These are homework exercises to accompany the Textmap created for "General Chemistry: Principles and Modern Applications" by Petrucci et al. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed here. In addition to these publicly available questions, access to private problems bank for use in exams and homework is available to faculty only on an individual basis; please contact Delmar Larsen for an account with access permission.

Q14.1

Looking at the following, write the balanced equation and Kc expression for each reversible reaction:

a. the dissociation of water:

b. the reaction between gaseous \(\text{CO}\) and \(\text{NO}_2\):

c. the addition of Hydrogen Iodide to water:

Q14.2

Balance the equilibrium equation and determine the corresponding \((K_p)\) expression for each reversible reaction.

a. \(\text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}(g)\)

b. \(\text{Ag}_2\text{O(s)} \rightleftharpoons 2\text{Ag(s)} + \frac{1}{2}\text{O}_2(g)\)

c. \(\text{Cs}_2(g) + 4\text{H}_2(g) \rightleftharpoons \text{CH}_4(g) + 2\text{H}_2\text{S(g)}\)

S14.2

a.) \(\text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO(g)}\)

\[K_p = \frac{(P_{\text{NO}})^2}{(P_{\text{N}_2\text{O}_4})}\]

b.) \(\text{Ag}_2\text{O(s)} \rightleftharpoons 2\text{Ag(s)} + \frac{1}{2}\text{O}_2(g)\)

\[K_p = (P_{\text{Ag}})^2 \times (P_{\text{O}_2})^{1/2}/(P_{\text{Ag}_2\text{O}})\]

c.) \(\text{Cs}_2(g) + 4\text{H}_2(g) \rightleftharpoons \text{CH}_4(g) + 2\text{H}_2\text{S(g)}\)

\[K_p = (P_{\text{CH}_4})^1 \times (P_{\text{H}_2\text{S}})^2/(P_{\text{Cs}_2})^1 \times (P_{\text{H}_2})^4\]

http://chemwiki.ucdavis.edu/Physical_Chemistry/Chemical_Equilibrium/The_Equilibrium_Constant/
Calculating_An_Equilibrium_Concentration_From_An_Equilibrium_Constant/
Writing_Equilibrium_Constant_Expressions_Involving_Gases/Gas_Equilibrium_Constants%3a_Kc_And_Kp
Q14.2
Based on the following, write a balanced equation and kp expression for each of the following:

a. \(N_2(g) + H_2(g) \rightleftharpoons NH_3(g)\)
b. The thermal decomposition of \(CaCO_3\)
c. The reaction of \(CO_2\) with a sodium hydroxide solution

Q14.3
Give the \(K_c\) expressions for the following reactions:

a. \(2HI(g) + O_2(g) \rightleftharpoons 2HIO_2(g)\)
b. \(ClNO_2(g) + NO(g) \rightleftharpoons NO_2(g) + ClNO(g)\)
c. \(H_2(g) + I_2(g) \rightleftharpoons 2HI(g)\)

Q14.33
A mixture consisting of 0.201 mol \(H_2\), and 0.201 mol of \(I_2\) is brought to equilibrium at 500°C, in a 4.5 L flask. What are the equilibrium amounts of \(H_2\), \(I_2\), and \(HI\)? \(K_c = 50.2\) at 500°C

S14.33
Video Solution

0.201 mol H2 /4.5 L = 0.0447 M H2 (First calculate molarity)

0.201 mol I2/4.5 L = 0.0447 M I2

\[H_2 + I_2 \rightleftharpoons 2HI\] (Balance equation and set up ICE table)

<table>
<thead>
<tr>
<th>ICE TABLE</th>
<th>(H_2)</th>
<th>(I_2)</th>
<th>(HI)</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>0.0447M</td>
<td>0.0447M</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>-X</td>
<td>-X</td>
<td>+2X</td>
</tr>
<tr>
<td>E</td>
<td>0.0447-X</td>
<td>0.0447-X</td>
<td>2X</td>
</tr>
</tbody>
</table>

\[K_c = \dfrac{2X}{(0.0447-X)^2}\]

\[50.2 = \dfrac{2X}{(0.0447-X)^2}\]

(use quad formula)
X = 0.0264

\([\text{H}_2]\) = 0.0183 \text{ M}

\([\text{I}_2]\) = 0.0183 \text{ M}

\([\text{HI}]\) = 0.0528 \text{ M}

http://chemwiki.ucdavis.edu/Physical_Chemistry/Acids_and_Bases/Ionization_Constants/Calculating_Equilibrium_Concentrations

Q14.43

Cadmium metal is added to 0.500 L of 1 M an aqueous solution of \(\text{[Cr}^{3+}\text{]}\). What are the concentrations of the different ionic species at equilibrium for the oxidation of \(\text{[Cd]}\) into \(\text{[Cd}^{2+}\text{]}\): \n
\[2\text{Cr}^{3+}_{\text{(aq)}} + \text{Cd}_{\text{(s)}} \rightleftharpoons 2\text{Cr}^{2+}_{\text{(aq)}} + \text{Cd}^{2+}_{\text{(aq)}}\]

with \(K_c = 0.288\).

**Solution**

The ICE table is first constructed

<table>
<thead>
<tr>
<th>ICE</th>
<th>(\text{[Cr}^{3+}\text{]}) (M)</th>
<th>(\text{[Cr}^{2+}\text{]}) (M)</th>
<th>(\text{[Cd}^{2+}\text{]}) (M)</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>1</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>-2x</td>
<td>+2x</td>
<td>+x</td>
</tr>
<tr>
<td>E</td>
<td>1-2x</td>
<td>2x</td>
<td>x</td>
</tr>
</tbody>
</table>

Now construct the relationship for the equilibrium constant (do not include solid Cd in equilibrium):

\[K_c = \frac{[\text{Cr}^{2+}_{\text{(aq)}}]^2[\text{Cd}^{2+}_{\text{(aq)}}]}{[\text{Cr}^{3+}_{\text{(aq)}}]^2} = \frac{(2x)^2(x)}{(1-2x)^2} \tag{1}\]

\[= \frac{4x^3}{(1-2x)^2} \tag{2}\]

Alternatively
\[K_c (1-2x)^2 = 4x^3 \tag{3}\]

\[4x^3 - 4K_c x + 4K_c x^2 = 0 \tag{4}\]

with \(K_c = 0.288\).
This is unfortunately is a cubic equation and can be solved analytically, but not easily. It is typically solved either graphically or numerically. Before solving numerical, we can evaluate one alternative by assuming that $x$ is small ($x/2 << 1$), then the equation 2 can be approximated and the cubic equation becomes solvable:

$$\sqrt[3]{\frac{4x^3}{(1-2x)^2}} \approx 4x^3$$

$$x \approx \sqrt[3]{\frac{K_c}{4}}$$

Unfortunately, this value for $x$ is clearly not small and hence the approximate is incorrect. Hence, the solution requires a numerical approach to identify the roots of equation 4. By plotting $(4x^3 - 4Kc^2x + 4Kc - Kc^3)$ vs. $x$, the physical solution (i.e., root between $x=0$ and $x=1$) can be resolved graphically (below).

From this graph, is clear that $x \approx 0.2571$, which is decently far off from the crude approximation in equation 5.

http://chemwiki.ucdavis.edu/Physical_Chemistry/Acids_and_Bases/Ionization_Constants/Calculating_Equilibrium_Concentrations

**Q14.71**

Use Le Chatelier’s Principal to explain what happens to the following reaction when you increase the temperature. Why does this happen?

$$\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \leftrightarrow 2 \text{NH}_3(\text{g}) \Delta H = -92\text{kJ/mol}$$

**S14.71**

Increasing the temperature will cause the equilibrium to shift towards the reactants (left). This is because the reaction is exothermic and releases heat. So if you add heat, more N$_2$(g) and H$_2$(g) will be created and balance out the reaction.

http://chemwiki.ucdavis.edu/Physical_Chemistry/Chemical_Equilibrium/Le_Chatelier's_Principle

**Q14.84**

For the synthesis of ammonia at 500 K,
\[\text{N}_2(g) + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3(g)\]

with \(K_p = 11.08 \times 10^{-2}\). Assume that \(\text{N}_2\) and \(\text{H}_2\) are mixed in the mole ratio 1:3 and that the total pressure is maintained at 1.00 atm. What is the mole percent \(\text{NH}_3\) at equilibrium? \textit{Hint: Use this equation} → \(K_p = \frac{(X_{\text{NH}_3})^2}{(X_{\text{N}_2})(X_{\text{H}_2})^2} \times \frac{1}{(P_{\text{tot}})^2}\)

\section*{S14.84}

\[11.08 \times 10^{-2} = \frac{(X_{\text{NH}_3})^2}{(1)(3)^2} \times (1/1)^2\]

\(X_{\text{NH}_3} = 0.998\ \text{mol}\)

\(0.998\ \text{mol NH}_3/6\ \text{tot mol} \times 100 = 16.63\%\ \text{mol NH}_3\)

http://chemwiki.ucdavis.edu/Physical_Chemistry/Chemical_Equilibrium/Calculating_an_Equilibrium_Concentration_from_the_Equilibrium_Constant/Calculating_an_Equilibrium_Constant%2c_Kp%2c_with_Partial_Pressures

\section*{Q14.29}

Consider a mixture of 2.7 mol O\(_2\), 3.2 mol SO\(_2\), and 2.1 mol SO\(_3\). Can this mixture be maintained in a 7.2 L container where \(K_c=20\)?

\section*{Q14.30}

In what direction will the following reaction proceed given that \(K_c=5.37E11\)?

\[2\text{COF}_2(g) \rightleftharpoons \text{CO}_2(g) + \text{CF}_4(g)\]

\section*{Q14.31}

In the reaction \(2\text{ClO}_2(g)+\text{O}_2(g)\rightleftharpoons 2\text{ClO}_3(g)\), 0.241 mol ClO\(_2\), 0.197 mol O\(_2\), 0.670 mol ClO\(_3\) are simultaneously placed into a 1.7 L vessel at a temperature of 1000K.

a. If \(K_c=2.53E2\) is the mixture at equilibrium?

b. If it is not in equilibrium then in what direction will the reaction shift?

\section*{Q14.33}

Consider a mixture of 0.263 mol H\(_2\) and .263 mol Br\(_2\)that is broughtto equilibrium at 500 degrees Celsius in a 4.0 L flask. What are concentrations at equilibrium of H\(_2\), Br\(_2\), and HBr?
Q14.35
For the reversible reaction between $\text{CO}$ and $\text{CO}_2$ in the gas phase

\[ \text{CO}(g) + \text{O}_2(g) \rightleftharpoons \text{CO}_2(g) \]

If 0.352 mol of $\text{CO}$ were added to 0.067 mol of $\text{CO}_2$ in a 3 L flask at 668 K, how many moles of $\text{O}_2(g)$ will be present at equilibrium? The $K_c$ of this reaction is $(1.2 \times 10^3)$ at 668 K.

Q14.37
Equilibrium is reached in a 2.5 L flask at 800 °C for the reaction of

\[ \text{CaCO}_3(s) \rightleftharpoons \text{CaCO}_3(s) + \text{CO}_2(g) \]

with $K_c=2.68 \times 10^{-3}$

a. How many moles of $\text{CaCO}_3(s)$ is present at equilibrium if 0.715 mols $\text{CaCO}_3(g)$ is introduced into an empty flask?

Q14.39
Consider the reaction, $\text{Ca(OH)}_2 + 2 \text{CH}_3\text{COOH} = \text{Ca(CH}_3\text{COO)}_2 + 2\text{H}_2\text{O}$ ; $K_c=4.0$

A reaction occurs in a mixture of 19.3 g $\text{Ca(OH)}_2$, 21.7 g $\text{CH}_3\text{COOH}$, 41.2 g $\text{Ca(CH}_3\text{COO)}_2$, and 70.1 g $\text{H}_2\text{O}$.

a. In what direction will the reaction proceed?

b. How many grams of each will be present at equilibrium?

Q14.41
Given, the thermal decomposition of $\text{CaCO}_3(g) \rightarrow \text{CaO}(g) + \text{CO}_2(g)$; $K_c=4.84 @300K$

if 0.362 mol $\text{CaCO}_3$ dissociates in a 2.47 L flask at 300 K what will be the TOTAL Pressure at equilibrium?

Q14.43
write the equilibrium expression for the reaction $[\text{citrate(aq)}]=0.027 \text{ M}$, $[\text{aconitate(aq)}]=4.31E-5$, and $[\text{H}_2\text{O}]= 53.6\text{M}$. Determine whether the reaction is at equilibrium / or which direction it will proceed if not at equilibrium.

\[ \text{Citrate(aq)↓⇌aconitate(aq)} + \text{H}_2\text{O(l)} \quad K=0.031 \]
Q14.45

Based on the equation \( \text{HNO}_3 \ (aq) \rightarrow \text{H}^+ \ (aq) + \text{NO}_3^- \ (aq) \); \( K = 3.21 \times 10^{11} \), if 2.0 M \( \text{HNO}_3 \) is added to a .5 L flask then what are the resulting ion concentrations at equilibrium?

Q14.53

Based on the following equilibrium reaction \( 2\text{H}(g) + \text{O}(g) \rightleftharpoons \text{H}_2\text{O}(g) \); \( K = 35 \)

What is the partial pressure of Oxygen?

Q14.55

Given the following situations, determine what will happen to the amount of \( \text{H}_2 \) that will be present in an equilibrium mixture in the reaction \( 3\text{Fe}(s) + 4\text{H}_2\text{O}(g) \rightleftharpoons \text{Fe}_3\text{O}_4(s) + 4\text{H}_2(g) \); \( \Delta H^\circ = -150 \text{ kJ} \)

a. Decreasing the temperature of the mixture  
b. Introducing more \( \text{Fe}(s) \)  
c. Halving the volume of the container  
d. Adding a catalyst

63

Consider the reaction \( 4\text{H}_2(g) + 2\text{O}_2(g) \rightleftharpoons 2\text{H}_2\text{O}(g) \); \( \Delta H^\circ = 271 \text{ kJ} \)

what affect does decreasing the temperature have on

a. The equilibrium production of \( \text{NO}(g) \)  
b. The rate at which the reaction takes place

Q14.65

Based on the reaction \( \text{N}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}(g) \) what will happen to the Pressure of \( \text{N}_2 \) if the volume of the mixture were to be increased by decreasing the pressure by half?

Q14.Q14.71

A container of \( \text{NO}(g) \) at 0.96 atm is brought into contact with 100 mL of a saturated aqueous solution with a \( \text{NO}_2 \) concentration of 5.53E-2 M. If the container is 1100 ml in volume at 30˚C, then a) what will be \( \text{Kc} \) for the equilibrium?  
And b) If .023 mol of \(^{250}\text{NO} \) is added to the flask, then how many moles of \(^{14}\text{CO}_2 \) (g) will form in the aqueous solution?
Q14.73

for the reaction \( V^{3+}(aq) + Cr^{2+}(aq) \rightleftharpoons V^{2+}(aq) + Cr^{3+}(aq) \) (K_c=7.2 \times 10^2) with initial concentrations \([V^{3+}] = [Cr^{2+}] = 0.0231 \text{ M} \) and \([V^{2+}] = [Cr^{3+}] = 0.260 \text{ M}\), what happens to the concentrations if the volume of the solution, 1L, is diluted to 2L with water?

Q14.75

Given the following generic reaction at 670 K what will be the total pressure at equilibrium if one starts with gas \( A \) at 0.837 atm?

\[ 2A(g) \rightleftharpoons 2B(g) + C(g) \]

with \( K_p = 1.3 \times 10^{-4} \)

Q14.77

Calculate the equilibrium amounts of the previous equation \( 2A \rightleftharpoons 2B + C \) with a K_c = 1.5 \times 10^2 if the initial \([A] = 1.3 \text{ M}\).

Q14.79

based on the reaction \( 2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \Delta H < 0 \), explain how each of the following will affect the amount of \( SO_3(g) \)

a. Increase pressure
b. Increase temperature
c. Increase the amount of \( O_2(g) \)

Q14.84

what will be the molar mass of the resulting gaseous mixture in the reaction \( COCl_2(g) \rightleftharpoons CO(g) + Cl_2(g) \) given that \( K_p = 4.44 \times 10^{-2} @122K \) and \( P_{tot} = 2.5 \text{ atm} \)?

Q14.102

given the following equilibrium equation, \( A+B \rightleftharpoons C \) with equal concentrations of \( A \) and \( C \) and a K_c=1000, what must the concentration of \( C \) be at equilibrium?

Q14.104

For the reaction \( 2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \) with \( K_p = 1.37 \times 10^{-2} \) at 1500°C. What do you expect the relationship to be between \( K_c \) and \( K_p \)?
SOLUTIONS

1) a) \( \text{H}_2\text{O} \leftrightarrow \text{H}^+ + \text{OH}^- \); \( k_c = [\text{H}^+][\text{OH}^-] \)

b) \( \text{Co}(l) + \text{NO}_2(l) \leftrightarrow \text{NO}(l) + \text{CO}_2(l) \); \( k_c = [\text{NO}][\text{CO}_2]/[\text{CO}][\text{NO}_2] \)

c) \( \text{HI}(g) + \text{H}_2\text{O}(l) \leftrightarrow \text{H}_3\text{O}^+(aq) + \text{I}^- (aq) \); \( k_c = [\text{H}_3\text{O}^+][\text{I}^-]/[\text{HI}] \)

2) a) \( \text{N}_2(g) + 3\text{H}_2(g) \leftrightarrow 2\text{NH}_3(g) \) \( K_p = \frac{P^2(\text{NH}_3)/P(\text{N}_2)P^3(\text{H}_2)} \)

b) \( \text{CaCO}_3 \leftrightarrow \text{CaO}(s) + \text{CO}_2(g) \) \( K_p = P(\text{CO}_2) \)

c) \( \text{CO}_2(g) + \text{NaOH}(aq) \leftrightarrow \text{NaHCO}_3(aq) \) \( K_p = \frac{[\text{NaHCO}_3]/P(\text{CO}_2)[\text{NaOH}]} \)

3) a) \( K_c = [\text{HIO}_2]^2/[\text{HI}]^2[\text{O}_2] \)

b) \( K_c = [\text{NO}_2][\text{ClNO}]/[\text{ClNO}_2][\text{NO}] \)

c) \( K_c = [\text{HI}]^2/[\text{H}_2][\text{I}_2] \)

29) \( 2\text{SO}_2(g) + \text{O}_2(g) \leftrightarrow 2\text{SO}_3(g) \)

\( K_c = 20 = \frac{(2.1/7.2)/(2.7/7.2)^*(3.2/7.2))}{1.75 < K_c = 20} \)

Therefore the mixture cannot be maintained indefinitely.

30) The reaction will proceed towards the products because \( K_c \) is much greater than one.

31)

a) \( K_c = 2.53E2 = \frac{.67 \text{ mol}/1.7 \text{ L}}{(.241 \text{ mol}/1.7 \text{ L})^*(.197 \text{ mol}/1.7 \text{ L})} = 23.99 \)

23.99 is not equal to 2.53E2, therefore it is not at equilibrium.

b) \( K_c > Q \) therefore there will be a shift to the right, towards the products.

33) \( \text{H}_2(g) + \text{Br}_2(g) \leftrightarrow 2\text{HBr}(g) \) \( K_c = 50.2 \) at 500 degrees C

\[ \begin{array}{c|ccc} \text{I} & .263 & .263 & 0 \\ \text{c} & -x & -x & +2x \\ \text{e} & .263-x & .263-x & 0+2x \end{array} \] 50.2 = \frac{(2x/4 \text{ L})^2}{((.263-x)/4 \text{ L})^2}

\( x = .205 \)

\( [\text{H}_2]_{eq} = .263-.205 = .058 \text{ M} \)
\[ [\text{Br}_2]_{eq} = 0.263 - 0.205 = 0.058 \text{ M} \]

\[ [\text{HBr}]_{eq} = 0 + 0.205 = 0.205 \text{ M} \]

35) \( \text{CO(g)} + \text{O}_2(g) \rightleftharpoons \text{CO}_2(g) \) \( \text{K}_c = 1.2 \times 10^3 \)

\[
\begin{array}{ccc}
\text{i} & 0.352 & 0.067 \\
\text{c} & +x & +x \ -x \\
\text{e} & 0.352 + x & 0.067 - x \\
\end{array}
\]

\[ \text{K}_c = 1.2 \times 10^3 = \frac{((0.067 - x)/3 \text{ L})/((0.352 + x)/3 \text{ L})(x/3 \text{ L})}{((0.352 + x)/3 \text{ L})(x/3 \text{ L})} \]

\[ x = 5.15 \times 10^{-4} \text{ mols O}_2 \]

37)

a) \( \text{CaCO}_3(g) \rightleftharpoons \text{CaCO}(g) + \text{CO}_2(g) \) \( \text{K}_c = 3.8 \times 10^{-2} \)

\[
\begin{array}{ccc}
\text{i} & 0.731 & 0 \\
\text{c} & -x & +x \ +x \\
\text{e} & 0.731 - x & 0.731 + x \ x \\
\end{array}
\]

\[ \text{K}_c = 3.8 \times 10^{-2} = \frac{(x/2.5 \text{ L})(0.731 + x/2.5 \text{ L})}{(0.731 - x/2.5 \text{ L})} \]

\[ x = 0.268 \]

\[ n(\text{CaCO}_3) = 0.731 - 0.268 = 0.463 \text{ mols} \]

\[ n(\text{CaCO}) = 0.731 + 0.268 = 0.999 \text{ mols} \]

\[ n(\text{CO}_2) = x = 0.268 \text{ mols} \]

b) \( \text{CaCO}_3(g) \rightleftharpoons \text{CaCO}(g) + \text{CO}_2(g) \) \( \text{K}_c = 3.8 \times 10^{-2} \)

\[
\begin{array}{ccc}
\text{i} & 0.715 & 0 \\
\text{c} & -x & +x \ +x \\
\text{e} & 0.715 - x & x \ x \\
\end{array}
\]

\[ \text{K}_c = 3.8 \times 10^{-2} = \frac{(x^2/2.5 \text{ L})}{(0.715 - x/2.5 \text{ L})} \]

\[ x = 0.147 \]
\[ n(\text{CaCO}_3) = 0.715 - 0.147 = 0.568 \text{ mols} \]

\[ n((\text{CaCO}_x) = 0.147 \text{ mols} \]

\[ n(\text{CO}_2) = x = 0.147 \text{ mols} \]

39) \( \text{Ca(OH)}_2 + 2 \text{CH}_3\text{COOH} = \text{Ca(CH}_3\text{COO})_2 + 2\text{H}_2\text{O} ; K_c = 4.0 \)

a) \[ n(\text{Ca(OH)}_2) = \frac{19.3 \text{g}}{74.093 \text{ g/mol}} = 0.260 \text{ mols} \]

\[ n(\text{CH}_3\text{COOH}) = \frac{21.7 \text{g}}{60.05 \text{ g/mol}} = 0.361 \text{ mols} \]

\[ n(\text{Ca(CH}_3\text{COO})_2) = \frac{41.2}{158.1669 \text{ g/mol}} = 0.260 \text{ mols} \]

\[ n(\text{H}_2\text{O}) = \frac{70.1}{18.02 \text{ g/mol}} = 3.89 \text{ mol} \]

\[ Q_c = \frac{(0.260)(3.89)}{(0.361)(0.260)} = 10.78 \]

\[ K_c = 4.0 < 10.78 = Q_c \]

Therefore the reaction will proceed to the left, towards reactants.

b) \( \text{Ca(OH)}_2 + 2 \text{CH}_3\text{COOH} = \text{Ca(CH}_3\text{COO})_2 + 2\text{H}_2\text{O} ; K_c = 4.0 \)

\[ \begin{array}{cccc}
1 & 0.260 & 0.361 & 0.260 & 3.89 \\
\text{c} & +x & +2x & -x & -2x \\
e & .26+x & .361+2x & .26-x & 3.89-2x \\
\end{array} \]

\[ K_c = 4.0 = \frac{(0.26-x)(3.89-2x)}{(0.26+x)(0.361+2x)} \]

\[ X = 0.184 \]

\[ \text{Mass(}\text{Ca(OH)}_2) = 0.260 + 0.184 = 0.444(74.093) = 32.9 \text{ g} \]

\[ \text{Mass(}\text{CH}_3\text{COOH}) = 0.361 + 2(0.184) = 0.729(60.05) = 43.77 \text{ g} \]

\[ \text{Mass(}\text{Ca(CH}_3\text{COO})_2) = 0.26 - 0.184 = 0.076(158.1669) = 12.02 \text{ g} \]

\[ \text{Mass(H}_2\text{O}) = 3.89 - 2(0.184) = 3.522(18.02) = 63.47 \text{ g} \]

41) \( \text{CaCO}_3(g) \rightarrow \text{CaO(g) + CO}_2(g) ; K_c = 4.84 \)
\[ K_c = 4.84 = \frac{x^2}{0.146-x} \]

\[ x = 0.142 \]

**Equilibrium concentrations.**

\[ [\text{CaCO}_3] = 0.146 - 0.142 = 0.004 \text{ M} \]

\[ [\text{CaO}] = 0.142 \text{ M} \]

\[ [\text{CO}_2] = 0.142 \text{ M} \]

\[ M_{\text{tot}} = 0.288 \]

\[ P_{\text{tot}} = (0.288)(0.08206)(300) = 7.1 \]

43)

\[ Q = \frac{4.31E-5}{0.00217} = 0.0199 \]

\( K > Q \) and therefore the reaction will proceed to the left (Not in initially in equilibrium)

45)

HNO\text{\textsubscript{3}} is a strong acid and therefore dissociates into its ions completely. Hence the ion concentrations at equilibrium are

\[ [\text{H}^+] = 2 \text{M} = [\text{NO}_3^-] \]

53)

\[ P(O) = nRT/V = 1 \text{mol} \times 0.08206 \text{ L atm mol}^{-1} \text{K}^{-1} \text{ atm} \times 600 \text{ K}/1 = 49.2 \text{ atm} \]

55)

a) Decreasing the temperature of the mixture

- Because the reaction is exothermic heat can be thought of as a product and therefore, if the temperature decreases, then the amount of H\text{\textsubscript{2}}O(g) will increase.

b) Introducing more Fe(s)

- There should be no change in the amount of H\text{\textsubscript{2}}(g)

c) Halving the volume of the container (increasing pressure)

- This will cause a shift from the side with the most moles of gas to the side with the least moles of gas. IN this case they are the same so there will be no change.
d) Adding a catalyst
- No change, catalysts only affect the speed at which the reaction occurs.

63) a) Less H$_2$O(g) formed at lower temperatures
b) The rate of the reaction slows with the decrease in temperature

65) The equilibrium will shift to the side with the most mols and therefore P$_{(N2)}$ should increase.

71) a) [NO]=$n/v=P/RT= .96/ .0862*303= .039 \ M \ Kc=5.53E-2/.039 =1.418$

NO(g)$<-->NO(aq) \ Kc=1.418=((5.53E-2+x)/.1L)/((.062-x)/1L)$

$$\text{[I]} \ .039 \ 5.53E-2 \ x= .0029$$

$$+.023 -----$$

$$\text{[c]} \ -x +x$$

$$\text{[e]} \ (.062-x) \ 5.53E-2+x \ \text{total mols (NO(g))}=(1L)(.0591)=.0591 \ \text{mol}$$

$$\text{total mols(NO(aq))}=(.1L)(.0582)=.00582 \ \text{mol}$$

$$\text{Q14.73}$$

The redox reaction in solution

$$V^{3+} (aq)+Cr^{2+} (aq) \rightleftharpoons V^{2+} (aq)+Cr^{3+} (aq) \ (K_c=7.2 \times 10^2)$$

has an equilibrium constant of $\langle K_c=7.2 \times 10^2 \rangle$. If the reaction were initiated with initial concentrations $\langle[V^{3+}]=0.0231, \ M\rangle$ and $\langle[Cr^{2+}]=0.260, \ M\rangle$, what happens to the concentrations if the volume of the solution, 1L, is diluted to 1.5L with water?

Equation: $V^{3+}(aq)+Cr^{2+}(aq)\rightleftharpoons V^{2+}(aq)+Cr^{3+}(aq) \ (K_c=7.2E2)$

Equilib: .0231 .0231 .260 .260
Dilution: .01155 .01155 .13 .13

Change: -x -x +x +x

New eq: .01155-x .01155-x .13+x .13+x

7.2E2= (.13+x)^2/(.01155-x)^2

x=.0065

new concentrations:

[V^3+]=[Cr^{2+}]=.0231-.0065=.0166

V^{2+}]=[Cr^{3+}]=.260+.0065=.2665

Q14.75

\[[2A(g)\rightleftharpoons 2B(g)+C(g) ]\]

\[K_p=1.3E-4\]

\[I .837 0 0\]

\[c -2x +2x +x\]

\[e (.837-2x) 2x x\]

Kp=1.3E-4=((2x)^2(x))/(.837-2x)^2

X=.8365 atm

P_{tot}= P(A)+P(B)+P(C)=(.837-2(.8365))+(2(.8365))+(.8365)=3.91 atm

77) 2A\rightleftharpoons 2B+C \quad K_c= 1.5E2 \quad [A]= 1.3 \text{ M}

\[I 1.3 M 0 0\]

\[c -2x +2x +x\]

\[e 1.3-2x 2x x\]

Kc=1.5E-2=(2x)^2(x)/(1.3-2x)^2

X=.185

Equilibrium concentrations:
\[ [A] = 1.3 - 1.185 = 1.115 \text{ M} \]
\[ [B] = 2 \times 1.185 = 0.37 \text{ M} \]
\[ [C] = 0.185 \text{ M} \]

**Q14.79**

a) Increase pressure
- more SO\(_3\) will be produced

b) Increase temperature
-SO\(_3\) is reduced because heat is treated as a product in exothermic reactions

c) Increase the amount of O\(_2\)(g)
-more SO\(_3\) will be produced in order to use up the excess oxygen

84)

\[ \text{COCl}_2(g) \leftrightarrow \text{CO}(g) + \text{Cl}_2(g) \] given that \( K_p = 4.44 \times 10^{-2} \) @122K

\[
\begin{array}{c|c}
\text{I} & \text{P} \\
\hline
\text{II} & 0 \ 0 \\
\text{III} & -x \ +x \ +x \\
\text{IV} & P-x \ x \ x \\
\end{array}
\]

\[ K_p = 4.44 \times 10^{-2} = \frac{(x \times x)}{(2.5 - 2x)} \]

\[ x = 0.292 \]

\[ P_{\text{tot}} = 2.5 \text{ atm} = P-x + x = P + x \]

\[ P = 2.5 - x \]

\[ P(\text{COCl}_2) = 2.5 - 2x = 1.916 \text{ atm} \]

\[ P(\text{CO}) = P(\text{Cl}_2) = 0.292 \text{ atm} \]

\[ \text{M}_{\text{avg}} = \frac{(P(\text{CO}) \times M(\text{CO}) / P_{\text{tot}}) + (P(\text{Cl}_2) \times M(\text{Cl}_2) / P_{\text{tot}}) + (P(\text{COCl}_2) \times M(\text{COCl}_2) / P_{\text{tot}})}{} = 87.36 \text{ g/mol} \]
Q102

A + B → C

K = \frac{[C]^2}{[A]^2[B]} = 1000

[B] = \frac{1}{1000} = 0.001 \text{ M}

104) because K_p = K_c \left(\frac{RT}{\Delta n}\right), we can determine that K_p > K_c

Q1

Based on these descriptions, write a balanced equation and the corresponding expression for each reversible reaction.

a. Carbonyl fluoride decomposes into gaseous carbon dioxide and gaseous carbon tetra fluoride.

b. Copper metal displaces silver (I) ion from aqueous solution, producing silver metal and an aqueous solution of copper (II) ion.

c. Peroxodisulfate ion oxidizes iron (II) ion to iron (III) ion in aqueous solution and is itself reduced to sulfate ion.

Q2

Based on these descriptions, write a balanced equation and the corresponding K_p expression for each reversible reaction.

a. Oxygen gas oxidizes gaseous ammonia to gaseous nitrogen and water vapor.

b. Hydrogen gas reduces gaseous nitrogen dioxide to gaseous ammonia and water vapor.

c. Nitrogen gas reacts with the solid sodium carbonate and carbon to produce solid sodium cyanide and carbon monoxide gas.

Q3

Write equilibrium constant expressions, \( K_c \), for the reactions.

Q29

Can a mixture of 2.2 mol O_2, 3.6 mol SO_2, and 1.8 mol SO_3 be maintained indefinitely in a 7.2 L flask at a temperature at which K_c = 100 in this reaction?

Q30

Is a mixture of 0.0205 mol NO_2 (g) and 0.750 mol N_2O_4 (g) in a 5.25 L flask at 25 °C, at equilibrium? If not, in which direction will the reaction proceed toward products or reactants?
Q31

In the reaction \(2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{SO}_3(\text{g})\), 0.455 mol SO\(_2\), 0.183 mol O\(_2\), and 0.568 mol SO\(_3\) are introduced simultaneously into a 1.90 L vessel at 1000 K.

a. If \(K_c = 2.8 \times 10^2\), is this mixture at equilibrium?
b. If not, in which direction will a net change occur?

Q33

A mixture consisting of 0.150 mol H\(_2\) and 0.150 mol I\(_2\) is brought to equilibrium at 445 °C, in a 3.25 L flask. What are the equilibrium amounts of H\(_2\), I\(_2\), and HI?

Q35

Starting with 0.3500 mol CO (g) and 0.05500 mol COCl\(_2\) (g) in a 3.050 L flask at 668 K, how many moles of Cl\(_2\) (g) will be present at equilibrium?

Q37

Equilibrium is established in a 2.50 L flask at 250 degrees Celsius for the reaction

\[
\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2 (\text{g})
\]

with \(|K_c = 3.8 \times 10^{-2}|\). How many moles of \(|\text{PCl}_5|\), \(|\text{PCl}_3|\), and \(|\text{Cl}_2|\) are present at equilibrium, if

a. 0.550 mol each of \(|\text{PCl}_5|\) and \(|\text{PCl}_3|\) are initially introduced into the flask?
b. 0.610 mol \(|\text{PCl}_5|\) alone is introduced into the flask?

Q39

In the following reaction, \(k_c = 4.0\).

\[
\text{C}_2\text{H}_5\text{OH} + \text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{COOC}_2\text{H}_5 + \text{H}_2\text{O}
\]

A reaction is allowed to occur in a mixture of 17.2 g C\(_2\)H\(_5\)OH, 23.8 g CH\(_3\)COOH, 48.6 g CH\(_3\)COOC\(_2\)H\(_5\), and 71.2 g H\(_2\)O.

a. In what direction will a net change occur?
b. How many grams of each substance will be present at equilibrium?

Q41

Formamide, used in the manufacture of pharmaceuticals, dyes, and agricultural chemicals, decomposes at high
temperatures.

\[ \text{HCONH}_2 (g) \rightleftharpoons \text{NH}_3 (g) + \text{CO} (g) \]

Kc = 4.84 at 400 K

If 0.186 mol HCONH$_2$(g) dissociates in a 2.16 L flask at 400 K, what will be the total pressure at equilibrium?

Q43

Cadmium metal is added to 0.350 L of an aqueous solution in which [Cr$^{3+}$] = 1.00 M. What are the concentrations of the different ionic species at equilibrium? What is the minimum mass of cadmium metal required to establish this equilibrium?

Q45

One sketch below represents an initial non-equilibrium mixture in the reversible reaction

SO$_2$ (g) + Cl$_2$ (g) $\rightarrow$ SO$_2$Cl$_2$ (g)

Kc = 4.0

Which of the other three sketches best represents an equilibrium mixture? Explain.

Q53

1.00 mol each of CO and Cl$_2$ are introduced into an evacuated 1.75 L flask and the following equilibrium is established at 668 K.

CO (g) + Cl$_2$ (g) $\rightarrow$ COCl$_2$ (g)

Kp = 22.5

For this equilibrium, calculate (a) the partial pressure of COCl$_2$ (g); (b) the total gas pressure.

Q55

Continuous removal of one of the products of a chemical reaction has the effect of causing the reaction to go to completion. Explain this fact in terms of Le Châtelier’s principle.

63

What effect does increasing the volume of the system have on the equilibrium condition in each of the following reactions?
Q71

Explain why the percent of molecules that dissociate into atoms in reactions of the type \(I_2(g) \rightleftharpoons 2I(g)\) always increases with an increase in temperature.

Q73

Refer to Example 15-13. Suppose that 0.100 L of the equilibrium mixture is diluted to 0.250 L with water. What will be the new concentrations when equilibrium is re-established?

Q75

Starting with \(SO_3(g)\) at 1.00 atm, what will be the total pressure when equilibrium is reached in the following reaction at 700 K?

77

Derive, by calculation, the equilibrium amounts of \(SO_2\), \(O_2\), and \(SO_3\) listed in (a) Figure 15-6(c); (b) Figure 15-7(b).

79

One of the key reactions in the gasification of coal is the methanation reaction, in which methane is produced from synthesis gas a mixture of \(CO\) and \(H_2\).

\[CO (g) + 3H_2 (g) \rightleftharpoons CH_4 (g) + H_2O (g)\]

\(\Delta{H} = - 230 \text{ kJ/mol}\) and \(K_c = 190\) at 1000 K

a. Is the equilibrium conversion of synthesis gas to methane favored at higher or lower temperatures? Higher or lower pressures?

b. Assume you have 4.00 mol of synthesis gas with a 3:1 mol ratio of \(H_2 (g)\) to \(CO (g)\) in a 15.0 L flask. What will be the mole fraction of \(CH_4 (g)\) at equilibrium at 1000 K?

84.

For the synthesis of ammonia at 500 K, 

\[N_2(g) + 3H_2 (g) \rightleftharpoons 2NH_3(g)\]

with \(K_p = 9.06 \times 10^{-2}\). Assume that \(N_2\) and \(H_2\) are mixed in the molar ratio of \(1:3\) in a \(1\) liter flask, respectively and that the total pressure is maintained at 1.00 atm. What is the mole percent \(NH_3\) at equilibrium?
102.

Equilibrium is established in the reaction $2\text{SO}_2 (g) + \text{O}_2 (g) \rightleftharpoons 2\text{SO}_3 (g)$ at a temperature where $K_c = 100$. If the number of moles of $\text{SO}_3 (g)$ in the equilibrium mixture is the same as the number of moles of $\text{SO}_2 (g)$, (a) the number of moles of $\text{O}_2 (g)$ is also equal to the number of moles of $\text{SO}_2 (g)$; (b) the number of moles of $\text{O}_2 (g)$ is half the number of moles of $\text{SO}_2$; (c) $[\text{O}_2]$ may have any of several values; (d) $[\text{O}_2] = 0.010 \text{ M}$.

104.

For the reaction

$$[2 \text{NO}_2 (g) \rightleftharpoons 2\text{NO} (g) + \text{O}_2 (g)]$$

\(K_c = 1.8 \times 10^{-6}\) at 184 °C.

At 184 °C, the value of \(K_c\) for the reaction

$$[\text{NO} (g) + 3/2 \text{O}_2 (g) \rightleftharpoons \text{NO}_2 (g)]$$

is:

a. 0.9 \times 10^6;

b. 7.5 \times 10^2;

c. 5.6 \times 10^5;

d. 2.8 \times 10^5.

Solutions

S1

S2

Video Solution

S3

29. Cannot be maintained indefinitely in a 7.2L flask.

30. Shift forward to products.

31. No. Forward to product.
33. I₂: 0.0330mol
HI: 0.234mol
35. 4.0*10⁻⁴ mol
37. (a) PCl₃: 0.623mol
PCl₅: 0.478mol
(b) PCl₃: 0.20mol
PCl₅: 0.41mol
39. (a) towards left
(b) 68g
41. 5.58atm
43. 10.1g
45. b
53. (a) 30.14 atm
(b) 32.4 atm
55. Continuous removal of the product of a chemical reaction has the effect of causing the reaction to go to completion that is which makes Q less than K and reaction shift the right as result and change in pressure of a constant temperature equilibrium mixture.
63. (a) Volume
(b) No effect
(c) More reactions
71. The percent of molecules that dissociates into atoms on reaction of the I₂ and 2I (g) always increase in temperature.
73. 2.98
75. 1.016 atm

S77.
   a. 0.27
   b. 0.085mol
Video Solution (Part A)

Video Solution (Part B)

a. low temperatures and high pressures favor conversion of synthesis gas to methane.

b. 0.134

28.06%

102
d

104. 750 (b)

1

Write a balanced equation that describes sodium chloride splits into Sodium and Chloride Ions.

\[ \text{NaCl} \rightarrow \text{Na}^{+} + \text{Cl}^{-} \]

2

Write a balanced equation that describes the decomposition of Hydrochloric Acid into its Ions.

\[ \text{HCl} \rightarrow \text{H}^{+} + \text{Cl}^{-} \]

3

Write equilibrium constant expressions, \( K_c \), for the reactions

a. \( \text{C} + 2\text{H}_2 \rightarrow \text{CH}_4 \)

b. \( 2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl} \)

c. \( \text{NaHCO}_3 \rightarrow \text{Na} + \text{HCO}_3^{-} \)

S3

(a) \( K_c = \frac{[\text{CH}_4]}{[\text{H}_2]^2} \)
(b) \( K_c = [\text{NaCl}]^2/\text{Cl}_2 \)

(c) \( K_c = [\text{HCO}_3]/[\text{NaHCO}_3] \)

### 29.

There are two gases in a 4.5 L vessel. The mass of \( C_6H_{16} \) is 6.68g and the mass of \( C_3H_8 \) is 2.28g. What is the value of \( K_c \) for this reaction.

**S29**

\[
\text{Mol } C_6H_{16} = 6.68 \text{g} \times \frac{1 \text{mol } C_6H_{16}}{88.22 \text{g } C_6H_{16}} = 0.0757 \text{ mol}
\]

\[
[C_6H_{16}] = \frac{0.0757 \text{ mol}}{4.5 \text{ L}} = 0.0168 \text{ M}
\]

\[
\text{Mol } C_3H_8 = 2.28 \text{ g } C_3H_8 \times \frac{1 \text{ mol } C_3H_8}{44.11 \text{ g } C_3H_8} = 0.0517 \text{ mol}
\]

\[
[C_3H_8] = \frac{0.0517 \text{ mol}}{4.5 \text{ L}} = 0.0115 \text{ M}
\]

\[
K_c = [C_3H_8]/[C_6H_{16}] = \frac{0.0115}{0.0168} = 0.685
\]

### 30.

Is a mixture of 1.30 L of water and 3.15 mol of HCN and .001 mol of CN- in equilibrium? If not, in which direction will the shift proceed, towards products or reactants?

**S30**

\( \text{K}_a (HCN) = 6.2 \times 10^{-10} \)

\[
[HCN] = \frac{3.15 \text{ mol}}{1.30 \text{ L}} = 2.42 \text{ M} \quad [CN^-] = [H+] = 7.69 \times 10^{-4} \text{ M}
\]

\[
Q = [CN^-] \times [H+] / [HCN] = (7.69 \times 10^{-4})^2 / (2.42) = 2.45 \times 10^{-7}
\]

\( Q > K \), therefore... reaction shifts left, towards the reactants.

### 31.

For the reaction in equilibrium,

\[
\text{\textit{H}_2\text{SO}_4 + \text{iH}_2\text{O } \rightleftharpoons \text{HSO}_4^- + \text{iH}_3\text{O}^+}\]

if the \( \text{K}_a \) value for \( \text{(H}_2\text{SO}_4 \text{)} \) is \( \times 10^{-5} \), the molar concentration of \( \text{(H}_2\text{SO}_4 \text{)} \) is 0.056 M, and the pH of the solution is 4.5, what is the concentration of \( \text{(HSO}_4^-\text{)} \)?
### S31

\[ [\text{H}_3\text{O}^+] = 10^{-4.5} = 3.16 \times 10^{-5} \text{ M} \]

\[ 2.42 \times 10^{-5} = [\text{HSO}_4^-] \times (3.16 \times 10^{-5}) / (0.056) \]

\[ [\text{HSO}_4^-] = 0.0429 \text{ M} \]

### S33

In an exothermic reaction increasing the temperature shifts the equilibrium in which direction?

**S33**

Left, more heat is more products so it shifts towards reactants.

### S35

If the number of moles remains constant, a change in volume does what to the equilibrium of a system?

**S35**

Nothing, \( K_{eq} \) also remains constant.

### S37

Using the above equilibrium reaction at 25 degrees Celsius with \( K_c = 7.1 \times 10^2 \)

\[ \text{H}_2 \text{ (g)} + \text{I}_2 \text{ (g)} \rightleftharpoons 2 \text{ HI (g)} \]

a. find the molarity of \( \text{HI} \), assuming an equilibrium concentration of 0.76 M \( \text{I}_2 \), and 0.32 M \( \text{H}_2 \)

b. find the molarity of \( \text{I}_2 \), assuming an equilibrium concentration of 2.36 M \( \text{HI} \), and 0.42 M \( \text{H}_2 \)

**S37**

a) \( K_c = [\text{HI}]^2 / [\text{I}_2] \times [\text{H}_2] \)

\[ 7.1 \times 10^2 = [\text{HI}]^2 / (0.76 \text{ M} \times 0.32 \text{ M}) \]

\[ = 13.14 \text{ M HI} \]

b) \( K_c = [\text{HI}]^2 / [\text{I}_2] \times [\text{H}_2] \)

\[ 7.1 \times 10^2 = (2.36 \text{ M})^2 / (0.42 \times [\text{I}_2]) \]

\[ (2.36 \text{ M})^2 / (0.42 \times (7.1 \times 10^2)) = [\text{I}_2] \]
In the following equilibrium reaction, \(K_p = 3.16\):
\[
[2 \text{PO}_4 \rightleftharpoons 2 \text{PO}_3 + \text{O}_2]\]

A reaction occurs at 273 K in a 1.0 L container with 4.56 g of \(\text{PO}_4\), 6.54 g of 2 \(\text{PO}_3\) and 1.32 g of \(\text{O}_2\).

a. In what direction will a net change occur?
b. What will be the partial pressures of each substance at equilibrium?

**S39**

A)

\[
n = (4.56 \text{ g PO}_4)(1 \text{ mol} / 94.97 \text{ g}) = 0.0480 \text{ mol PO}_4
\]

\[
n = (6.54 \text{ g PO}_3)(1 \text{ mol} / 78.97 \text{ g}) = 0.0828 \text{ mol PO}_3
\]

\[
n = (1.32 \text{ g O}_2)(1 \text{ mol} / 32 \text{ g}) = 0.0413 \text{ mol O}_2
\]

\[
PV=nRT
\]

\[
P_{\text{PO}_4} = (0.0480 \text{ mol})(0.08206)(273 \text{ K}) / (1.0 \text{ L}) = 1.08 \text{ atm}
\]

\[
P_{\text{PO}_3} = (0.0828 \text{ mol})(0.08206)(273 \text{ K}) / (1.0 \text{ L}) = 1.85 \text{ atm}
\]

\[
P_{\text{O}_2} = (0.0413 \text{ mol})(0.08206)(273 \text{ K}) / (1.0 \text{ L}) = 0.925 \text{ atm}
\]

\[
Q = (1.85)^2 * (0.925) / (1.08)^2 = 2.71
\]

2.71 < 3.16... Q < \(K_p\)... therefore... reaction shifts right, towards products.

B)

\[
K = 3.16 = (1.85 + 2x)^2 * (0.925+x) / (1.08 – 2x)^2
\]

\[x = 0.0218\]

\[
P_{\text{PO}_4} = 1.08 – 2(0.0218) = 1.04 \text{ atm}
\]

\[
P_{\text{PO}_3} = 1.85 + 2(0.0218) = 1.89 \text{ atm}
\]

\[
P_{\text{O}_2} = 0.925 + 0.0218 = 0.947 \text{ atm}\]
Dinitrogen tetroxide, $(\text{N}_2\text{O}_4)$ in combination with specific forms of hydrazine can burn on contact without a separate ignition source, make it popular as a component of rocket fuel. At high temperatures the equilibrium is:

$$\text{N}_2\text{O}_4 \rightleftharpoons 2 \text{NO}_2$$

What are the partial pressures at equilibrium at 83 Celsius and 1.1 atm, if the density of the two equilibrium products is 2.44 g/l?

**S41**

**Knowns:**
- $p=1.1$ atm
- $W=\text{average molar mass}$
- $R=0.08206$ atm-L/mol-K
- $T=83+273=356$K

**Calculations:**

\[
p \cdot W = \text{density} \cdot RT
\]

\[
1.1 \text{ atm} \cdot W = (2.44 \text{ g/L}) \cdot (0.08206 \text{ atm-L/mol-K}) \cdot (356\text{K})
\]

\[
W = 64.8 \text{ g/mol}
\]

If $X$ is the mol fraction of N2O4, then $1 - X$ is the mol fraction of NO2

\[
92 \text{ g/mol} \cdot X + 46 \text{ g/mol} \cdot (1-X) = 64.8 \text{ g/mol}
\]

\[
X = 0.4087
\]

and

\[
1-X = 0.5913
\]

\[
p(\text{N2O4}) = 0.4087 \cdot 1.1\text{atm} = 0.4496 \text{ atm}
\]

\[
p(\text{NO2}) = 0.5913\cdot 1.1\text{atm} = 0.6504 \text{ atm}
\]

Therefore, the partial pressure of N2O4 is 0.4496 atm and the partial pressure of NO2 is 0.6504 atm.
What will a catalyst do to the composition of an equilibrium mixture?

Nothing, catalysts effect time, but once at equilibrium a system won’t take effect to a catalyst.

For the reaction

\[ \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2 \text{HI}(g) \]

\( K_c = 50.3 \) at 731 K. Equal amounts (1.00 M each of the three components) are added to a beaker, and then the temperature is raised to 731 K.

Calculate the equilibrium concentrations.

**S45**

\[ \text{H}_2(g) + \text{I}_2(g) = 2 \text{HI}(g) \]

1.00 1.00 1.00 initial concentration

\(-x\) \(-x\) 2x amount of change

1-x 1-x 1+2x equilibrium concentration

\( K_c = 50.3 = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(1+2x)^2}{(1-x)^2} \)

\( \therefore (1+4x+4x^2) = 50.3(1-2x-x^2) = 50.3-100.6x-50.3x^2 \)

\(-46.3x^2 + 104.6x - 49.3 = 0\)

Solving the quadratic equation: The roots are: 1.589 and .670

Since \( x \) cannot be greater than 1, the only solution is \( x = .670 \).

Therefore, the equilibrium concentrations are:

\[ [\text{H}_2] = 1-.670 = .33 \]

\[ [\text{I}_2] = 1-.670 = .33 \]

\[ [\text{HI}] = 1+2*.670 = 2.34 \]

Determine the partial pressure of the Hydrogen in the following reaction held within a 2L flask
\[H^+ + Br^- \rightleftharpoons HBr\]

P[Br] = (nRT)/(V) = ((1mol)(0.08206)(668K))/(2L) = 27.4 atm

55.

What is Le Chatelier’s principle?

Answer:

Le Chatelier’s principle is when a new equilibrium has occurred due to a change of change in temperature or pressure on a chemical equilibrium. It can be known as the Equilibrium Law.

63

What affect does decreasing the volume of the system have on the equilibrium condition in each of the following reactions:

a. \(2 H_2(g) + C(s) \rightleftharpoons CH_4(g)\)
b. \(2 CO_2 (g) \rightleftharpoons 2 CO (g) + O_2 (g)\)
c. \(2 N_2O (g) + O_2 (g) \rightleftharpoons 4 NO (g)\)

S63

a. shifts right
b. shifts left
c. shifts left

65

Hb(aq) + O2(aq) àHb:O2(aq) \(\Delta H<0\)

How does decreasing temperature effect this system?

S65

Because the reaction is endothermic, we know that decreasing the temperature adds more to the reactants side, shifting the equilibrium towards the products.

69.

The standard enthalpy of reaction for the combustion of ethanol is 1,418kJ/mole. Write the equation:

\[C_2H_5OH(l) + 3O_2 (g) \rightleftharpoons 2CO_2 (g) + 3H_2O(l)\]
How does pressure affect the equilibrium?

As 3 molecules of O₂ give 2 molecules of CO₂, increasing pressure drives the equilibrium to the right side of the reaction, as well as decrease the volume and increase the number of moles.

# 77

In a reversible reaction,

\[ 2Na + Cl_2 \rightleftharpoons 2NaCl \]

initial conditions indicate 2 mol \(Na\) and 1 mol \(Cl_2\). What is the expected amount of NaCl at equilibrium?

Less than 2mols, because the \(Cl_2\) is the limiting reagent at 1mol, but the theoretical yield of NaCl is 2mols, so it can never hit that point unless more Cl is added.

79.

One of the key reactions in the gasification of coal is the ethanation reaction, in which ethane is produced from synthesis gas – a mixture of CO and H₂.

\[ 2CO(g) + 2H_2(g) \rightleftharpoons C_2H_4(g) + 2H_2O (g) \]

\(DH = -230\) kJ; \(K_c = 178\) at 1200K

1. Is the equilibrium conversion of synthesis gas to methane favored at higher or lower temperatures? Higher or lower pressures? Products are favored with higher temperatures and lower pressures.

90.

Consider the following reaction:

\[ PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g) \]

An equilibrium mixture at 345 Kelvin in a 6.50 L vessel contains:

- 1.35 mol \(PCl_3\),
- 0.450 mol \(Cl_2\), and
- 3.20 mol \(PCl_5\).

The equilibrium constant for the reaction is \(K_p = 1.21\). If 0.50 mol of PCl₅ are removed from the system, what will the pressures of each species be once the reaction returns to equilibrium?

S90

\[ PV = nRT \]

\[ P = \frac{nRT}{V} \]
\[ \frac{(1.35 \text{ mol}) \cdot (0.08206) \cdot (345 \text{ K})}{6.50 \text{ L}} = 5.88 \text{ atm} \]

\[ \frac{(0.450 \text{ mol}) \cdot (0.08206) \cdot (345 \text{ K})}{6.50 \text{ L}} = 1.96 \text{ atm} \]

\[ \frac{(3.20 \text{ mol}) \cdot (0.08206) \cdot (345 \text{ K})}{6.50 \text{ L}} = 13.9 \text{ atm} \]

\[ \frac{(3.20 - 0.50 \text{ mol}) \cdot (0.08206) \cdot (345 \text{ K})}{6.50 \text{ L}} = 11.8 \text{ atm} \]

\[ 1.21 = \frac{11.8 + x}{5.88 - x} \cdot \frac{1.96 - x}{1.96 - x} \]

\[ x = 0.210 \]

\[ P_{\text{PCl}_3} = 5.88 - 0.210 = 5.67 \text{ atm} \]

\[ P_{\text{Cl}_2} = 1.96 - 0.210 = 1.75 \text{ atm} \]

\[ P_{\text{PCl}_5} = 11.8 + 0.210 = 12.01 \text{ atm} \]

103.

For the oxidation of ethanol:

\[ \text{C}_2\text{H}_5\text{OH}(l) + 3\text{O}_2 (g) \rightleftharpoons 2\text{CO}_2 (g)+ 3\text{H}_2\text{O}(l) \]

How does pressure affect the equilibrium?

S103

Answer: As 3 molecules of O2 give 2 molecules of CO2, increasing pressure drives the equilibrium to the right side of the equation, as well as decreasing the volume and the number of moles.

Q1

Based on these descriptions, write a balanced equation and the corresponding \( K_c \) expression for each reversible reaction:

a. Sulfur trioxide, \( \text{SO}_3(\text{g}) \), decomposes into gaseous oxygen and gaseous sulfur dioxide.

b. Zinc metal displaces silver(I) ion from aqueous solution, producing silver mean and an aqueous solution of zinc(II) ion.

c. Oxalate ion, \( \text{C}_2\text{O}_4^- \), oxidizes copper(I) ion to copper(II) ion in aqueous solution and is itself reduced to carbon dioxide.

Solution

a)

\[ 2 \text{SO}_3(\text{g}) \rightleftharpoons \text{O}_2(\text{g}) + 2 \text{SO}_2(\text{g}) \]

\[ K_c = \frac{[\text{SO}_2]^2 \cdot [\text{O}_2]}{[\text{SO}_3]^2} \]
b) 
\[ \text{\textit{Zn}}_{(s)} + 2\text{Ag}^{+}_{(aq)} \rightleftharpoons 2\text{Ag}_{(s)} + \text{Zn}^{2+}_{(aq)} \]  
\[ K_c = \dfrac{[\text{Zn}^{2+}]}{[\text{Ag}^{+}]^2} \]

\[ K_c = \dfrac{[\text{CO}_2]^2 [\text{Cu}]^{2+}}{[\text{C}_2 \text{O}_4^-][\text{Cu}^+]^2} \]

Q2

Based on these descriptions, write a balanced equation and the corresponding \(K_p\) expression for each reversible reaction:

a. \( \text{K}_2\text{O}_{(s)} + \text{CO}_{2(g)} \rightleftharpoons \text{KCO}_3 + \text{O}_{2(g)} \)

b. \( \text{HF}_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{F}^-_{(aq)} \)

c. \( 2\text{HCl}_{(g)} + \text{I}_2_{(g)} \rightleftharpoons 2\text{HI}_{(g)} + \text{Cl}_2{(g)} \)

Solution

a) 
\[ 4\text{K}_2\text{O}_{(s)} + 2\text{CO}_{2(g)} \rightleftharpoons 2\text{KCO}_3 + 3\text{O}_{2(g)} \]
\[ K_c = \dfrac{[\text{O}_2]^3}{[\text{CO}_2]^2} \]

b) 
\[ \text{HF}^-_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{F}^-_{(aq)} \]
\[ K_c = \dfrac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} \]

c) 
\[ 2\text{HCl}_{(g)} + \text{I}_2{(g)} \rightleftharpoons 2\text{HI}_{(g)} + \text{Cl}_2{(g)} \]
\[ K_c = \dfrac{[\text{HI}]^2 [\text{Cl}_2]}{[\text{HCl}]} \]

Q3

Write equilibrium constant expressions, \(\text{K}_c\), for the reactions:

a. \( 2\text{CO}^-_{(g)} + \text{O}^-_{2(g)} \rightleftharpoons 2\text{CO}_{2(g)} \)

b. \( \text{Fe}^{2+}_{(aq)} + 2\text{H}_2\text{O}_{(l)} \rightleftharpoons \text{FeOH}^{2+}_{(aq)} + \text{H}_3\text{O}^+_{(aq)} \)

c. \( \text{HNO}^-_{(2(aq))} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{NO}^-_{2{(aq)}} \)
Solutions

\[ K_c = \dfrac{[CO_2]^2}{[CO]^2[O_2]} \]

\[ K_c = \dfrac{[Fe(OH)][H_3O^+]}{[Fe^{3+}]} \]

\[ K_c = \dfrac{[H_3O^+][NO_2^-]}{[HNO_2]} \]

29

Can a mixture of 3.1 mol \( CO \), 3.9 mol \( O_2 \), and 2.2 mol \( CO_2 \) be maintained indefinitely in a \( 9.2 \; L \) container at a temperature at which \( K_c = 95 \) in this reaction (and why or why not):

\[ 2CO_{(g)} + O_{2 (g)} \rightleftharpoons 2CO_{2 (g)} \]

SOLUTION

\[ Q_C = \left( \dfrac{\left(\dfrac{2.2 \; mol \; [CO]}{9.2 \; L} \right)^2}{\left(\dfrac{3.1 \; mol \; CO}{9.2 \; L} \right)^2 \left(\dfrac{3.9 \; mol \; O_2}{9.2 \; L} \right)} \right) = 1.19 \]

1.19 < \( K_c = 95 \). Thus the mixture described cannot be maintained indefinitely. Because \( Q_C < K_c \), the reaction will proceed toward the products until the equilibrium is established.

30

Is a mixture of 0.0196 mol \( CO_{2 (g)} \) and 0.815 mol \( C_2O_{4 (g)} \) in a 6.15 L flask at 30°C at equilibrium? If not, in which direction will the reaction proceed: toward the reactants or products?

\[ C_2O_{4 (g)} \rightleftharpoons 2 CO_{2 (g)} \]

with

\[ K_c = 7.66 \times 10^{-5} \]

\[ Q = \dfrac{[CO_{2 (g)}]^2}{[C_2O_{4 (g)}]} \]

\[ = \dfrac{\left( \dfrac{0.0196 \; mol \; CO_2}{6.15 \; L} \right)^2}{\dfrac{0.815 \; mol \; C_2O_4}{6.15 \; L}} = 7.66 \times 10^{-5} \]

\[ \therefore Q = K \]

The reaction is already at equilibrium.

31

In the reaction:
\[2H_2(g) + O_2(g) \rightleftharpoons 2H_2O(l)\]

if 0.05 mol \(H_2\) and 0.02 mol \(O_2\), 20 mol of \(H_2O\) are introduced simultaneously into a 1.0 L vessel.

a. If \(K_c = 1.4 \times 10^{83}\) is the mixture initially at equilibrium.

b. If not, in which direction will a net change occur?

Solution

a) \(K_c = \dfrac{1}{[H_2(g)]^2 [O_2(g)]} = 1.4 \times 10^{83}\)

\(Q = \dfrac{1}{(0.05)^2(0.02)} = 4 \times 10^5\)

\(\therefore Q \neq K_c\)

The mixture is initially not at equilibrium.

b) The reaction will experience a net change in the direction of the products because since \(Q<K\).

33

A mixture consisting of 0.5 mol \(Cu^{2+}\) and 0.5 mol \(Sn^{2+}\) is brought to equilibrium in a 1 L flask. What is the equilibrium concentration of \(Sn^{4+}\). The reaction proceeds as follows:

\[2Cu^{2+}_{(aq)} + Sn^{2+}_{(aq)} \rightleftharpoons 2Cu^{+} + Sn^{4+}\]

with \(K_c= 1.48\)

Solution:

An ICE table is necessary to solve this problem.

<table>
<thead>
<tr>
<th>ICE</th>
<th>(2Cu^{2+}_{(aq)})</th>
<th>(Sn^{2+}_{(aq)})</th>
<th>(\rightleftharpoons)</th>
<th>(2Cu^{+})</th>
<th>(Sn^{4+})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.5 mol</td>
<td>0.5 mol</td>
<td></td>
<td>0 mol</td>
<td>0 mol</td>
</tr>
<tr>
<td>Change</td>
<td>-2x</td>
<td>-x</td>
<td></td>
<td>+2x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.5 mol - 2x</td>
<td>0.5 mol - x</td>
<td></td>
<td>+2x</td>
<td>+x</td>
</tr>
</tbody>
</table>

\(K_c = \dfrac{(x^2)(x)}{(0.5-2x)^2(0.5-x)} = 1.48\)

\(x=0.1752\)

\(\therefore [Sn^{4+}]=0.1752\)
Starting with 0.5 mol of \(\text{PCl}_5(g)\) and 1 mol of \(\text{Cl}_2(g)\), in a 2.0 L flask, how many moles of \(\text{PCl}_3(g)\) will be present at equilibrium?

\[
\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2 (g)
\]

with

\[
K_c = 3.8 \times 10^{-2}
\]

**Solution:**

Reaction is already balanced:

\[
\text{PCl}_5(g) = \text{PCl}_3(g)+ \text{Cl}_2 (g)
\]

An **ICE table** is necessary to solve this problem (constructed in terms of molarity, not mol).

<table>
<thead>
<tr>
<th>ICE</th>
<th>(\text{PCl}_5(g))</th>
<th>(\text{PCl}_3(g))</th>
<th>(\text{Cl}_2(g))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.25 mol</td>
<td>0 mol</td>
<td>0.5 mol</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.25 mol - x</td>
<td>+x</td>
<td>0.5 mol +x</td>
</tr>
</tbody>
</table>

\[
K_c = 3.8 \times 10^{-2}
\]

\[
3.8 \times 10^{-2} = \frac{x(0.5+x)}{0.25-x}
\]

\[
x = 0.0335
\]

\[
\therefore [\text{PCl}_3] = 0.0335 \text{ M}
\]

Equilibrium is established in a 3.0 L flask for the following reaction.

\[
2\text{H}_2\text{S} + \text{CH}_4 \rightleftharpoons 4\text{H}_2 + \text{CS}_2
\]

with

\[
K_c = 5.27 \times 10^{-8}
\]

How many moles of \(\text{CS}_2(g)\) and \(\text{CH}_4(g)\) are present at equilibrium in the two initial situations:

a. If 3 mol of \(\text{H}_2\text{S}((g))\) and \(\text{CH}_4(g)\) are initially introduced into the flask.
b. If 4 mol of \(4H_{2(g)}\) and \(CS_{2(g)}\) are initially introduced into the flask.

**Solution:**

The reaction is already balanced

\[
\begin{align*}
2H_2S_{(g)} + CH_{4(g)} & \rightleftharpoons 4H_{2(g)} + CS_{2(g)}
\end{align*}
\]

**Scenario A)**

An **ICE table** is necessary to solve this problem.

<table>
<thead>
<tr>
<th>ICE</th>
<th>(2H_2S_{(g)})</th>
<th>(CH_4(g))</th>
<th>(\text{rightleftharpoons})</th>
<th>(4H_2(g))</th>
<th>(CS_2(g))</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Initial</strong></td>
<td>1 M</td>
<td>1 M</td>
<td>0 M</td>
<td>0 M</td>
<td>0 M</td>
</tr>
<tr>
<td><strong>Change</strong></td>
<td>-x</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td><strong>Equilibrium</strong></td>
<td>1 M - 2x</td>
<td>1 M - x</td>
<td>+x</td>
<td>+x</td>
<td>+x</td>
</tr>
</tbody>
</table>

\[
5.27 \times 10^{-8} = \dfrac{(x^5)}{(1-x)^3}\]

\[x = 0.0343 = [CS_2] \]

\[(1 - 0.0343 = 0.9657 = [CH_4])\]

**Scenario B)**

An **ICE table** is necessary to solve this problem.

<table>
<thead>
<tr>
<th>ICE</th>
<th>(2H_2S_{(g)})</th>
<th>(CH_4(g))</th>
<th>(\text{rightleftharpoons})</th>
<th>(4H_2(g))</th>
<th>(CS_2(g))</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Initial</strong></td>
<td>0</td>
<td>0</td>
<td>4/3 M</td>
<td>4/3 M</td>
<td>4/3 M</td>
</tr>
<tr>
<td><strong>Change</strong></td>
<td>+x</td>
<td>+x</td>
<td>-x</td>
<td>-x</td>
<td>-x</td>
</tr>
<tr>
<td><strong>Equilibrium</strong></td>
<td>+x</td>
<td>+x</td>
<td>4/3 M - x</td>
<td>4/3 M - x</td>
<td></td>
</tr>
</tbody>
</table>

\[
5.27 \times 10^{-8} = \dfrac{(4/3-x)^5}{x^3}\]

\[x = 1.2925 = [CH_4] \]

\[4/3 - 1.2925 = 0.04083 = [CS_2] \]

39

The following reaction occurs in the presence of 10 grams of \((CH_4(g))\), 15 grams of \((C_2H_2(g))\), and 2 grams of \((H_2(g))\).
\[2\text{CH}_4(g) \rightleftharpoons \text{C}_2\text{H}_2(g) + 3\text{H}_2(g)\] 

with 

\[K_c = 0.154\]

What direction will a net change occur and now many grams of each substance will be present at equilibrium?

**Solution:**

First one must find the moles of each substance.

- 10 grams \(\text{CH}_4\)/16 grams/mol \(\text{CH}_4\) = 0.625 mol \(\text{CH}_4\)
- 15 grams \(\text{C}_2\text{H}_2\)/26 grams/mol = 0.5769 mol \(\text{C}_2\text{H}_2\)
- 2 grams \(\text{H}_2\)/2 grams/mol = 1 mol \(\text{H}_2\)

\[Q = \dfrac{0.5769 \times 1^3}{(0.625)^2} = 1.454 > 0.154\]

Since \(Q > K_c\), the reaction shifts to the left.

\[2\text{CH}_4(g) \rightleftharpoons \text{C}_2\text{H}_2(g) + 3\text{H}_2(g)\]

An **ICE table** is necessary to solve this problem (in terms of moles).

<table>
<thead>
<tr>
<th>ICE</th>
<th>(2\text{CH}_4(g))</th>
<th>(\text{C}_2\text{H}_2(g))</th>
<th>(3\text{H}_2(g))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.625</td>
<td>0.5769</td>
<td>1</td>
</tr>
<tr>
<td>Change</td>
<td>+2x</td>
<td>-x</td>
<td>-3x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.625 + 2x</td>
<td>0.5769 - x</td>
<td>1 - 3x</td>
</tr>
</tbody>
</table>

\[0.154 = \dfrac{(0.5769 - x)(1 - x)^3}{(0.625 + 2x)^3}\]

\((x = 0.277)\)

- \(\text{[CH}_4\) = 0.625 + 0.277 = 0.902\)
- \(\text{[C}_2\text{H}_2\) = 0.299\)
- \(\text{[H}_2\) = 0.722\)

41

Gaseous hydrogen iodide is placed in a closed container, where it partially decomposes.

\[2\text{HI}(g) \rightleftharpoons \text{H}_2(g) + \text{I}_2(g)\]
If 0.172 mol $\text{HI}_{(g)}$ is initially added to a 1.67 L flask at 698 K, what will be the total partial pressure at equilibrium?

Solution

- $\text{HI}_i = \frac{0.172 \text{ mol}}{1.67 \text{ L}} = 0.103 \text{ M}$

An ICE table is necessary to solve this problem (in terms of mole).

<table>
<thead>
<tr>
<th>ICE</th>
<th>$\text{2HI}_{(g)}$</th>
<th>$\text{H}_2(g)$</th>
<th>$\text{I}_2(g)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.172 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-2x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.172 M - 2x</td>
<td>+x</td>
<td>+x</td>
</tr>
</tbody>
</table>

$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = 5.5 \times 10^5$

$x^2 = 94600 - 5.5 \times 10^5 x = 0$

Solving the quadratic equation:

$x = -5.5 \times 10^5 \pm \sqrt{(3.03 \times 10^11 + 378400)(2)} = 227 \text{ M}, -550227 \text{ M}$

$[\text{total}] = [\text{H}_2] + [\text{I}_2] + [\text{2HI}] = x + x + 0.172 - x = 0.172 + 227 = 227 \text{ M}$

$P_{tot} = (227 \text{ mol})/(0.08206 \text{ L atm}/(\text{mol K}) \times 698 \text{ K} = 1.3 \times 10^4 \text{ atm}$

55

The continuous removal of reactants of a chemical reaction has the effect of causing the reaction to go backwards. Explain this in terms of Le Châtelier's principle.

Solution

Continuous removal of the reactants has the effect of decreasing the concentrations of the reactants below their equilibrium values. Therefore, the equilibrium system is disturbed by the removal of the reactant; the system will remedy this by trying to re-establish the equilibrium to the left to generate more reactants.

63

How would decreasing the volume of the following systems affect their equilibrium?

Decreasing the volume of a system causes the equilibrium to shift towards the side whose sum of the stoichiometric
coefficients of only the species in a gaseous state is smaller. If $\Delta n_{\text{gas}}$ is greater than 0, it shifts to the left. If $\Delta n_{\text{gas}}$ is less than 0, it shifts to the right.

a) $\text{CH}_4 (g) + 2 \text{H}_2 \text{O} (g) \rightleftharpoons \text{CO}_2 (g) + 4\text{H}_2 (g)$

$\Delta n_{\text{gas}}$ is 2, since there are 3 moles of reactants and 5 moles of products, the reaction shifts to the left, towards the reactants.

b) $\text{CaCO}_3 (s) \rightleftharpoons \text{CaO}(s) + \text{CO}_2 (g)$

$\Delta n_{\text{gas}}$ is 1, since we only count the moles of gas, there is only one mole of gas in the products. Thus, the reaction shifts towards the reactants.

c) $\text{C}(s) + \text{O}_2 (g) \rightleftharpoons \text{CO}_2 (g)$

$\Delta n_{\text{gas}}$ is 0, since we only count the number of moles of species in their gaseous state. Thus, there is no shift in equilibrium despite change in volume.

71

Explain why dissociation diatomic elements from molecules to elements, such as the reaction $\text{O}_2 (g) \rightleftharpoons 2\text{O}(g)$, will always increase with an increase in temperature.

First, let's think about this logically. Breaking bonds, especially strong bonds in diatomic elements, requires energy. When you add more energy to the system, there is a higher amount of energy, and more bonds being broken.

Second, let's consider the equilibrium of this reaction. If we add heat to the reactants, the equilibrium shifts to the right, towards the products.

$\text{O}_2 (g) + \text{heat} \rightleftharpoons 2\text{O}(g)$

$\rightarrow$ reaction shifts to products

1

Write a balanced equation & $K_c$ expression for each reaction.

a) $\text{CH}_4$ undergoes combustion to become gaseous carbon dioxide and liquid water

$$\text{CH}_4 + 2\text{O}_2 \rightleftharpoons \text{CO}_2 + 2\text{H}_2\text{O} ; K_c = \frac{[\text{CO}_2]}{[\text{CH}_4][\text{O}_2]^2}$$

b) $\text{Cl}_2\text{O}_5 (g)$ decomposes into $\text{Cl}_2$ and $\text{O}_2$

$$2\text{Cl}_2\text{O}_5 (g) \rightleftharpoons 2\text{Cl}_2 + 5\text{O}_2 ; K_c = \frac{[\text{Cl}_2]^2[\text{O}_2]^5}{[\text{N}_2\text{O}_5]^2}$$
c) $\text{N}_2\text{O}_5$ decomposes into $\text{O}_2$ and $\text{NO}_2$

$$2\text{N}_2\text{O}_5 \leftrightarrow \text{O}_2 + 4\text{NO}_2 ; K_c = \frac{[\text{NO}_2]^4[\text{O}_2]}{[\text{N}_2\text{O}_5]^2}$$

2

Based on the descriptions, write a balanced equation and its $K_p$ expression for each reaction

a) $\text{SO}_2 (g)$ is converted to sulphur trioxide

$$2\text{SO}_2(g) + \text{O}_2(g) \leftrightarrow 2\text{SO}_3(g) ; K_p = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]}$$

b) $\text{N}_2$ and $\text{H}_2$ are converted to $\text{NH}_3$

$$\text{N}_2(g) + 3\text{H}_2(g) \leftrightarrow 2\text{NH}_3(g) ; K_p = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

c) Solid carbon is heated with steam to produce hydrogen gas and carbon monoxide

$$\text{C}(s) + \text{H}_2\text{O}(g) \leftrightarrow \text{H}_2(g) + \text{CO}(g) ; K_p = \frac{[\text{H}_2][\text{CO}]}{[\text{H}_2\text{O}]}$$

3

Write the $K_c$ expression for:

a) $\text{CaCO}_3(s)$

$$\text{CaO}(s) + \text{CO}_2(g) \leftrightarrow \text{CaCO}_3(s) ; K_c = [\text{CO}_2]$$

b) $\text{N}_2 + \text{O}_2$

$$2\text{NO} \leftrightarrow \text{2NOCl} ; K_c = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]}$$
29
Can a mixture of 1.5 mol CH₃, 3 mol O₂, 2 mol CO₂, and 0.5 mol H₂O be maintained in a 5 L flask (K_c = 175)? Explain.

-No. Q < K, so the reaction will shift to the right.

30
Is the following reaction in an 8 L flask at 273 K at equilibrium? Explain.

\[
\text{BaCl}_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \quad \leftrightarrow \quad \text{BaSO}_4(\text{s}) + 2\text{NaCl(aq)}
\]

-No. Q > K, so the reactants are favored.

31
If 0.68 mol CH₃OH, 0.54 mol CO, and 0.94 H₂ are all added to a 5 L container at 500 K (K_c = 150), is the mixture at equilibrium? If not, which way will it shift?

No. Q < K, so it will shift to the right.

33
A mixture of 0.28 mol H₂, 0.28 mol CO₂, 0.22 mol CO, and 0.22 mol H₂O is brought to equilibrium at 562 K in a 3.5 L flask. What are the equilibrium amounts of each? (K_c = 9.84)

\[
\text{CO}_2(\text{g}) + \text{H}_2(\text{g}) \quad \leftrightarrow \quad \text{CO}(\text{g}) + \text{H}_2\text{O(l)}
\]

H₂: 0.05657 mol
CO₂: 0.05657 mol
CO: 0.080 mol
H₂O: n/a
35
Starting with 0.217 mol NO\textsubscript{2} and 0.423 mol O\textsubscript{2} in a 2 L flask at 500 K, how many moles of NO\textsubscript{} will be present at equilibrium? \(K_c = 1.4 \times 10^{-3}\)

\[\text{NO} + \text{O}_2 \rightleftharpoons \text{NO}_2\]
- Moles NO: 0.008

37
How many moles of each substance will there be at equilibrium if 1 mol each of CO\textsubscript{2} and COF\textsubscript{2} are added?

\[2\text{COF}_2(g) \rightleftharpoons \text{CO}_2(g) + \text{CF}_4(g)\] at 300 K in a 4 L flask \(K_c = 3.4 \times 10^{-3}\)

COF\textsubscript{2}: 0.215 mol
CO\textsubscript{2}: 0.215 mol
CF\textsubscript{4}: 0.305 mol

39.
CS\textsubscript{2}(g) + 4H\textsubscript{2}(g)
\[\rightleftharpoons\]
CH\textsubscript{4}(g) + 2H\textsubscript{2}S(g) \(K_c = 3.5\)

a) Which way will the reaction shift?
- To the right

b) How many moles of each will be present at equilibrium?
- 0.45, 0.3, 0.15, respectively

41
If 1.5 mols of PCl\textsubscript{5} dissociate in a 2 L flask, what is the total equilibrium pressure at 375 K? \(K_c = 3.0\)

PCl\textsubscript{5}(g)
Cl\(_2(g)\) + PCl\(_3(aq)\)
- 69.2 atm

43

In the reaction below, what are the concentrations of Cr\(^{3+}\) and Cr\(^{2+}\)?

\[
Pb(s) + 2Cr^{3+} (aq) \rightleftharpoons Pb^{2+} (aq) + 2Cr^{2+} (aq)
\]

a) Cr\(^{3+}\): 2
b) Cr\(^{2+}\): 2

45.

One sketch below represents an initial nonequilibrium mixture in the reversible reaction:

\[
SO_2\; (g)+Cl_2\; (g) \rightleftharpoons SO_2Cl_2\; (g)
\]

with

\[
K_c= 4.0
\]

Which of the other three sketches best represents an equilibrium mixture? Explain.

Solution

represents the best due to it having numbers of SO2Cl2,SO4,Cl2 molecules that are consistent with Kc. (b) has 12 SO2Cl2 molecules per unit volume, 1 Cl2 molecules per unit volume, 3 SO2 molecules per unit volume. They are the requisite number of each type of molecules needed to generate the expected Kc.

53. Problem

1.11 mol each of \(CO\) and \(Cl_2\) are introduced into an evacuated 2.97L flask, and the following equilibrium is established at 771 K.

\[
CO\; (g)+Cl_2\; (g) \rightleftharpoons COCl_2\; (g)
\]

with
Calculate (a) the partial pressure of COCl₂(g) (b) total gas pressure

Solution:

(a) P(CO) = P(Cl₂) = \( \frac{nRT}{V} = 1.11 \text{mol} \times 0.08206 \text{Latm}^{-1}\text{mol}^{-1}\text{K}^{-1} \times 771 \text{K} / 2.97 \text{K} = 23.6 \text{atm} \)

Equation: CO(g) + Cl₂(g) \( \rightleftharpoons \) COCl₂ Kp = 22.5

Initial: 23.6atm 23.6atm 0atm

Changes: -xatm -xatm +xatm

Equil: 23.6-xatm 23.6-xatm xatm

Kp = \( \frac{P(\text{COCl₂})}{P(\text{CO})P(\text{Cl₂})} = 22.5 \)

\( \frac{x}{(23.6-x)^2} = x/(557.0-47.2x+x^2) \)

22.5(557.0-47.2x+x^2) = x = 12532.5 - 1062x + 22.5x^2 = x

12532.5 - 1062x + 22.5x^2 = 0

x = \( \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} \) = \( \frac{-(-1063) \pm \sqrt{(-1063)^2 - 4(22.5)(12532.5)}}{2(22.5)} \) = 69.04 (too large), 22.62

P(CO) = P(Cl₂) = 23.6atm - 22.62atm = .98atm P(COCl₂) = 22.62 atm

(a) P_total = P(CO) + P(Cl₂) + P(COCl₂) = .98atm + .98atm + 22.62atm = 24.58atm

55. Problem

Explain how Le Chatelier's principle supports the fact that the continuous removal of one of the products of a chemical reaction has the effect of causing the reaction to go to completion.

Solution:

Makes Q less than K and reaction shift the right as result and change in pressure of a constant temp equilibrium mixture.

63. Problem

If the volume increases the system, what effect does it have on the equilibrium condition?

a. \( (\text{C(s)} + \text{H}_2\text{O}(g)) \rightleftharpoons \text{CO(g)} + \text{H}_2(g) \)

b. \( (4\text{NH}_3(g) + 5\text{O}_2(g)) \rightleftharpoons 4\text{NO(g)} + 6\text{H}_2\text{O(g)} \)

c. \( (\text{Ca(OH)}_2(s) + \text{CO}_2(g)) \rightleftharpoons \text{CaCO}_3(s) + \text{H}_2\text{O(g)} \)
71. Problem

Why the % molecules dissociate into atoms in reactions of \( \text{I}_2(\text{g}) \rightarrow 2\text{I}(\text{g}) \) always raises with a. raise temp?

Solution:

Increasing temp adds heat and decreasing temp takes away the heat. When the temp raises, heat is absorbed and helps in breaking molecules into atoms. Lower temp, lower the % of dissociation and higher temp, higher % of dissociation.

73. Problem

Refer to Example 14-13. Suppose that 0.1L of the equilibrium mixture is diluted to 0.25L with water. What will be the new concentrations when equilibrium is re-established?

Solution:

\[
\text{Eq: } \text{Ag}^+(\text{aq}) + \text{Fe}^{2+}(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{Ag}(\text{s}) \quad \text{K}_c=2.98
\]

Init: 0.31M 0.21M 0.19M

Ch: 0.12M 0.084M 0.076M

Equil: +xM +xM -xM

New: 0.12+x 0.0084+x 0.076-x

\[
\text{K}_c=\frac{[\text{Fe}^{3+}]}{[\text{Ag}^+][\text{Fe}^{2+}]}=2.98=0.076-x/(0.12+x)(0.084+x)
\]

\[
29.8(0.12+x)(0.084+x)=0.076-x
\]

\[
0.076-x=0.03+0.61x+2.98x^2+2.98x^2+1.16x-0.046=0
\]

\[
x=\frac{-16.1+\sqrt{(2.59+0.55)^2}}{5.96}=0.027,-0.57
\]

\[
[\text{Fe}^{2+}]=0.084-0.57=-0.49\text{M}
\]

\[
[\text{Fe}^{2+}]=0.084+0.027=0.111\text{M} \quad [\text{Ag}^+]=0.12+0.027=0.15\text{M}
\]

\[
[\text{Fe}^{3+}]=0.076-0.027=0.049\text{M}
\]
75. Problem

Started with $\text{(SO}_3\text{(g))}$ at 1.00 atm, what will be the total pressure when equilibrium is reached in the following reaction at 700 K?

$$2\text{SO}_3\text{(g)} \rightleftharpoons 2\text{SO}_2\text{(g)} + \text{O}_2\text{(g)}$$

with

$$K_p = 1.65 \times 10^{-5}$$

Solution:

Pressure: $2\text{SO}_3\text{(g)} \rightarrow 2\text{SO}_2\text{(g)} + \text{O}_2\text{(g)}$

Initi: 1.00 atm 0 0

Change: -2x +2x +x

Equil: 1.00-2x 2x x

$$K_p = (PSO_2)^2(PO_2)/(PSO_3)^2$$

$$1.65 \times 10^{-5} = (2x)^2(x)/(1.00-2x)^2$$

$$x = P_{O_2} = 0.0155 \text{ atm}$$

$$PSO_2 = 2x = 2(0.0155) = 0.0310 \text{ atm}$$

$$PSO_3 = 1.00-2x = [1.00-2(0.0155)] = [1.00-0.0310] = 0.969 \text{ atm}$$

$$P_{tot} = P_{O_2} + P_{SO_2} + P_{SO_3} = (0.0115 + 0.0310 + 0.969) \text{ atm} = 1.016 \text{ atm}$$

77. Problem

Calculate the amount of SO2, O2, SO3 listed on (a) Figure 14-6(c); (b) Figure 14-7(b)

Solution:

(a) $2\text{SO}_3\text{(g)} \rightarrow 2\text{SO}_2\text{(g)} + \text{O}_2\text{(g)}$ $K_c = 1/2.8 \times 10^2$

Init: 1.68 mol 0.32 0.16

C: -2x +2x x

E: 1.68-2x 0.32+2x 0.16+x
3.57*10^-3 = (((0.32 + 2x)^2)/10)*((0.16 + x)/10))/((1.68 - 2x)^2)/10

X = 0.10755 M

[SO3] = 1.68 - 2(0.10755) = 1.46 mol

[SO2] = 0.32 + 2(0.10755) = 0.54 mol

[O2] = 0.16 + 0.10755 = 0.27 mol

(b) 2SO2(g) + O2(g) \rightarrow 2SO3(g) K_c = 2.8*10^2

I: 0.32 mol 0.16 0.68

C: -2x -x +2x

E: 0.32 - 2x 0.16 - x 0.68 + 2x

K_c = [SO3]^2/[SO2]^2[O2]

2.8*10^2 = ((0.68 + 2x)^2)/(0.32 - 2x)^2(0.16 - x)

X = 0.0749 M

[SO3] = 0.68 + 2(0.0749) = 0.83 mol

[SO2] = 0.32 - 2(0.0749) = 0.17 mol

[O2] = 0.16 - 0.0749 = 0.085 mol

79. Problem

Methanation reaction: methane is produced from a mixture of CO and H2.

CO(s) + 3H2(g) \rightarrow CH4(g) + H2O(g) \text{ change} = -230 \text{ kJ} K_c = 190 \text{ at } 1000 K

(a) Is the reaction favored at higher or lower temp? Higher or lower pressure? (b) 4 mol of synthesis gas with a 3:1 ratio of H2(g) to CO(g) in a 15.0 L flask. What is the mole fraction of CH4 at 1000K?

Solution:

(a) Reaction is exothermic. Lowering temp, shifts right and increase conversion of synthesis gas to methane. High pressure, shifts right and increase conversion of synthesis gas to methane

(b) CO(s) + 3H2(g) \rightarrow CH4(g) + H2O(g)

I: 1 mol/15 L 3 mol/15 L 0 0
C: -x -3x +x +x
E: 0.0667 - 0.2 - 3x x x
Kc = \[\text{[CH}_4\text{][H}_2\text{O}]/\text{[CO][H}_2\text{]}\]^3
190 = (x)((x)/(0.0667-x)(0.2-3x))^3
x = 0.0424 M
Mole fraction of CH4 = equilibrium mol of CH4/total number of mol at equilibrium
= 0.0243 mol/(0.0424 + 0.0424 + 0.0243 + 0.0728) = 0.134 mol

84. Problem:
Synthesis of ammonia at 500K, N2+(g)+3H2(g) → 2NH3(g), Kp = 8.71 * 10^-2. N2 and H2 are mixed in a 1:7 ratio and total pressure is 1.00 atm. Find the mole % NH3 at equilibrium.

Solution:
Kp = (PNH3)^2/(PN2)(PH2)^3
= ((XNH3)^2/(XN2)(XH2)^3)*1/(Ptotal)^2
= ((XNH3)^2/1*(7)^2)*1/(1^2)
= 8.71*10^-2
(xNH3)^2/49 = 8.71*10^-2
(xNH3)^2 = 4.2679
xNH3 = 2.07
%Mol of NH3 = mol/total mol*100% = 2.07/1+7+2.07*(100%) = 20.52%

102. Problem
Equilibrium is established in the reaction 2SO2(g)+O2(g) → 2SO3(g) at a Temp where Kc = 123. If the number of mol of SO3(g) in the equilibrium mixture is the same as the number of mol of SO2(g) (a) the number of mol of O2(g) is also equal
to the number of mol of SO\(_2\) (b) the number of mol of O\(_2\) is half the number of mol of SO\(_2\) (c) [O\(_2\)] may have any of several values (d) [O\(_2\)] = 0.008M

Solution:

\[
K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}
\]

\[= \frac{(x/V)^2}{(x/V)^2[O_2]}\]

\[K_c = \frac{1}{[O_2]}\]

\[123 = \frac{1}{[O_2]}\]

\[[O_2] = 0.008 M\]

D is the correct answer

104. Problem

\(2NO_2(g) \rightarrow 2NO(g) + O_2(g)\), \(K_c = 3.19 \times 10^{-12}\) at 184°C. At 184°C, the value of \(K_c\) for this reaction is (a) \(0.9 \times 10^6\) (b) \(7.5 \times 10^2\) (c) \(5.6 \times 10^5\) (d) \(2.8 \times 10^5\).

Solution:

\[
K'_c = \frac{[NO_2]}{[NO][O_2]^{1/2}}
\]

\[= \frac{1}{(K_c)^{1/2}}\]

\[= \frac{1}{(3.19 \times 10^{-12})^{1/2}}\]

\[= 5.6 \times 10^5\] (c) is correct

Q14.1

based on these descriptions, write a balanced equation for each reversible reaction.

a. Oxygen gas oxidizes gaseous ammonia vapor to gaseous nitrogen and water vapor.

b. Hydrogen gas reduces gaseous nitrogen dioxide to gaseous ammonia and water vapor.

c. Liquid acetone is in equilibrium with its vapor.

S31

a. \(3 O_2(g) + 4 NH_3(g) \leftrightarrow 2 N_2(g) + 6 H_2O(g)\)

b. \(7 H_2(g) + 2 NO_2(g) \leftrightarrow 2 NH_3(g) + 4 H_2O(g)\)
c. \((\text{CH}_3\text{C})_2\text{O}(l) \leftrightarrow (\text{CH}_3\text{C})_2\text{O}(g)\)

Q13

For the following reactions, determine the equilibrium constant expression in terms of concentration.

a. a) \(\text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g)\)
   b. b)\(\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)\)
   c. c)\(\text{C}(s) + \text{H}_2\text{O}(g) \rightleftharpoons \text{H}_2(g) + \text{CO}(g)\)

S13

a)\(K_c = [\text{NO}_2]^2/[\text{N}_2\text{O}_4]\)
   b)\(K_C = [\text{NH}_4^+][\text{OH}^-]/[\text{NH}_3]\)
   c)\(K_C = [\text{H}_2][\text{CO}]/[\text{H}_2\text{O}]\)

Q25

At a certain temperature 1L flask with .298mol \(\text{PCI}_3(g)\) and \(8.70\times10^{-3}\text{mol}\)\(\text{PCI}_5(g)\). After the system reached equilibrium, \(2\times10^{-3}\text{mol}\)\(\text{Cl}_2(g)\) was found in the flask.

Calculate equilibrium concentrations and the value of \(K_c\) and \(K_p\) assuming 298K.

\(\text{PCI}_5 \rightleftharpoons \text{PCI}_3 + \text{Cl}_2\)

\[8.7\times10^{-3}\text{mol} \times 0.298 = 0\]

\[-x + x + x = 2\times10^{-3}\]

\[8.7\times10^{-3}\text{mol} - x \times 0.298 + x = 2\times10^{-3}\]

\[K_c = (0.002)(0.3)/0.0067 = 0.0896\]

\[K_p = K_c(\text{RT})\Delta n\]

\[K_p = (0.0821)(298)2-1\]

\[K_p = 2.19\]

Q26

At 2000 K, \(K_c = 0.154\) for the reaction \(\text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g).\)
a. Is the conversion of N2 (g) to 2NH3(g) favored at high or low pressures?

b. If the equilibrium mixture is transferred from a 1.00 L flask to a 2.00L flask, will the number of moles of N2(g) increase, decrease, or remain unchanged?

c. If H2 is added, which direction would the reaction shift?

S26

a. High because the reactants have more moles than the products

b. Increase

c. Right

Q29

In the following reaction, which direction will the reaction favor if the initial concentration of O2 and CO2 is 4 M, and Kc is 1.16 at 298 K?

\[4\text{KO}_2(s) + 2\text{CO}_2(g) \rightleftharpoons 2\text{KCO}_3(s) + 3\text{O}_2(g)\]

S29

\[Q = \frac{[\text{O}_2]^3}{[\text{CO}_2]^2}\]

\[Q = \frac{43}{42} = 4K < Q = \text{reaction will shift left}\]

Q49

At 338 K, pure PCl5 gas is present in a flask at a pressure of 26.7 kPa. At 473 K this is completely dissociated into PCl3 gas and Cl2 gas. Calculate the pressure in the flask at 473 K. No volume is specified, so take a convenient value, say V = 1.00 L. The equation for the reaction that occurs at 473 K is

S49

\[\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)\]

At 338 K, moles of PCl5 gas = \(\frac{PV}{RT}\)

\[= \frac{(26.7 \times 1.00)}{(8.314 \times 338)} = 9.50 \times 10^{-3} \text{ mol.}\]

At 473 K, moles of gas present = \(2 \times 9.50 \times 10^{-3} = 0.0190 \text{ mol.}\)

Then, using the Ideal Gas Equation, the new pressure can be deduced.

\[P = \frac{nRT}{V}\]

\[= 0.0190 \times 8.314 \times 473/1.00 = 74.7 \text{ kPa}\]
Q57

Predict the effects on the amount of $NH_3$ in the following reaction

\[ N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)} \]

with $\Delta H = -46.11$ kcal

when

a. temperature is lowered
b. introducing more H2
c. doubling the vol. of the container.

S57

a. amount will increase because this is an exothermic reaction
b. amount will increase to obtain equilibrium
c. amount will decrease to produce more moles of gas to fill the larger container

Q59

If the system is heated which direction will the equilibrium shift?

a. $2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$ $\Delta H^\circ = 566$ kJ
b. $C_2H_4(g) + H_2(g) \rightleftharpoons C_2H_6(g)$

S59

a. shifts to the right
b. shifts to the left

Q61

If the volume of an equilibrium mixture of N2(g), H2 (g), and NH3 (g) is reduced by doubling the pressure, will PNH3 have increased, decreased, or remained the same when equilibrium is re established? Explain.

\[ N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)} \]

Increase because the number of moles on the product side is less than the amount of moles on the reactant side, therefore, if there is an increase in pressure, the reaction will move toward the side with less moles.
Q63

If we increased the pressure for the following reactions, what would the effect on equilibrium be?

a. \(2A (g) + B (s) \rightleftharpoons 4C (g)\)

b. \(5A (g) + 2B(l) \rightleftharpoons C(s) + 4D(g)\)

c. \(A(l) + 4B(g) + 2C (g) \rightleftharpoons 3D(s) + 5E(g)\)

S63

Pressure is the inverse of volume, so when we increase pressure, we are decreasing volume. A will result in a shift left, B will shift right, and C will shift left.