + RT \ln K$$

$$\Delta G^\circ = -RT \ln K \quad \text{and by substitution}$$

$$\Delta H^\circ - T\Delta S^\circ = -RT \ln K$$

$$\ln K = \frac{\Delta H^\circ - T\Delta S^\circ}{-RT} = \frac{(-304.2 \text{ kJ}) - (498 \text{ K} \cdot (-414.2 \times 10^{-3} \text{ kJ} \cdot \text{K}^{-1}))}{-(0.008314 \cdot 498 \text{ K})}$$

$$\ln K = 23.65$$

$$e^{\ln K} = e^{23.65}$$

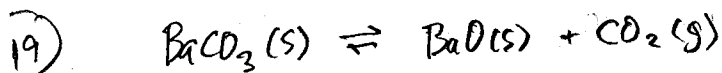
$$K = 1.87 \times 10^{10}$$

#5 11-17 not covered

18)  $\Delta H = \ominus$  b/c bonds are being formed

$\Delta S = \ominus$  b/c molecules are becoming more ordered (fewer gas molecules)

$\Delta G = \ominus$  b/c the question stem states the rxn is spontaneous



$$K = P_{\text{CO}_2} = 6 \times 10^{-39} \text{ at } 298 \text{ K}$$

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

$$\Delta G^\circ = 269.3 \text{ kJ} - 1450 \text{ K} (0.17192 \text{ kJ/mol} \cdot \text{K})$$

$$\Delta G^\circ = 20.0 \text{ kJ}$$

$$\Delta G^\circ = -RT \ln K$$

$$K = e^{-\Delta G^\circ / RT}$$

$$K = e^{-20.0 \text{ kJ} / (0.008314 \text{ kJ/mol} \cdot \text{K})(1450 \text{ K})}$$

$$K = 0.19 \text{ atm}$$

$$P_{\text{CO}_2} = 0.19 \text{ atm}$$

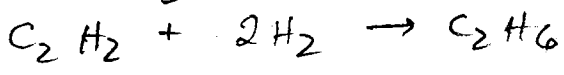
\* Remember,  $\Delta H^\circ$  and  $\Delta S^\circ$  can be apx. to be constant as temp varies, but

$\Delta G^\circ$  changes w/ Temp.

\* If you use the  $\Delta G^\circ$  at 298 K you will get the wrong answer

(20)

Balanced Equation:



$$\Delta G^\circ = -RT \ln K$$

$$\sum \Delta G^\circ_{\text{products}} - \sum \Delta G^\circ_{\text{reactants}}$$

$$\Delta G^\circ = 1 \text{ mol} (-32.89 \text{ kJ/mol}) - [1 \text{ mol} (209.2 \text{ kJ/mol}) + 2 \text{ mol} (0)]$$

$$\Delta G^\circ = -32.89 \text{ kJ} - 209.2 \text{ kJ}$$

$$\Delta G^\circ = -242.09 \text{ kJ}$$

Note:  $\Delta G^\circ$  is  $\ominus$   
 $\therefore K > 1$

$$\Delta G^\circ = -RT \ln K$$

\* Be sure to convert to kJ

$$K = e^{-\Delta G^\circ / RT} = e^{-242.09 \text{ kJ} / 0.008314 \text{ kJ} (298 \text{ K})}$$

$$K = 2.7 \times 10^{42} \rightarrow \text{None of the above}$$

(21)

$$\ln \left( \frac{K_2}{K_1} \right) = \frac{\Delta H^\circ}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$K_{sp} = [Ag^+]^2 [CrO_4^{2-}]$$

$$K_{sp} = 4x^3 = 4(6.7 \times 10^{-5})^3 = 1.2 \times 10^{-12}$$

$$\ln \left( \frac{K_2}{1.2 \times 10^{-12}} \right) = \frac{51.8 \text{ kJ/mol}}{0.008314 \text{ kJ/mol} \cdot K} \left( \frac{1}{298 \text{ K}} - \frac{1}{348.5 \text{ K}} \right)$$

$$\ln K_2 - \ln 1.2 \times 10^{-12} = 3.03$$

$$\ln K_2 = 3.03 + \ln 1.2 \times 10^{-12}$$

$$\ln K_2 = -24.4$$

$$K_2 = 2.5 \times 10^{-11}$$

← This is the Eq constant, we need the molar sol.

$$K_{sp} = 4x^3 = 2.5 \times 10^{-11}$$

$$x = 1.8 \times 10^{-4}$$