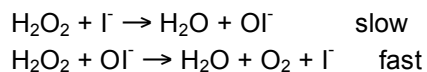


Use the following information to answer questions 1 through 3

A proposed mechanism for the decomposition of hydrogen peroxide by iodide ion is:



1. The rate law consistent with this mechanism is:

- |   |   |
|---|---|
| (A) $R = k[\text{H}_2\text{O}_2]^2$           | (D) $R = k[\text{H}_2\text{O}_2]^2[\text{I}^-]/[\text{H}_2\text{O}]$            |
| (B) $R = k[\text{H}_2\text{O}_2][\text{I}^-]$ | (E) $R = k[\text{H}_2\text{O}][\text{OI}^-]/[\text{H}_2\text{O}_2][\text{I}^-]$ |
| (C) $R = k[\text{H}_2\text{O}][\text{OI}^-]$  |   |

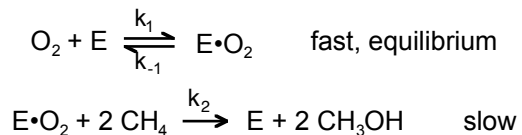
2. A catalyst in this mechanism is:

- |                   |                          |                   |
|-------------------|--------------------------|-------------------|
| (A) $\text{I}^-$  | (C) $\text{O}_2$         | (E) there is none |
| (B) $\text{OI}^-$ | (D) $\text{H}_2\text{O}$ |                   |

3. An intermediate in this mechanism is:

- |                   |                          |                   |
|-------------------|--------------------------|-------------------|
| (A) $\text{I}^-$  | (C) $\text{O}_2$         | (E) there is none |
| (B) $\text{OI}^-$ | (D) $\text{H}_2\text{O}$ |                   |

4. Methanogens are a class of bacteria that metabolize methane. They have an enzyme that allows them to convert methane to methanol. A very simplified mechanism has two steps: First, oxygen binds to the enzyme (E) to produce an enzyme- $\text{O}_2$  complex ( $\text{E}\cdot\text{O}_2$ ), then methane reacts with the  $\text{E}\cdot\text{O}_2$  complex to produce methanol:



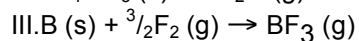
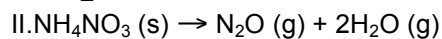
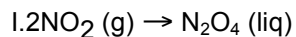
What is the rate law for this reaction ? (Note: k is a combination of  $k_1$ ,  $k_{-1}$  and  $k_2$ )

- |   |   |                               |
|---|---|-------------------------------|
| (A) Rate = $k[\text{O}_2][\text{CH}_4][\text{E}]$   | (C) Rate = $k[\text{O}_2][\text{E}]$    | (E) Rate = $k[\text{CH}_4]^2$ |
| (B) Rate = $k[\text{O}_2][\text{CH}_4]^2[\text{E}]$ | (D) Rate = $k[\text{E}][\text{CH}_4]^2$ |                               |

5. A reaction with an activation energy of  $E_a = 100 \text{ kJ/mol}$  has a rate constant  $k = 10 \text{ s}^{-1}$  at a temperature of  $20^\circ\text{C}$ . What is the rate of this reaction (in  $\text{s}^{-1}$ ) at  $50^\circ\text{C}$  ?

- (A) 7.1 (B) 10.0 (C) 14.0 (D) 450 (E) 3300

6. Predict the sign of  $\Delta S^\circ_{\text{rxn}}$  for the three reactions below:



	<u>I</u>	<u>II</u>	<u>III</u>		<u>I</u>	<u>II</u>	<u>III</u>
(A)	+	+	+	(D)	-	-	+
(B)	-	+	-	(E)	+	-	+
(C)	+	+	-				

7. Under what circumstances will  $\Delta G$  for a chemical reaction always be positive?

(A) An endothermic reaction that generates fewer moles of gaseous products than gaseous reactants.

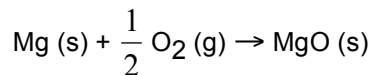
(B) An exothermic reaction that generates solids from liquid reactants.

(C) An exothermic reaction that generates fewer moles of gaseous products than gaseous reactants.

(D) An endothermic reaction that generates more moles of gaseous products than gaseous reactant.

(E) both (B) and (C) above.

For questions 8 and 9 use the thermodynamic data at 298 K given below and the following reaction.



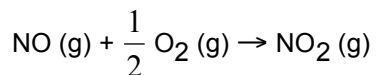
	$\Delta H^\circ_f$ (kJ/mol)	$\Delta G^\circ_f$ (kJ/mol)	$S^\circ$ (J/mol K)
Mg (s)	0	0	33
O <sub>2</sub> (g)	0	0	205
MgO (s)	-602	-569	27

8. What is  $\Delta S^\circ_{\text{rxn}}$  in J/mol K?

- (A) -175 (C) -27 (E) -108  
(B) 27 (D) 108

9. If magnesium metal is ignited in an oxygen atmosphere in an isolated system at 300 K:
- (A) no reaction will take place because  $\Delta S_{\text{rxn}}^{\circ}$  is negative.
  - (B) no reaction will take place because  $\Delta S_{\text{surroundings}} = 0$ .
  - (C) a reaction will take place because  $\Delta S_{\text{universe}}$  is positive.
  - (D) no reaction will take place because  $\Delta G_{\text{reaction}}^{\circ}$  is positive.
  - (E) a reaction will take place because  $\Delta S_{\text{surroundings}}$  is negative.

Questions 10 and 11 refer to the reaction



for which  $\Delta H_{\text{rxn}} = -57 \text{ kJ/mol}$  and  $\Delta S_{\text{rxn}} = -73 \text{ J/(mol K)}$  at 298 K.

10. What is  $\Delta G_{\text{rxn}}$  in kJ/mol at 25 °C for this reaction?

(A) -35    (B) -79    (C) -22    (D) 68    (E) 85

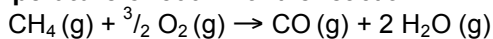
11. What is the temperature in °C above/below which this reaction would have  $K_p$  greater than one?

(A) below 508 °C                      (D) above 128 °C  
 (B) above 508 °C                      (E) There is none.  
 (C) below 128 °C

Thermodynamic values for Problems 12 and 13:

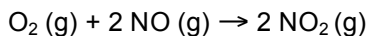
Species	$\Delta H_f^\circ(298\text{ K})$ kJ/mol	$S^\circ(298\text{ K})$ J/(K mol)	$\Delta G_f^\circ(298\text{ K})$ kJ/mol
C(s, graphite)	0	5.6	0
CH <sub>4</sub> (g)	-74.87	186.26	-50.8
C <sub>2</sub> H <sub>4</sub> (g)	52.47	219.36	68.35
C <sub>2</sub> H <sub>6</sub> (g)	-83.85	229.2	-31.89
CO(g)	-110.53	197.67	-137.17
CO <sub>2</sub> (g)	-393.51	213.74	-394.36
H <sub>2</sub> (g)	0	130.7	0
H <sub>2</sub> O(g)	-241.83	188.84	-228.59
H <sub>2</sub> O(l)	-285.83	69.95	-237.15
NO(g)	90.29	210.76	86.58
NO <sub>2</sub> (g)	33.1	240.04	51.23
O <sub>2</sub> (g)	0	205.07	0

12. Use the thermodynamic values above to calculate  $\Delta G_{\text{rxn}}$  in kJ/mol  
at a temperature of 500 K for the reaction



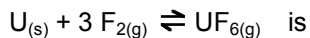
- (A) -275      (B) -280      (C) -479      (D) -544      (E) -560

13. Use the thermodynamic values at the top of this page to calculate the equilibrium constant  $K_p$   
at 298 K for the reaction



- (A)  $4.0 \times 10^{-13}$     (B)  $6.4 \times 10^{-7}$     (C)  $4.8 \times 10^{-4}$     (D)  $1.6 \times 10^6$     (E)  $2.5 \times 10^{12}$

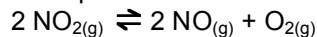
14. The equilibrium constant  $K_c$  for the reaction



- (A)  $K_c = \frac{[\text{UF}_6]}{[\text{U}][\text{F}_2]}$       (D)  $K_c = \frac{[\text{UF}_6]}{[\text{F}_2]}$   
 (B)  $K_c = \frac{[\text{UF}_6]}{[\text{U}][3\text{F}_2]}$       (E)  $K_c = \frac{[\text{UF}_6]}{[\text{F}_2]^3}$   
 (C)  $K_c = \frac{[\text{UF}_6]}{[\text{U}][\text{F}_2]^3}$

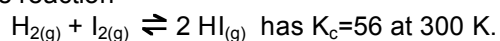
15. At a particular temperature, an equilibrium mixture contains the following concentrations of gases:  $[\text{NO}_2] = 0.32 \text{ M}$ ,  $[\text{NO}] = 10^{-4} \text{ M}$ ,  $[\text{O}_2] = 0.045 \text{ M}$ .

What is the equilibrium constant  $K_c$  for the reaction



- (A)  $4.5 \times 10^{-10}$  (D)  $1.4 \times 10^{-5}$   
 (B)  $1.4 \times 10^{-9}$  (E)  $4.5 \times 10^{-5}$   
 (C)  $4.4 \times 10^{-9}$

16. The reaction



A 1 liter flask is filled with 0.2 moles of  $\text{H}_2$ , 0.2 moles of  $\text{I}_2$  and 0.4 moles of  $\text{HI}$ .

Will any reaction occur? If so, is  $\text{HI}$  produced or consumed?

- (A) No reaction will occur  
 (B) A reaction will occur;  $\text{HI}$  will be produced  
 (C) A reaction will occur;  $\text{HI}$  will be consumed

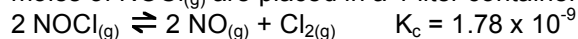
17. 2 moles of  $\text{NH}_4\text{Cl}_{(s)}$  are put into an evacuated 1 liter container at 550 K, and the following reaction occurs:



At equilibrium,  $[\text{NH}_{3(g)}] = 2.2 \times 10^{-3} \text{ M}$ . What is  $K_c$  for the reaction?

- (A)  $2.2 \times 10^{-3}$  (B)  $2.4 \times 10^{-6}$  (C)  $4.8 \times 10^{-6}$  (D)  $9.6 \times 10^{-6}$  (E)  $1.9 \times 10^{-5}$

18. 0.1 moles of  $\text{NOCl}_{(g)}$  are placed in a 1 liter container at 372 K. The following reaction occurs:

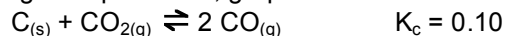


What is the  $\text{Cl}_2$  concentration at equilibrium?

(Hint:  $K_c$  is so small that very little of the  $\text{NOCl}$  decomposes)

- (A)  $6.67 \times 10^{-6}$  (D)  $1.64 \times 10^{-4}$   
 (B)  $1.33 \times 10^{-5}$  (E)  $3.28 \times 10^{-4}$   
 (C)  $2.66 \times 10^{-5}$

19. At high temperatures, graphite reacts with  $\text{CO}_2$  to produce  $\text{CO}$ :

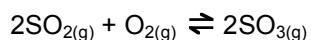


If 0.07 moles of  $\text{CO}_2$  are placed in a 1 liter container, what is the  $\text{CO}$  concentration at equilibrium?

(Note:  $K_c$  is fairly large, so a significant amount of the  $\text{CO}_2$  reacts)

- (A) 0.022 M (D) 0.062 M  
 (B) 0.031 M (E) 0.13 M  
 (C) 0.043 M

Questions 20 through 22 refer to the following gas phase equilibrium for which  $K_c = 12$  at 1100K and  $\Delta H^\circ = -198 \text{ kJ/mol}$ .



20. Increasing the concentration (the pressure) on an equilibrium mixture by decreasing the volume at constant temperature would cause:

- (A)  $K$  to decrease and the amount of  $\text{O}_{2(g)}$  to increase.
- (B)  $K$  to decrease and the amount of  $\text{O}_{2(g)}$  to decrease.
- (C)  $K$  to increase and the amount of  $\text{O}_{2(g)}$  to decrease.
- (D) no change in  $K$  but an increase in the amount of  $\text{O}_{2(g)}$ .
- (E) no change in  $K$  but a decrease in the amount of  $\text{O}_{2(g)}$ .

21. Addition of  $\text{SO}_{3(g)}$  to an equilibrium mixture of the three gases at constant volume and temperature

would cause:

- (A)  $K$  to decrease and the amount of  $\text{O}_{2(g)}$  to increase.
- (B)  $K$  to decrease and the amount of  $\text{O}_{2(g)}$  to decrease.
- (C)  $K$  to increase and the amount of  $\text{O}_{2(g)}$  to decrease.
- (D) no change in  $K$  but an increase in the amount of  $\text{O}_{2(g)}$ .
- (E) no change in  $K$  but a decrease in the amount of  $\text{O}_{2(g)}$ .

22. An equilibrium mixture of the three gases is initially at 1100K. The temperature is increased to 1300K, at constant volume. This would cause:

- (A)  $K$  to decrease and the amount of  $\text{O}_{2(g)}$  to increase.
- (B)  $K$  to decrease and the amount of  $\text{O}_{2(g)}$  to decrease.
- (C)  $K$  to increase and the amount of  $\text{O}_{2(g)}$  to decrease.
- (D) no change in  $K$  but an increase in the amount of  $\text{O}_{2(g)}$ .
- (E) no change in  $K$  but a decrease in the amount of  $\text{O}_{2(g)}$ .