86. $$N_{2(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)}$$

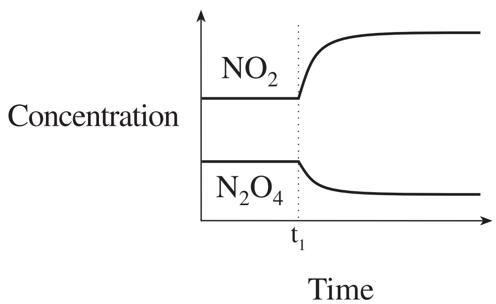
Which of the following factors will not alter the position of equilibrium?

A. a pressure decrease
B. a temperature increase
C. the presence of a catalyst
D. the addition of more $N_{2(g)}$

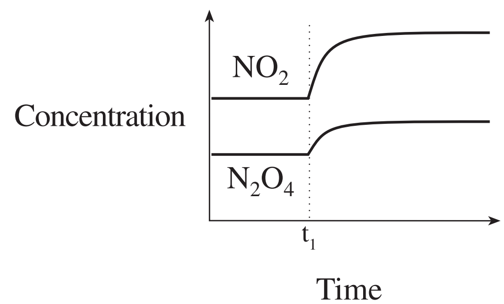
87. $$N_2O_{4(g)} + \text{energy} \rightleftarrows 2NO_{2(g)}$$

Which of the following graphs shows the result of increasing the temperature at time t_1 ? **(1 mark)**

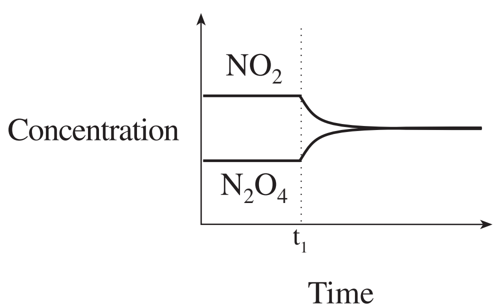
A.



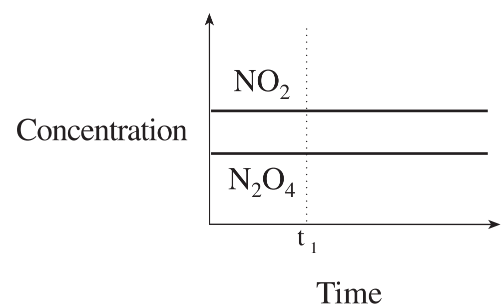
B.



C.



D.



88. $$N_2O_{4(g)} \rightleftarrows 2NO_{2(g)}$$

	Initial		Equilibrium	
	$[N_2O_4]$	$[NO_2]$	$[N_2O_4]$	$[NO_2]$
Trial 1	0.0400	0.0000	0.0337	0.0125
Trial 2	0.0200	0.0600	0.0429	0.0141

Which of the following represents the K_{eq} value?

A. 4.64×10^{-3}
B. 3.71×10^{-1}
C. 7.42×10^{-1}
D. 2.16×10^{2}

89. Which of the following is **least** likely to favour the formation of products?

A. $2H_2O_{(g)} \rightleftarrows 2H_{2(g)} + O_{2(g)}$ $K_{eq} = 7.3 \times 10^{-18}$

B. $N_2O_{(g)} + NO_{2(g)} \rightleftarrows 3NO_{(g)}$ $K_{eq} = 4.2 \times 10^{-4}$

C. $N_2O_{4(g)} \rightleftarrows 2NO_{2(g)}$ $K_{eq} = 4.5$

D. $SO_{2(g)} + NO_{2(g)} \rightleftarrows NO_{(g)} + SO_{3(g)}$ $K_{eq} = 85$

90. Consider the following equilibrium:

$$N_{2(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)}$$

What is the final result of adding some NH_3 gas to the system at constant volume?

A. K_{eq} increases.

B. $[H_2]$ decreases.

C. $[NH_3]$ decreases.

D. K_{eq} remains unchanged.

91.

$$2CO_{(g)} + O_{2(g)} \rightleftarrows 2CO_{2(g)}$$

A container is initially filled with CO and O_2. How will the $[CO]$ and $[CO_2]$ change as the system reaches equilibrium?

	$[CO]$	$[CO_2]$
A.	increase	decrease
B.	increase	increase
C.	decrease	decrease
D.	decrease	increase

92. Two experiments were performed involving the following equilibrium. The temperature was the same in both experiments.

$$H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$$

In experiment A, $1.0\,M\ H_2$ and $1.0\,M\ I_2$ were initially added to a flask and equilibrium was established. In experiment B, $2.0\,M\ HI$ was initially added to a second flask and equilibrium was established. Which of the following statements is always true about the equilibrium concentrations?

A. $[H_2]$ equals $[HI]$ in experiment A.

B. $[HI]$ equals $2[H_2]$ in experiment A.

C. $[HI]$ in experiment A equals $[HI]$ in experiment B.

D. $[HI]$ in experiment A equals $\frac{1}{2}[I_2]$ in experiment B.

93. Which of the following reactions is accompanied by an increase in enthalpy?

A. $2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)} + 113\,kJ$

B. $2H_{2(g)} + O_{2(g)} - 484\,kJ \rightleftarrows 2H_2O_{(g)}$

C. $2SO_{3(g)} \rightleftarrows 2SO_{2(g)} + O_{2(g)} \qquad \Delta H = +197\,kJ$

D. $4HCl_{(g)} + O_{2(g)} \rightleftarrows 2H_2O_{(g)} + 2Cl_{2(g)} \qquad \Delta H = -111.4\,kJ$

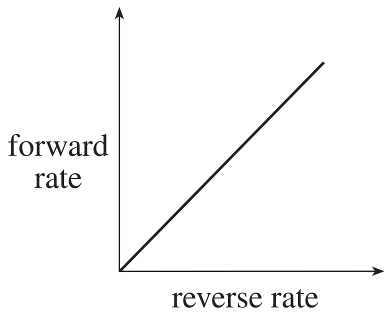
94. Two substances are mixed and no reaction occurs. With respect to enthalpy and entropy, which of the following could explain why no reaction occurs?

	Enthalpy	Entropy
A.	increases	increases
B.	increases	decreases
C.	decreases	increases
D.	decreases	decreases

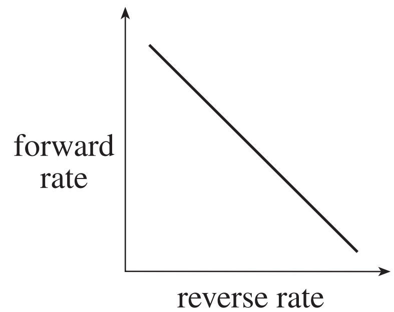
95. $$N_{2(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)}$$

Which of the following diagrams represents what happens to the forward and reverse reaction rates when the catalyst Fe_3O_4 is added?

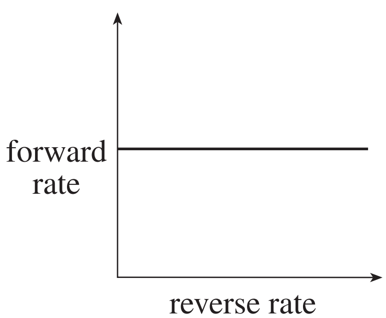
A.



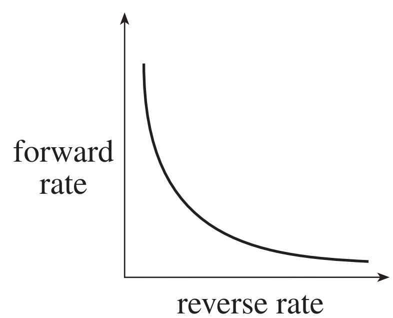
B.



C.

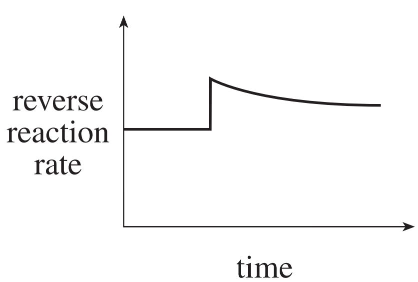


D.

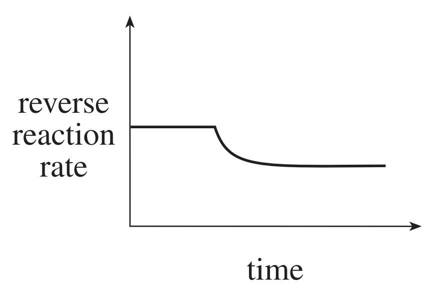


96. Temperature is gradually decreased then held constant in an exothermic equilibrium. Which of the following represents the change in the reverse reaction rate?

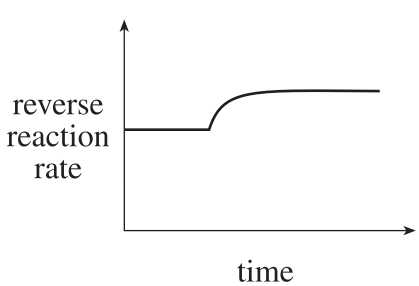
A.



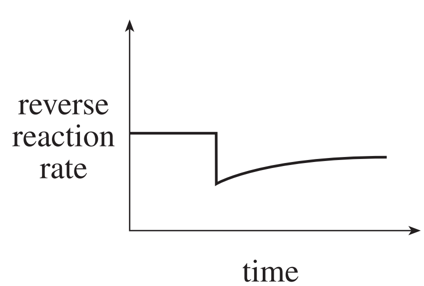
B.



C.



D.



97.

$$CaCO_{3(s)} + 2HF_{(g)} \rightleftarrows CaF_{2(s)} + H_2O_{(g)} + CO_{2(g)}$$

Which of the following represents the equilibrium $[H_2O]$?

A. $[H_2O] = \dfrac{[HF]^2}{K_{eq}[CO_2]}$ C. $[H_2O] = \dfrac{[HF]^2[CaCO_3]}{K_{eq}[CO_2][CaF_2]}$

B. $[H_2O] = \dfrac{K_{eq}[HF]^2}{[CO_2]}$ D. $[H_2O] = \dfrac{K_{eq}[HF]^2[CaCO_3]}{[CO_2][CaF_2]}$

98. Which of the following reactions will proceed furthest toward completion?

A. $Si_{(s)} + O_{2(g)} \rightleftarrows SiO_{2(s)}$ $K_{eq} = 2.0 \times 10^{142}$

B. $2HBr_{(g)} \rightleftarrows H_{2(g)} + Br_{2(g)}$ $K_{eq} = 7.0 \times 10^{-20}$

C. $2H_2O_{(g)} \rightleftarrows 2H_{2(g)} + O_{2(g)}$ $K_{eq} = 7.3 \times 10^{-18}$

D. $C_2H_{4(g)} + H_{2(g)} \rightleftarrows C_2H_{6(g)}$ $K_{eq} = 9.0 \times 10^{19}$

99.

$$CH_{4(g)} + H_2O_{(g)} \rightleftarrows CO_{(g)} + 3H_{2(g)} \qquad K_{eq} = 5.67$$

An equilibrium mixture of this system was found to contain the following concentrations: $[CH_4] = 0.59\,M$, $[H_2O] = 0.63\,M$, $[CO] = 0.25\,M$.
What was the equilibrium $[H_2]$?

A. 0.26 M
B. 0.64 M
C. 2.0 M
D. 8.4 M

100. $H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$ $K_{eq} = 49.5$ at $440°C$

If $5.0\,M$ HI is initially placed into a container, what will be the equilibrium $[HI]$?

A. 0.33 M
B. 3.9 M
C. 4.4 M
D. 4.8 M

101. $H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$ $K_{eq} = 49.5$ at $440°C$

If $0.120\,M$ H_2, $0.120\,M$ I_2 and $0.844\,M$ HI are placed into a container at $440°C$, which of the following is true as equilibrium is approached?

A. $[I_2]$ decreases significantly.
B. $[HI]$ decreases significantly.
C. $[H_2]$ decreases significantly.
D. $[H_2]$ remains the same.

102. A small amount of HCl is added to the following equilibrium system:

$$2CrO_4{}^{2-}{}_{(aq)} + 2H^+{}_{(aq)} \rightleftarrows Cr_2O_7{}^{2-}{}_{(aq)} + H_2O_{(\ell)}$$

How do the $[CrO_4{}^{2-}]$ and the reverse reaction rate change as equilibrium is re-established?

	$[CrO_4{}^{2-}]$	Reverse Rate
A.	increases	increases
B.	increases	decreases
C.	decreases	decreases
D.	decreases	increases

103. $N_2O_{4(g)} \rightleftarrows 2NO_{2(g)}$

colourless brown

If N_2O_4 is placed in a flask at a constant temperature, which of the following is true as the system approaches equilibrium?

A. The colour gets darker as $[NO_2]$ increases.
B. The colour gets lighter as $[NO_2]$ decreases.
C. The colour gets darker as $[N_2O_4]$ increases.
D. The colour gets lighter as $[N_2O_4]$ decreases.

104. $N_2O_{4(g)} \rightleftarrows 2NO_{2(g)}$

colourless brown

The system above reaches equilibrium. Considering enthalpy and entropy factors, which of the following is true with respect to the forward reaction?

A. The entropy is increasing and the reaction is exothermic.
B. The entropy is decreasing and the reaction is exothermic.
C. The entropy is increasing and the reaction is endothermic.
D. The entropy is decreasing and the reaction is endothermic.

105. In which of the following reactions do the tendencies for minimum enthalpy and maximum entropy both favour reactants?

A. $3O_{2(g)} \overset{?}{\rightleftarrows} 2O_{3(g)}$ $\Delta H = +285\,kJ$

B. $N_{2(g)} + 3H_{2(g)} \overset{?}{\rightleftarrows} 2NH_{3(g)}$ $\Delta H = -92\,kJ$

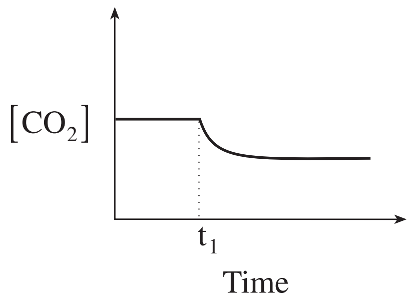
C. $2BrCl_{(g)} \overset{?}{\rightleftarrows} Br_{2(g)} + Cl_{2(g)}$ $\Delta H = -29.3\,kJ$

D. $CaCO_{3(s)} \overset{?}{\rightleftarrows} CaO_{(s)} + CO_{2(g)}$ $\Delta H = +175\,kJ$

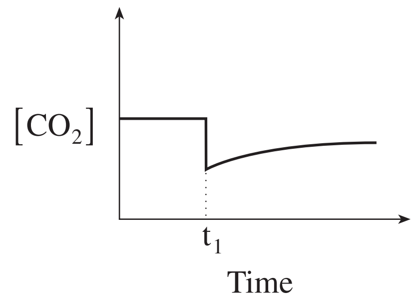
106. $$CaCO_{3(s)} \rightleftarrows CaO_{(s)} + CO_{2(g)} \qquad \Delta H = +175\,\text{kJ}$$

Which of the following diagrams best represents the change in the concentration of CO_2 as temperature is decreased at time t_1 ?

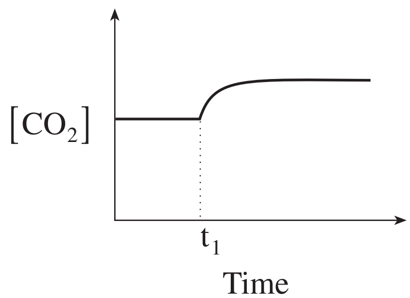
A.



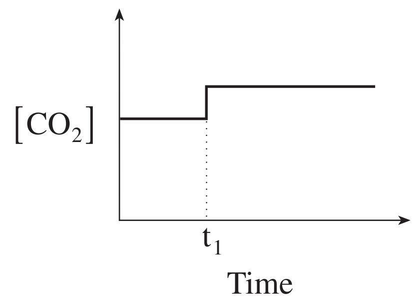
B.



C.



D.



107. The Haber Process is used to produce ammonia commercially according to the following equilibrium:

$$N_{2(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)} + \text{energy}$$

Which of the following conditions will produce the highest yield of ammonia?

A. increase temperature and increase pressure
B. increase temperature and decrease pressure
C. decrease temperature and increase pressure
D. decrease temperature and decrease pressure

108. $$Cr_2O_7{}^{2-}{}_{(aq)} + 2\,OH^-{}_{(aq)} \rightleftarrows 2CrO_4{}^{2-}{}_{(aq)} + H_2O_{(\ell)} \qquad K_{eq} = 4.14$$

The concentration of ions at equilibrium was measured at a specific temperature and found to be $\left[Cr_2O_7{}^{2-}\right] = 0.100\,\text{M}$ and $\left[OH^-\right] = 0.020\,\text{M}$. What is the equilibrium $\left[CrO_4{}^{2-}\right]$?

A. 1.7×10^{-4} M
B. 3.1×10^{-3} M
C. 1.3×10^{-2} M
D. 2.0×10^{-1} M

109. In which of the following equilibria does the concentration of reactants equal the concentration of products?

A. $N_2O_{4(g)} \rightleftarrows 2NO_{2(g)} \quad K_{eq} = 0.71$

B. $H^+_{(aq)} + OH^-_{(aq)} \rightleftarrows H_2O_{(\ell)} \quad K_{eq} = 1.0 \times 10^{14}$

C. $CO_{2(g)} + H_{2(g)} \rightleftarrows CO_{(g)} + H_2O_{(g)} \quad K_{eq} = 0.279$

D. $SnO_{2(s)} + 2H_{2(g)} \rightleftarrows Sn_{(s)} + 2H_2O_{(g)} \quad K_{eq} = 1.00$

110.

$$C_{(s)} + H_2O_{(g)} \rightleftarrows CO_{(g)} + H_{2(g)}$$

At equilibrium, 4.0×10^{-2} mol H_2, 4.0×10^{-2} mol CO, 1.0×10^{-2} mol H_2O and 1.0×10^{-2} mol C were present in a 1.0 L container. What is the value of K_{eq} ?

A. 0.063
B. 0.16
C. 6.3
D. 16

111.

$$2COF_{2(g)} \rightleftarrows CO_{2(g)} + CF_{4(g)}$$

Initially, 0.12 M CO_2 and 0.20 M CF_4 are placed in a container. At equilibrium, it is found that the $[COF_2]$ is 0.040 M. What is the value of K_{eq} ?

A. 0.089
B. 0.45
C. 8.0
D. 11

112. Reacting systems tend toward which of the following?

	Entropy	Enthalpy
A.	minimum	maximum
B.	minimum	minimum
C.	maximum	minimum
D.	maximum	maximum

113.

$$NH_4NO_{3(s)} \overset{?}{\rightleftarrows} N_2O_{(g)} + 2H_2O_{(g)} \qquad \Delta H = -37\,\text{kJ}$$

Which of the following is true?

	Enthalpy	Entropy	Outcome
A.	favours reactants	favours reactants	reaction does not occur
B.	favours products	favours products	reaction goes to completion
C.	favours reactants	favours products	reaction reaches equilibrium
D.	favours products	favours reactants	reaction reaches equilibrium

114.

$$\text{energy} + 2KClO_{3(s)} \rightleftarrows 2KCl_{(s)} + 3O_{2(g)}$$

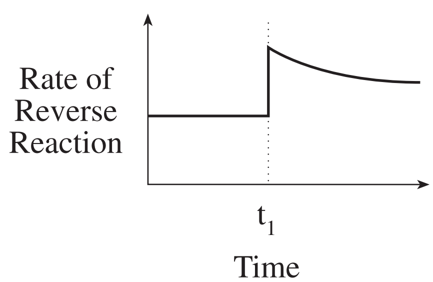
Which of the following will cause a shift to the left?

A. adding more O_2
B. adding more KCl
C. removing some $KClO_3$
D. increasing the temperature

115.

$$N_{2(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)} + \text{energy}$$

The following diagram represents the rate of the reverse reaction.



Which of the following stresses explains what happened at t_1 ?

A. $[H_2]$ increased.
B. $[N_2]$ decreased.
C. $[NH_3]$ increased.
D. $[NH_3]$ decreased.

116. Limestone is decomposed to make quicklime (CaO) according to the following equilibrium:

$$CaCO_{3(s)} + 175\text{kJ} \rightleftarrows CaO_{(s)} + CO_{2(g)}$$

Which of the following conditions would produce the greatest yield of $CaO_{(s)}$?

	Temperature	Pressure
A.	low	low
B.	low	high
C.	high	low
D.	high	high

117. $$H_{2(g)} + \tfrac{1}{2}O_{2(g)} \rightleftarrows H_2O_{(\ell)}$$

Which of the following represents the concentration of O_2 at equilibrium?

A. $[O_2] = \left(\dfrac{1}{K_{eq}[H_2]}\right)^2$

B. $[O_2] = K_{eq}[H_2]$

C. $[O_2] = \left(\dfrac{[H_2O]}{K_{eq}[H_2]}\right)^2$

D. $[O_2] = \sqrt{\dfrac{1}{K_{eq}[H_2]}}$

118. $$2CrO_4{}^{2-}{}_{(aq)} + 2H^+{}_{(aq)} \rightleftarrows Cr_2O_7{}^{2-}{}_{(aq)} + H_2O_{(\ell)}$$

A solution of $Ba(NO_3)_2$ is added, and a precipitate of $BaCrO_4$ forms. In which direction will the equilibrium shift, and what will happen to the value of K_{eq}?

A. Equilibrium shifts left, and K_{eq} decreases.
B. Equilibrium shifts right, and K_{eq} increases.
C. Equilibrium shifts left, and K_{eq} remains constant.
D. Equilibrium does not shift, and K_{eq} remains constant.

119.

$$CH_{4(g)} + H_2O_{(\ell)} \rightleftarrows CO_{(g)} + 3H_{2(g)}$$

At equilibrium, 1.2 mol CH_4, 1.2 mol H_2O, 0.080 mol CO and 0.040 mol H_2 are present in a 1.0 L container. What is the value of K_{eq} ?

A. 4.3×10^{-6}
B. 3.6×10^{-6}
C. 2.7×10^{-3}
D. 2.3×10^{5}

120.

$$2NO_{(g)} + 2H_{2(g)} \rightleftarrows N_{2(g)} + 2H_2O_{(g)}$$

Initially, 0.100 mol NO, 0.0500 mol H_2 and 0.100 mol H_2O are placed in a 1.0 L container. At equilibrium, the $[H_2O] = 0.138\,M$. What is the value of K_{eq} ?

A. 3.5
B. 6.5×10^{2}
C. 1.5×10^{-3}
D. 1.3×10^{3}

Written-Solution Questions:

121.

$$2COF_{2(g)} \rightleftarrows CO_{2(g)} + CF_{4(g)} \qquad K_{eq} = 2.00$$

A 2.00 L container is filled with 0.500 mol of COF_2.
Calculate the $[COF_2]$ at equilibrium.

122.

$$\underset{\textit{blue}}{Cu^{2+}_{(aq)}} + \underset{\textit{colourless}}{4Br^{-}_{(aq)}} \rightleftarrows \underset{\textit{green}}{CuBr_{4}{}^{2-}_{(aq)}}$$

Cooling the equilibrium changes the colour from green to blue. What effect will the decrease in temperature have on K_{eq} ? Explain, using Le Chatelier's Principle.

123. A flask is initially filled with some HI. At equilibrium, the $[HI] = 0.80\,mol/L$
What is the $[H_2]$ at equilibrium?

$$2HI_{(g)} \rightleftarrows H_{2(g)} + I_{2(g)} \qquad K_{eq} = 0.25$$

124. $$2NOCl_{(g)} \rightleftarrows 2NO_{(g)} + Cl_{2(g)} \qquad K_{eq} = 1.6 \times 10^{-5}$$

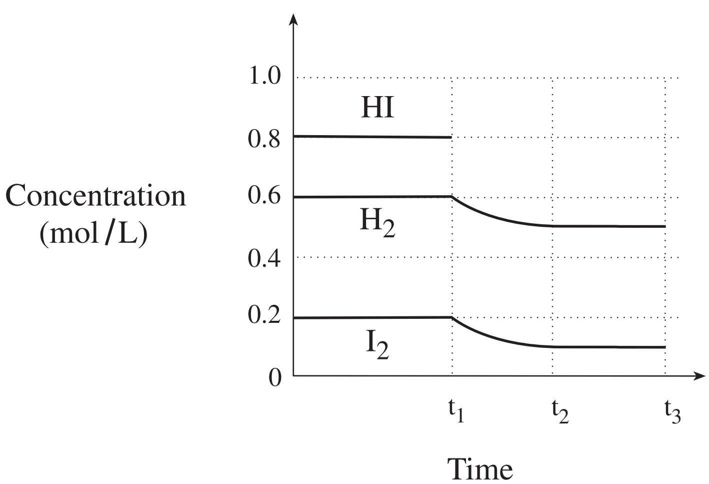
A $1.00\,\text{L}$ flask is filled with $0.20\,\text{mol}$ NOCl, $0.10\,\text{mol}$ NO and $0.10\,\text{mol}$ Cl_2. State and show by calculation the direction in which the reaction proceeds to reach equilibrium.

125. $$H_{2(g)} + Br_{2(g)} \rightleftarrows 2HBr_{(g)} \qquad K_{eq} = 14.8$$

A closed container was initially filled with equal moles of H_2 and Br_2.
When equilibrium is reached, the $[HBr]$ is $0.329\,\text{mol/L}$. What was the initial $[H_2]$?

126. Write **four** statements that apply to all chemical equilibrium systems.

127. $$H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$$



The temperature is increased at t_1 and equilibrium is re-established at t_2.

a) On the above graph, sketch the line representing the $[HI]$ between time t_1 and t_3.

b) Calculate the value of K_{eq} after t_2.

128. $$2SO_{2(g)} + O_{2(g)} \rightleftarrows 2SO_{3(g)}$$

A $1.00\,\text{L}$ container is initially filled with $0.100\,\text{mol}$ SO_2 and $0.100\,\text{mol}$ O_2.
At equilibrium the O_2 concentration is $0.060\,\text{mol/L}$. Calculate the value of K_{eq}.

129.

$$2NF_{2(g)} \rightleftarrows N_2F_{4(g)}$$

Equilibrium shifts to the right when volume is decreased. Describe the changes in reaction rates that cause this shift to the right.

130.

$$H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$$

Initially, 0.200 mol H_2 and 0.200 mol I_2 are added to an empty 2.00 L container. At equilibrium, the $[I_2] = 0.020$ mol/L. What is the value of K_{eq} ?

131. State *Le Chatelier's Principle*.

132.

$$CH_{4(g)} + H_2O_{(g)} \rightleftarrows CO_{(g)} + 3H_{2(g)}$$

Initially, 0.060 mol CH_4 , 0.080 mol H_2O , 0.280 mol CO and 0.740 mol H_2 are placed into a 4.00 L container. At equilibrium, the $[H_2] = 0.200$ mol/L. What is the value of K_{eq} ?

133. Chemical reactions tend toward a position of minimum enthalpy and maximum entropy.

a) What is meant by the term *enthalpy*?

b) What is meant by the term *entropy*?

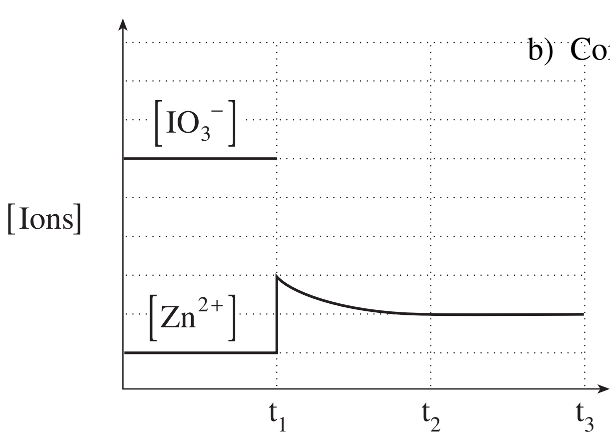
134.

$$H_{2(g)} + Br_{2(g)} \rightleftarrows 2HBr_{(g)} \qquad K_{eq} = 12.0$$

Initially, 0.080 mol H_2 and 0.080 mol Br_2 are placed into a 4.00 L container. What is the $[HBr]$ at equilibrium?

135.

$$Zn(IO_3)_{2(s)} \rightleftarrows Zn^{2+}_{(aq)} + 2IO_3^-{}_{(aq)}$$



a) Identify the stress applied at t_1 .

b) Complete the above graph from t_1 to t_3 for the $[IO_3^-]$.

136. Describe the nature of *dynamic equilibrium.*

137.

$$N_2O_{4(g)} \rightleftarrows 2NO_{2(g)} \qquad K_{eq} = 9.5 \times 10^{-3}$$

Initially, 0.060 mol N_2O_4 and 0.020 mol NO_2 are placed into a 2.00 L container. Use calculations to determine the direction in which the reaction proceeds in order to reach equilibrium.

138.

$$CH_{4(g)} + H_2O_{(g)} \rightleftarrows CO_{(g)} + 3H_{2(g)}$$

K_{eq}	Temperature
1.78×10^{-3}	800°C
4.68×10^{-2}	1000°C

Is the forward reaction in this equilibrium exothermic or endothermic? Explain your answer.

139.

$$CO_{(g)} + Cl_{2(g)} \rightleftarrows COCl_{2(g)}$$

At equilibrium, the system contains 2.00 mol CO, 1.00 mol Cl_2 and 0.200 mol $COCl_2$ in a 2.0 L container. Calculate the value of K_{eq}.

140.

$$N_{2(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)} \qquad \Delta H = -92\text{ kJ}$$

The system is normally maintained at a temperature of approximately 500°C.

a) Explain why 1000°C is not used.

b) Explain why 100°C is not used.

141.

$$SO_{3(g)} + NO_{(g)} \rightleftarrows NO_{2(g)} + SO_{2(g)}$$

In an experiment, 0.100 moles of SO_3 and 0.100 moles of NO are placed in a 1.00 L container. When equilibrium is achieved, $[NO_2] = 0.0414\text{ mol/L}$. Calculate the K_{eq} value.

142.

$$C_3H_{8(g)} + 5O_{2(g)} \overset{?}{\rightarrow} 3CO_{2(g)} + 4H_2O_{(g)}$$

Explain, in terms of increasing or decreasing entropy and enthalpy, whether or not the reaction will reach equilibrium.

143. $$H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)} \qquad K_{eq} = 64$$

Equal moles of H_2, I_2 and HI are placed in a 1.0 L container. Use calculations to determine the direction the reaction will proceed in order to reach equilibrium.

144. $$H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$$

The system is said to "shift right" as the result of the addition of **extra** $H_{2(g)}$. Describe the sequence of changes in both forward and reverse reaction rates as the system goes from the original equilibrium to the new equilibrium.

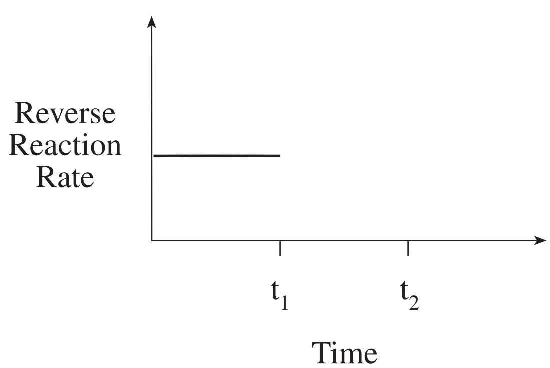
145. $$2NO_{(g)} + Cl_{2(g)} \rightleftarrows 2NOCl_{(g)} \qquad K_{eq} = 8.5$$

A closed flask is found to contain 0.40 M $NO_{(g)}$, 0.32 M $Cl_{2(g)}$ and 5.6 M $NOCl_{(g)}$. Use appropriate calculations to determine the direction the reaction proceeds to reach equilibrium.

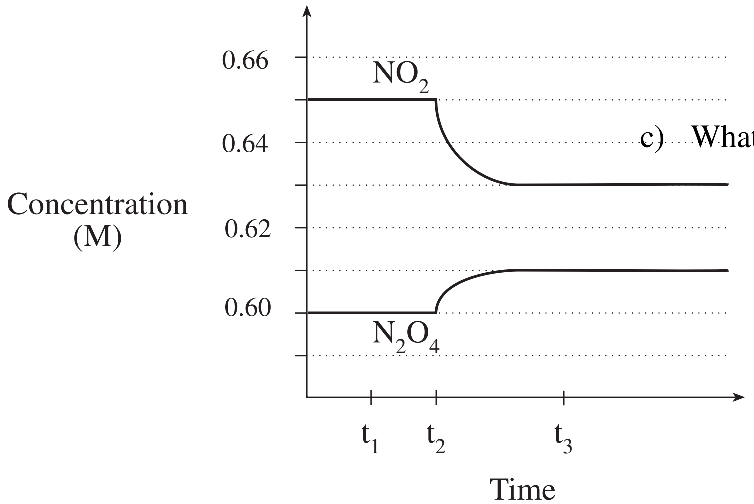
146. $$Al(H_2O)_4(OH)_2{}^{+}{}_{(aq)} \rightleftarrows Al(H_2O)_3(OH)_{3(s)} + H^{+}{}_{(aq)}$$

a) Some $HCl_{(aq)}$ is added to the equilibrium. What happens to the amount of solid $Al(H_2O)_3(OH)_3$? Explain.

b) The HCl is added at time t_1 and equilibrium is re-established at time t_2. On the axis below, sketch what happens to the reverse reaction rate.



147. $N_2O_{4(g)} \rightleftarrows 2NO_{2(g)}$ $\Delta H = +59\,kJ$



a) Calculate the value of K_{eq} at t_1.

b) Calculate the value of K_{eq} at t_3.

c) What stress was applied at time t_2 ? Explain.

148.

$$N_{2(g)} + O_{2(g)} \rightleftarrows 2NO_{(g)} \qquad K_{eq} = 0.0095$$

Initially, $0.15\,mol\ N_2$ and $0.15\,mol\ O_2$ were placed in a $1.0\,L$ container. Calculate the concentration of all species at equilibrium.

149.

$$N_{2(g)} + O_{2(g)} \rightleftarrows 2NO_{(g)} \qquad K_{eq} = 8.1 \times 10^{-3}$$

A $2.0\,L$ container is filled with $0.15\,mol\ N_2$, $0.15\,mol\ O_2$ and $0.050\,mol\ NO$. Does the $[NO]$ increase or decrease as equilibrium is established? Support your answer with appropriate calculations.

Multiple-Choice Answers

1	A	6	C	11	D	16	A	21	C	26	B	31	B	36	C	41	A	46	D
2	D	7	A	12	B	17	A	22	C	27	B	32	C	37	B	42	B	47	B
3	C	8	C	13	A	18	D	23	C	28	A	33	D	38	C	43	D	48	D
4	C	9	A	14	D	19	C	24	B	29	D	34	A	39	D	44	A	49	D
5	D	10	B	15	A	20	A	25	A	30	D	35	C	40	B	45	A	50	B

51	C	56	C	61	D	66	C	71	D	76	D	81	A	86	C	91	D	96	B
52	A	57	A	62	B	67	A	72	C	77	A	82	D	87	A	92	C	97	B
53	C	58	A	63	C	68	C	73	B	78	A	83	B	88	A	93	C	98	A
54	A	59	B	64	A	69	C	74	A	79	A	84	A	89	A	94	B	99	C
55	C	60	B	65	A	70	B	75	D	80	C	85	C	90	D	95	A	100	B

101	D	106	A	111	D	116	C												
102	D	107	C	112	C	117	A												
103	A	108	C	113	B	118	C												
104	C	109	D	114	A	119	A												
105	A	110	B	115	C	120	B												

Written-Solution Answers:

121.

	$2COF_{2(g)}$	$\rightleftarrows$	$CO_{2(g)}$	+	$CF_{4(g)}$
[I]	$\frac{0.500\text{ mol}}{2.00\text{ L}} = 0.250\text{ mol/L}$		0		0
[C]	$-2x$		$+x$		$+x$
[E]	$0.250 - 2x$		x		x

$$K_{eq} = \frac{[CO_2][CF_4]}{[COF_2]^2}$$

$$2.00 = \frac{(x)(x)}{(0.250 - 2x)^2}$$

$$x = 0.0923\text{ mol/L}$$

$$\begin{aligned}[COF_2] &= 0.250 - 2x \\ &= 0.250 - 2(0.0923) \\ &= 0.065\text{ mol/L}\end{aligned}$$

122. The value of K_{eq} decreased.

When the temperature stress is applied to the equilibrium system, the system shifts to the left to offset the stress and in the new equilibrium, the [products] has decreased.

123.

$$K_{eq} = \frac{[H_2][I_2]}{[HI]^2}$$

$$0.25 = \frac{(x)(x)}{(0.80)^2}$$

$$x = 0.40$$

$$[H_2] = 0.40\text{ M}$$

124. Direction: Reaction proceeds to the left.

Calculations:

$$2NOCl_{(g)} \rightleftarrows 2NO_{(g)} + Cl_{2(g)}$$

$$K_{Trial} = \frac{[NO]^2[Cl_2]}{[NOCl]^2}$$

$$= \frac{(0.10)^2(0.10)}{(0.20)^2}$$

$$= 0.025$$

$$K_{Trial} > K_{eq}$$

125.

	$H_{2(g)}$	+	$Br_{2(g)}$	$\rightleftarrows$	$2HBr_{(g)}$
[I]	x		x		0
[C]	-0.1645		-0.1645		$+0.329$
[E]	$x-0.1645$		$x-0.1645$		0.329

$$K_{eq} = \frac{[HBr]^2}{[H_2][Br_2]}$$

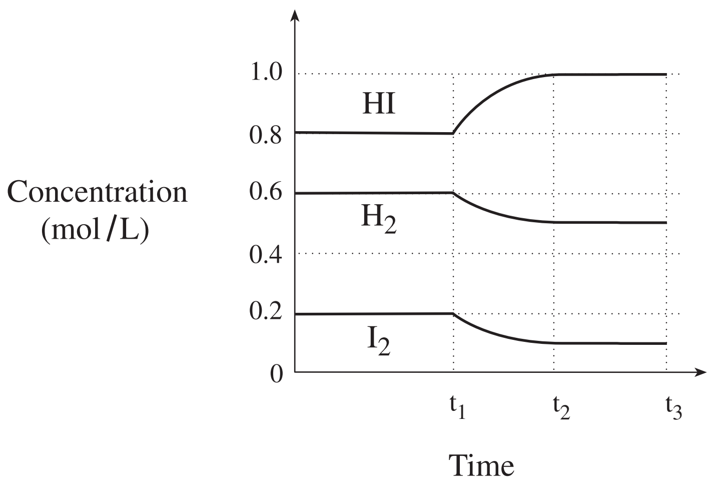
$$14.8 = \frac{(0.329)^2}{(x-0.1645)(x-0.1645)}$$

$$x = 0.250\,\text{M}$$

$$[H_2] = 0.250\,\text{mol/L}$$

126.
- system must be closed
- temperature is constant
- forward and reverse rates are equal
- macroscopic properties are constant
- can be achieved from either direction
- concentration of reactants and products are constant

127. a)



b)

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

$$= \frac{(1.0)^2}{(0.5)(0.1)}$$

$$= 20$$

128.

	$2SO_{2(g)}$	+	$O_{2(g)}$	$\rightleftarrows$	$2SO_{3(g)}$
[I]	0.100		0.100		0
[C]	−0.080		−0.040		+0.080
[E]	0.020		0.060		0.080

$$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

$$= \frac{(0.080)^2}{(0.020)^2(0.060)}$$

$$= 2.7 \times 10^2$$

129. Both forward and reverse rates increase as a result of increased concentration.

The forward rate increases more than the reverse rate, so the equilibrium shifts to the right.

130.

	H_2	+	I_2	$\rightleftarrows$	$2HI$
[I]	0.100 mol/L		0.100		0
[C]	−0.080		−0.080		+0.160
[E]	0.020		0.020		0.160

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]} = \frac{(0.160)^2}{(0.020)(0.020)} = 64$$

131. Le Chatelier's Principle states that when a stress is placed on an equilibrium system, the system will shift to offset this stress until a new equilibrium is reached.

132.

	CH_4	+	H_2O	$\rightleftarrows$	CO	+	$3H_2$
[I]	0.015 mol/L		0.020		0.070		0.185
[C]	−0.005		−0.005		+0.005		+0.015
[E]	0.010		0.015		0.075		0.200

$$K_{eq} = \frac{[CO][H_2]^3}{[CH_4][H_2O]}$$

$$= \frac{(0.075)(0.200)^3}{(0.010)(0.015)}$$

$$= 4.0$$

133. Enthalpy is a measure of heat content. Entropy is a measure of randomness.

134.

	H_2	+	Br_2	$\rightleftarrows$	2HBr
[I]	0.020		0.020		0
[C]	$-x$		$-x$		$+2x$
[E]	$0.020-x$		$0.020-x$		$2x$

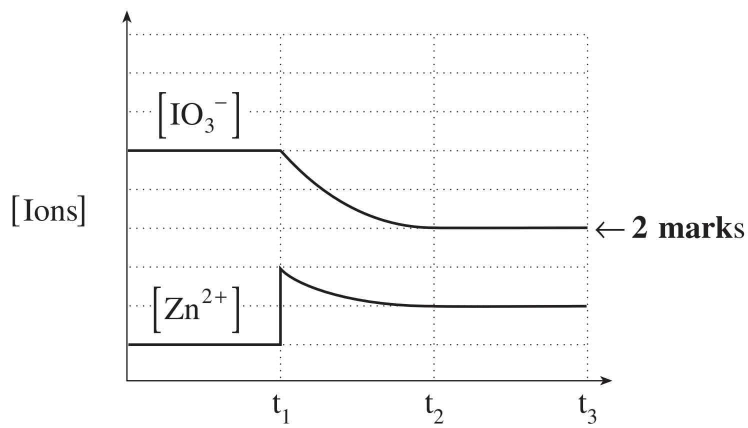
$$K_{eq} = \frac{[HBr]^2}{[H_2][Br_2]} = \frac{(2x)^2}{(0.020-x)^2} = 12.0$$

$$\sqrt{\frac{(2x)^2}{(0.020-x)}} = \sqrt{12.0}$$

$$x = 0.0127$$

$$[HBr] = 2x = 0.025\,M$$

135. a) More $Zn^{2+}_{(aq)}$ has been added.



136. In a dynamic equilibrium, the forward reaction and reverse reaction continue to proceed at equal rates.

137.

$$[N_2O_4] = \frac{0.060\,mol}{2.00\,L} = 0.030\,mol/L$$

$$[NO_2] = \frac{0.020\,mol}{2.00\,L} = 0.010\,mol/L$$

$$K_{trial} = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(0.010)^2}{(0.030)} = 3.3 \times 10^{-3}$$

$$K_{trial} < K_{eq}$$

$\therefore$ The reaction proceeds to the right.

138. This equilibrium is endothermic.

Since K_{eq} increases as a result of a temperature increase, equilibrium has shifted to the right.

139.

$$K_{eq} = \frac{[COCl_2]}{[CO][Cl_2]}$$

$$= \frac{(0.200\,mol/2.0\,L)}{(2.00\,mol/2.0\,L)(1.00\,mol/2.0\,L)}$$

$$= \frac{(0.100)}{(1.00)(0.500)}$$

$$= 0.20$$

140. Equilibrium will be shifted to the left, reducing the yield of NH_3.
The rate of the reaction would be too low.

141. Entropy increases in the forward reaction.

Enthalpy decreases in the forward reaction.

Since both favour products, equilibrium will not be attained; or the reaction will go to completion.

142.

	SO_3	+	NO	$\rightleftarrows$	NO_2	+	SO_2
[I]	0.100		0.100		0		0
[C]	−0.0414		−0.0414		+0.0414		+0.0414
[E]	0.059		0.059		0.0414		0.0414

$$K_{eq} = \frac{[NO_2][SO_2]}{[SO_3][NO]} = \frac{(0.0414)(0.0414)}{(0.059)(0.059)} = 0.50$$

143.

	H_2	+	I_2	$\rightleftarrows$	2HI
[I]	x		x		x

$$\text{Trial } K_{eq} = \frac{[HI]^2}{[H_2][I_2]} = \frac{(x)^2}{(x)(x)} = 1$$

Since Trial $K_{eq} < K_{eq}$, equilibrium is established by proceeding to the right.

144. As $[H_2]$ is increased the forward rate increases. The forward rate will be greater than the reverse rate, resulting in more HI being produced.

The $[H_2]$ is consumed as the shift occurs and the forward rate starts to decrease. The increasing [HI] results in an increasing reverse rate.

At the new equilibrium the forward and reverse rates will be equal.

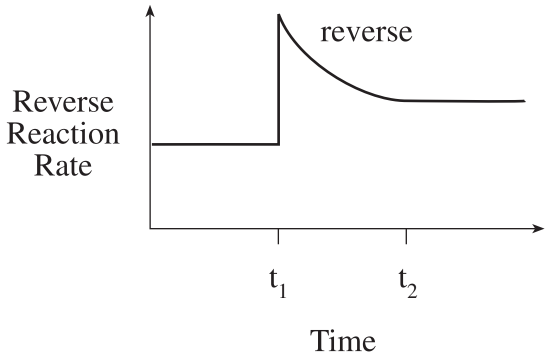
145. $$\text{Trial } K_{eq} = \frac{[NOCl]^2}{[NO]^2[Cl_2]} = \frac{(5.6)^2}{(0.40)^2(0.32)} = 6.1 \times 10^2$$

Since Trial $K_{eq} > K_{eq}$,

the equilibrium will proceed left to reduce the Trial K_{eq} value to K_{eq}.

146. a) The amount of solid decreases because the equilibrium shifts left.

b)



147. a) $$K_{eq} = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(0.65)^2}{0.60} = 0.70$$

b) $$K_{eq} = \frac{(0.63)^2}{0.61} = 0.65$$

c) Stress: Temperature was decreased

Explanation: because K_{eq} decreased.

The appearance of the graph is consistent with a temperature shift.

148.

	N_2	+	O_2	$\rightleftarrows$	$2NO$
[I]	0.15		0.15		0
[C]	$-x$		$-x$		$+2x$
[E]	$0.15-x$		$0.15-x$		$2x$

$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$$

$$0.0095 = \frac{(2x)^2}{(0.15-x)^2}$$

$$\sqrt{0.0095} = \sqrt{\frac{(2x)^2}{(0.15-x)^2}}$$

$$x = 6.97 \times 10^{-3}$$

$$[N_2] = [O_2] = 0.15 - x = 0.14\,\mathrm{M}$$

$$[NO] = 2(x) = 0.014\,\mathrm{M}$$

149.

$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]} = 8.1 \times 10^{-3}$$

$$\text{Trial } K_{eq} = \frac{(0.025)^2}{(0.075)(0.075)} = 0.11$$

Trial $K_{eq} > K_{eq}$ so reaction proceeds to the left

[NO] decreases