Nuggets: Equilibrium Calculations; Reaction Quotient (Q)

EQUILIBRIUM CALCULATIONS (these are typical problems)

1. Given concentrations or moles at equilibrium $\rightarrow$ find $K_{eq}$
2. Given concentrations or moles not at equilibrium $\rightarrow$ find $K_{eq}$ (usually requires an ICE Table)
3. Given $K_{eq} \rightarrow$ find equilibrium concentrations (usually requires an ICE Table; “x” values in ICE Table)
   
   If algebraic expression is difficult to solve easily, try:
   
   a. taking square root of both sides; if not possible try
   
   b. the quadratic equation
   
   c. make an approximation (presented in the next unit with acids and bases)

ICE Table: A table with I = Initial concentrations/pressures, C = Change (C row is always stoichiometric), and E = Equilibrium concentrations/pressures for a reaction

Example 1: (problem type #1 above: given concentration at equilibrium $\rightarrow$ find $K_{eq}$)
If a 2.0L flask at 25°C contains 0.0860g H₂, 0.963g I₂, and 12.2g HI at equilibrium, what is $K_{eq}$?

$\text{H}_2(g) + \text{I}_2(g) \rightarrow 2\text{HI}(g)$

Answer 1: $K_{eq} = 56.2$

1. Determine concentrations (M): $[\text{HI}] = \frac{(12.2 \text{ g HI})(1 \text{ mol HI} / 127.9 \text{ g HI})}{2.0 \text{ L}} = 0.0477 \text{ M}$; $[\text{H}_2] = \frac{(0.0860 \text{ g H}_2)(1 \text{ mol H}_2 / 2.016 \text{ g H}_2)}{2.0 \text{ L}} = 0.0213 \text{ M}$; $[\text{I}_2] = \frac{(0.963 \text{ g I}_2)(1 \text{ mol I}_2 / 253.8 \text{ g I}_2)}{2.0 \text{ L}} = 0.00190 \text{ M}$;

2. Write $K_{eq}$ expression: $K_{eq} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$;

3. Plug in equilibrium concentrations: $K_{eq} = \frac{(0.0477)^2}{(0.0213)(0.00190)} = 56.2$

Example 2: (problem type #2 above: given concentrations not at equilibrium $\rightarrow$ find $K_{eq}$)
A 5.0L container initially has 0.50mol NO(g) and 0.30mol O₂(g). Once equilibrium has been established, 0.10mol NO₂(g) exist. What is $K_{eq}$?

$2\text{NO(g)} + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)$

Answer 2: $K_{eq} = 1.25$

Step 1. Set up ICE Table; units should be M.
If no quantity is given then initial amount is assumed to be 0;
initial mol NO₂ = 0 and therefore $[\text{NO}_2]_{\text{initial}} = 0$ M
$[\text{NO}]_{\text{initial}} = 0.50 \text{ mol} / 5.0 \text{ L} = 0.10 \text{ M}$
$[\text{O}_2]_{\text{initial}} = 0.30 \text{ mol} / 5.0 \text{ L} = 0.06 \text{ M}$
$[\text{NO}_2]_{\text{final}} = 0.10 \text{ mol} / 5.0 \text{ L} = 0.02 \text{ M}$

<table>
<thead>
<tr>
<th></th>
<th>2NO(g)</th>
<th>O₂(g)</th>
<th>2NO₂(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>0.10</td>
<td>0.06</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>-0.02</td>
<td>-0.01</td>
<td>+0.02</td>
</tr>
<tr>
<td>E</td>
<td>0.08</td>
<td>0.05</td>
<td>0.02</td>
</tr>
</tbody>
</table>

Step 2. Complete ICE Table by filling in missing parts.

C row is always stoichiometric: ratio of 2 : 1 : 2 in this example.

a. Write C row NO₂ value (use I and E row NO₂ values to determine the C row value of NO₂).
b. Write C row NO and O₂ values based on C row NO₂ value and the stoichiometric ratio of 2 : 1 : 2.
c. Write E row NO and O₂ values based on I and C row NO and O₂ values, respectively.

Step 3.

a. Write $K_{eq}$ expression
b. Plug in E row molarity values

$K_{eq} = \frac{[\text{NO}_2]^2}{[\text{NO}]^2[\text{O}_2]} = \frac{(0.02)^2}{(0.08)^2(0.05)} = 1.25$
**Example 3:** (problem type #3 above: given \( K_{\text{eq}} \rightarrow \) find concentration)

If the initial concentrations of \( \text{H}_2(g) = 1_2(g) = 0.20\text{M} \), and the \( K_{\text{eq}} \) for the reaction below at 35°C is 144, what are the **final concentrations** of all chemicals?

\[
\text{H}_2(g) + 1_2(g) \rightleftharpoons 2\text{H}(g)
\]

**Answer 3:** \( [\text{H}_2]_{\text{eq}} = [1_2]_{\text{eq}} = 0.03\text{M}; \) \[\text{[HI]}_{\text{eq}} = 0.34\text{M}\]

**Step 1.** Set up ICE Table using units of M.  
If no quantity is given then initial is assumed to be 0; therefore initial concentration

\[
[\text{HI}]_{\text{initial}} = 0\text{M and } [1_2]_{\text{initial}} = 0.20\text{M}
\]

\[
\begin{array}{ccc}
\text{H}_2(g) & + & 1_2(g) \\ \text{I} & 0.20 & 0.20 & 0 \\
\text{C} & \times & \times & \times \\
\text{E} & 0.20-x & 0.20-x & 2x
\end{array}
\]

**Step 2.** Complete ICE Table.  
C row is **always stoichiometric**: ratio of \( 1 : 1 : 2 \) in this example.

a. Write C row unknown values in correct stoichiometric ratios.

b. Write E row values based on I and C row values.

\[
\begin{array}{ccc}
\text{H}_2(g) & + & 1_2(g) \\ \text{I} & 0.20 & 0.20 & 0 \\
\text{C} & -x & -x & +2x \\
\text{E} & 0.20-x & 0.20-x & 2x
\end{array}
\]

**Step 3.** a. Write \( K_{\text{eq}} \) expression.

b. Plug in K value and E row unknown values.

c. Solve for \( x \).  
In this case take square root of both sides.

d. Check to see if \( "x" \) answers the question asked.

\[
K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][1_2]} \; \text{; \; 144} = \frac{(2x)^2}{(0.20-x)(0.20-x)} = \frac{(2x)^2}{(0.20-x)^2}
\]

\[
\sqrt{144} = \frac{(2x)^2}{(0.20-x)^2} \; \text{; \; 12} = \frac{2x}{0.20-x} \; \text{; \; 2.4-12x = 2x; 14x = 2.4; } x = 0.17
\]

**Step 4.** The value of “\( x \)” does **not answer the question!**  
To complete the question: \( [\text{H}_2]_{\text{eq}} = [1_2]_{\text{eq}} = 0.20-x = 0.20-0.17 = 0.03\text{M}; \)

\[\text{[HI]}_{\text{eq}} = 2x = 2(0.17) = 0.34\text{M}\]

**REACTION QUOTIENT \( Q \) –** a snapshot in time of the reaction progress; \( Q \) expression is the same as a \( K_{\text{eq}} \) expression but \( Q \) may or may not be a system at equilibrium.  
\( Q \) **can be used to determine if system is at equilibrium, and if not, which way the reaction will shift to reach equilibrium**

\[
aA(g) + b\text{B}(g) \rightleftharpoons c\text{C}(g) \rightarrow Q = \frac{[\text{C}]^c}{[\text{A}]^a[\text{B}]^b}
\]

Q > \( K \) → system will shift left (to reactant side) to reach equilibrium  
(when reaction shifts L \( \Rightarrow \) reactants ↑ and products ↓ \( \Rightarrow \) Q ↓ and approaches K)

Q < \( K \) → system will shift right (to product side) to reach equilibrium  
(when reaction shifts R \( \Rightarrow \) reactants ↓ and products ↑ \( \Rightarrow \) Q ↑ and approaches K)

Q = \( K \) → at equilibrium (no shift)

**Example 4:** If \( K \) of \( \text{PCl}_2 \) is \( 7.2 \times 10^{-4} \) for the equilibrium below and the concentrations of \( \text{PCl}_3 \) is 0.50M, \( \text{PCl}_1 \) is 0.15M, and \( \text{Cl}_2 \) is 0.025M, is the system at equilibrium?  
If not, would this reaction proceed to the right or left to reach equilibrium?

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

**Answer 4:** Determine Q.

\[
Q = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_2]} = \frac{(0.15)(0.025)}{(0.50)} = 7.5 \times 10^{-3}
\]

Since \( Q \neq K \) the reaction is **not** at equilibrium.

Since Q > K, the reaction will shift **left**.

**Quadratic equation:** for \( ax^2 + bx + c = 0 \rightarrow x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} \)

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**Types of Equilibria**

<table>
<thead>
<tr>
<th>Types of Equilibria</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid ( \rightleftharpoons ) Liquid</td>
<td>ice mixed with water; ( \text{H}_2\text{O}(s) \rightleftharpoons \text{H}_2\text{O}(l) )</td>
</tr>
<tr>
<td>Liquid ( \rightleftharpoons ) Gas</td>
<td>water with water gas above it; ( \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{O}(g) )</td>
</tr>
<tr>
<td>Solid (condensed phase)</td>
<td>under normal pressures ( \text{Fe}(s) ) is found in 3 forms: ( \delta-, \gamma-, ) and ( \alpha- )-allotropes (BCC = body-centered cubic structure; FCC = face-centered cubic structure)</td>
</tr>
<tr>
<td>Solubility</td>
<td>glucose, ( \text{C}_6\text{H}_12\text{O}_6(s) ), dissolving in water; ( \text{C}_6\text{H}_12\text{O}_6(s) + \text{H}_2\text{O}(l) \rightleftharpoons \text{C}_6\text{H}_12\text{O}_6(aq) )</td>
</tr>
<tr>
<td>Partition</td>
<td>a substance such as ( 1_2 ) (nonpolar) dissolved in a layered-mixture of water (polar) and hexane (nonpolar); the ( 1_2 ) dissolves more readily into the hexane liquid (nonpolar likes nonpolar) and is partitioned to that layer; ( 1_2 ) (in ( \text{H}_2\text{O}(l) )) \rightleftharpoons \text{I}_2 ) (in ( \text{C}_6\text{H}_14(l) ))</td>
</tr>
</tbody>
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**Cooling Temperature**

- \( \delta- \)-iron
- \( \gamma- \)-iron
- \( \alpha- \)-iron

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**Solid**

- Solid (condensed phase)

**Liquid**

- Liquid
PHASE DIAGRAMS: A diagram of Pressure (y-axis) versus Temperature (x-axis) that shows what phase (solid, liquid or gas) is present at a particular T and P.

**Triple Point**: All 3 phases at equilibrium at this T and P.

**Along the lines**: Equilibrium of 2 phases (solid+liquid; liquid+gas; or solid+gas)

1. Given the equilibrium concentrations for the reaction below, \([\text{PCl}_5] = 0.0350\text{M, \[\text{PCl}_3] = 0.0101\text{M, \[Cl}_2] = 0.00250\text{M, what is the } K_c \text{ value?}\)

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

2. If the reaction shown below had an equilibrium constant of \(K_c = 6.0 \times 10^5\)

\[
\text{CH}_3\text{COCH}_3(\text{g}) + 2\text{H}_2(\text{g}) \rightleftharpoons \text{H}_2\text{CO}(\text{g}) + 2\text{CH}_4(\text{g})
\]

which of the following statements would be correct? If none are correct, select answer “e”.

a. The concentration of \(\text{CH}_3\text{COCH}_3\) will be the same as the concentration of \(\text{H}_2\text{CO}\).

b. The concentration of \(\text{H}_2\) will be twice the concentration of \(\text{CH}_3\text{COCH}_3\).

c. The concentration of \(\text{H}_2\text{CO}\) will be the same as the concentration of \(\text{CH}_3\text{COCH}_3\).

d. The value of \([\text{H}_2\text{CO}] \times [\text{CH}_4]^2\) will be much greater than the value of \([\text{CH}_3\text{COCH}_3] \times [\text{H}_2]^2\).

e. None of these statements are correct.

3. A mixture at equilibrium in a 0.45L flask contains 2.8g \(\text{N}_2(\text{g})\), 0.063g \(\text{H}_2(\text{g})\), and 0.85g \(\text{NH}_3(\text{g})\). What is the value of \(K_c\) for this reaction?

\[
\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})
\]

4. Given the reaction at equilibrium, if \(K_c = 25\), \([\text{H}_2] = 2.5\text{M, and } [\text{HI}] = 15.0\text{M, what is } [\text{I}_2]?\)

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

5. Initially, a 250. ml flask contains 0.063mol \(\text{COBr}_2(\text{g})\) and 0.250mol \(\text{CO}(\text{g})\). When the reaction reached equilibrium, the number of moles of \(\text{CO}(\text{g})\) is 0.260.

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

a. Calculate the equilibrium concentrations of \(\text{COBr}_2(\text{g}), \text{CO}(\text{g}), \text{and Br}_2(\text{g})\).  

b. Calculate \(K_c\).
6. Initially, a flask contains 1.0M N\(_2\)(g) and 1.0M O\(_2\)(g). When the reaction has reached equilibrium, the final concentration of N\(_2\)(g) is 0.9M.  
\[ \text{N}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}(g) \]
a. Calculate the final concentrations of O\(_2\)(g) and NO(g).  
b. Calculate K\(_c\).

7. The equilibrium constant for the reaction below is K\(_c\) = 0.0150, what are the final concentrations of all chemicals if the initial concentration of HI(g) was 1.50M and the initial concentrations of H\(_2\)(g) and I\(_2\)(g) were 0.00500M?  
\[ 2\text{HI}(g) \rightleftharpoons \text{H}_2(g) + \text{I}_2(g) \]

8. The equilibrium constant for the reaction below is K\(_p\) = 4.50 \times 10^{-5}, what are the final pressures of all chemicals if the initial pressures of O\(_2\)(g) and N\(_2\)(g) are 0.750atm?  
\[ \text{O}_2(g) + \text{N}_2(g) \rightleftharpoons 2\text{NO}(g) \]

9. Initially, 1.50moles of A(g) are added to a 3.00L reaction vessel and the system is allowed to come to equilibrium. At equilibrium, the amount of A(g) has decreased by 35.0%.  
\[ 2\text{A}(g) \rightleftharpoons 2\text{B}(g) + \text{C}(g) \]
a. What are the concentrations of A(g), B(g), and C(g) at equilibrium?  
b. What is the value of K\(_c\)?

10. In the reaction, A(g) \rightleftharpoons B(g), with K\(_p\) = 5.00, if the initial partial pressures of A(g) = 2.50atm and B(g) = 10.00atm, what are the final pressures of A(g) and B(g)?  
a. A = 1.00atm, B = 5.00atm  
b. A = 4.58atm, B = 7.91atm  
c. A = 5.00atm, B = 10.00atm  
d. A = 2.08atm, B = 10.42atm  
e. A = 5.00atm, B = 7.50atm

11. K\(_c\) is 0.75 for the reaction A(g) \rightleftharpoons B(g) + C(g). If a 2.0 L flask contains 5.0mol A(g), how many mol of A(g), B(g), and C(g) will exist once equilibrium is achieved? (Hint: Requires the quadratic equation.)

12. What are the equilibrium concentrations of NO\(_2\)(g) and N\(_2\)O\(_4\)(g) if initially there was 0.0015mol of N\(_2\)O\(_4\)(g) in a 5.0L flask, and the K\(_c\) at 25 °C is 0.125? (Hint: Requires the quadratic equation.)  
\[ \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \]

13. If [N\(_2\)(g)] = 2.5M, [O\(_2\)(g)] = 7.0M, [NO(g)] = 4.5M, and K\(_c\) = 65 which statement is correct?  
\[ \text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g) \]
a. The reaction will not shift since the system is at equilibrium.  
b. The reaction will shift to produce N\(_2\)(g) and O\(_2\)(g) and thereby increase the value of Q.  
c. The reaction will shift to produce N\(_2\)(g) and O\(_2\)(g) and thereby decrease the value of Q.  
d. The reaction will shift to produce NO(g) and thereby increase the value of Q.  
e. The reaction will shift to produce NO(g) and thereby decrease the value of Q.
14. For each of the following questions, determine if the system is at equilibrium. If the system is not in equilibrium, which way will the reaction shift to reach equilibrium (left or right)?

I. For the following reaction, $K_c = 10.5$.
\[ \text{CO}(g) + 2\text{H}_2(g) \rightleftharpoons \text{CH}_3\text{OH}(g) \]
The starting concentrations are: $[\text{CO}] = 0.00135 \text{ M}$, $[\text{H}_2] = 0.00226 \text{ M}$, $[\text{CH}_3\text{OH}] = 1.33 \text{ M}$

II. For the following reaction, $K_c = 4.8 \times 10^3$.
\[ \text{CO}(g) + 3\text{H}_2(g) \rightleftharpoons \text{CH}_4(g) + \text{H}_2\text{O}(g) \]
The starting concentrations are: $[\text{CO}] = 0.020 \text{ M}$, $[\text{H}_2] = 0.60 \text{ M}$, $[\text{CH}_4] = 1.2 \text{ M}$, and $[\text{H}_2\text{O}] = 0.30 \text{ M}$.

III. For the following reaction, $K_c = 8.0$.
\[ 2\text{NO}_2(g) \rightleftharpoons \text{N}_2\text{O}_4(g) \]
The starting concentrations of $[\text{NO}_2]$ and $[\text{N}_2\text{O}_4]$ are 0.50M.

15. (This is the same problem as Question #6 from Help Sheet #9 but it has additional parts added.)
a. For the reaction, $A_2(g) + B_2(g) \rightleftharpoons 2\text{AB}(g)$, it was found that at equilibrium $[A_2] = 1.0 \text{ M}$, $[B_2] = 2.0 \text{ M}$, and $[\text{AB}] = 4.0 \text{ M}$. What is the value of $K_c$?
b. While at equilibrium, the concentration of $A_2(g)$ was quickly changed from 1.0M to 2.0M, and all the concentrations changed and the system returned to equilibrium. What are the final concentrations for $A_2(g)$, $B_2(g)$, and $\text{AB}(g)$?
c. Plug in the new concentrations for $A_2(g)$, $B_2(g)$, and $\text{AB}(g)$ determined in part “a” to show that the equilibrium value $K$ has returned to the prior value of 8.0.

ANSWERS

1. $K_c = 7.21 \times 10^{-4} \quad \{ K_c = \frac{[\text{PCL}_3][\text{Cl}_2]}{[\text{PCL}_5]} \} \quad K_c = \frac{(0.0101)(0.00250)}{(0.0350)} = 0.0007214$

2. $d \quad \{ K_c = \frac{[\text{H}_2\text{CO}][\text{CH}_4]^2}{[\text{CH}_3\text{COCH}_3][\text{H}_2]^3} \}$ and since $K > 1$ products are favored; hence, $[\text{H}_2\text{CO}] \times [\text{CH}_4]^2$ must be greater than $[\text{CH}_3\text{COCH}_3] \times [\text{H}_2]^2$

3. $K_c = 170 \quad \{ K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} \}$; for $[\text{N}_2]$: \( \left(2.8 \text{ g N}_2\right) \left(\frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2}\right) = 0.010 \text{ mol N}_2; \left(\frac{0.010 \text{ mol N}_2}{0.45 \text{ L}}\right) = 0.222 \text{ M N}_2; \)
for $[\text{H}_2]$: \( \left(0.063 \text{ g H}_2\right) \left(\frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2}\right) = 0.0313 \text{ mol H}_2; \left(\frac{0.0313 \text{ mol H}_2}{0.45 \text{ L}}\right) = 0.0694 \text{ M H}_2; \)
for $[\text{NH}_3]$: \( \left(0.85 \text{ g NH}_3\right) \left(\frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3}\right) = 0.0500 \text{ mol NH}_3; \left(\frac{0.0500 \text{ mol H}_2}{0.45 \text{ L}}\right) = 0.111 \text{ M NH}_3; \) $K_c = \frac{(0.111)^2}{(0.222)(0.0694)^3} = 166$

4. $[I_2] = 3.6 \text{ M} \quad \{ \text{write K expression: } K_c = \frac{[\text{HI}]^2}{[\text{H}_2][I_2]} \}; \text{ plug in the values: } 25 = \frac{(15)^2}{(2.5)[I_2]}; \text{ solve for } [I_2]; \ [I_2] = \frac{(15)^2}{(2.5)(25)} = 3.6 \text{ M} \}$
5. a. \([\text{COBr}_2] = 0.212\text{M} ; \ [\text{CO}] = 1.040\text{M} ; \ [\text{Br}_2] = 0.040\text{M}\)
\[\text{initial } [\text{Br}_2] = 0 \text{ since none was stated to exist initially; convert moles to M:} \]
\[
\frac{0.063\text{mol COBr}_2}{0.250\text{L}} = 0.252\text{M COBr}_2 ; \quad \frac{0.250\text{mol CO}}{0.250\text{L}} = 1.000\text{M CO} ; \ [\text{Br}_2] = 0\text{M} ;
\]
\[\text{at equilibrium: } \frac{0.260\text{mol CO}}{0.250\text{L}} = 1.040\text{M CO} \text{; write ICE table and add values for initial row} \]
\[\text{for all 3 chemicals and equilibrium row for CO (non-bold values given in problem; bold values added after). Using initial and} \]
\[\text{equilibrium row for CO determine change row for CO: } 1.000 + x = 1.040 ; x = +0.040 \text{; using the reaction stoichiometry,} \]
\[1 : 1 : 1, \text{ add in values for change row for COBr}_2 \text{ and Br}_2 ; \text{finish ICE Table equilibrium row}\]

b. \(K = 0.196 \quad \{K_c = \left[\frac{\text{CO}}{[\text{Br}_2]}\right] ; \ K_c = \left(\frac{0.140(0.040)}{0.212}\right) = 0.196\} \]

6. a. \([\text{O}_2] = 0.9\text{M} ; \ [\text{NO}] = 0.2\text{M} \quad \{\text{initial } [\text{NO}] = 0 \text{ since none was stated to exist initially;} \]
\[\text{write ICE table and add values for initial row for all 3 chemicals and equilibrium row for N}_2 \]
\[\text{(non-bold values given in problem; bold values added after). Using initial and equilibrium row} \]
\[\text{for N}_2 \text{ determine change row for N}_2 : 1 + x = 0.9 ; x = -0.1 \text{; using the reaction stoichiometry,} \]
\[1 : 1 : 2, \text{ add in values for change row for O}_2 \text{ and NO; finish ICE Table equilibrium row}\]

b. 0.05 \( \{K_c = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} ; \ K_c = \frac{(0.2)^2}{(0.9)(0.9)} = 0.0494\} \]

7. \([\text{HI}] = 1.21\text{M} ; \ [\text{H}_2] = [\text{I}_2] = 0.149\text{M} \quad \{\text{write ICE Table (non-bold values given in} \]
\[\text{problem; bold values added after)} ; \ K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} ; \ 0.0150 = \frac{(0.00500 + x)(0.00500 + x)}{(1.50 - 2x)^2} ;
\[
0.0150 = \frac{(0.00500 + x)}{(1.50 - 2x)^2} ; \text{take square root of both sides: } \sqrt{0.0150} = \sqrt{\frac{(0.00500 + x)^2}{(1.50 - 2x)^2}} ;
\]
\[
0.1225 = \frac{(0.00500 + x)}{(1.50 - 2x)^2} ; \text{cross multiply: } 0.1838 - 0.245x = 0.00500 + x; \text{ simplify: } 1.245x = 0.1788; \ x = 0.1436; \]
\[\text{[HI]} = 1.50 - 2x = 1.50 - 2(0.1436) = 1.213\text{M} ; \ [\text{H}_2] = [\text{I}_2] = 0.00500 + x = 0.00500 + 0.1436 = 0.1486\text{M}\]
9. a. \([A] = 0.325\text{M}; [B] = 0.175\text{M}, [C] = 0.0875\text{M}\) \{write ICE Table (non-bold values given in problem; bold values added after); from the information provided that A decreases by 35.0%, then \(2x = (0.35)(1.50) = 0.525\text{mol}; at equilibrium (E Row):\)
\[A = 1.50 - 0.525 = 0.975\text{mol}; B = 2x = 0.525\text{mol}; C = x = 0.525/2 = 0.2625\text{mol};\]
\([A] = \frac{0.975\text{mol}}{3.00\text{L}} = 0.325\text{M}; [B] = \frac{0.525\text{mol}}{3.00\text{L}} = 0.175\text{M}; [C] = \frac{0.2625\text{mol}}{3.00\text{L}} = 0.0875\text{M}\}

b. \(K_c = 2.54 \times 10^{-2}\) \{\(K_c = \frac{[B]^2[C]}{[A]^2}\); \(K_c = \frac{(0.175)^2(0.0875)}{(0.325)^2} = 0.02537\}\n
d. \(\text{initial } P_A = 2.50 \text{ and } P_B = 10.00; \text{ write } K_P; \text{ } K_p = \frac{P_B}{P_A}; \text{ write ICE table; }\)
and enter the equilibrium row values into the \(K_P\) expression: \(5.00 = \frac{10.00 + x}{2.50 - x}\); cross multiply: \(12.50 - x = 10.00 + x\); collect terms: \(6x = 2.5\); solve for \(x\): \(x = 0.42\; \text{atm}\); \(P_A = 2.50 - x = 2.50 - 0.42 = 2.08\; \text{atm};\)
\(P_B = 10.00 + x = 10.00 + 0.42 = 10.42\; \text{atm};\)
It is also possible to solve this by trial and error by looking at the answers; for choices “b” and “e” if you evaluate the \(K_P\) it does not equal 5.00 (“b”): \(K_P = \frac{P_B}{P_A} = \frac{7.91}{4.58} = 1.73\) and “e”: \(K_P = \frac{P_B}{P_A} = \frac{7.50}{5.00} = 1.50\); the other choices, “a”, “c”, and “d”, the \(K_P\) does equal 5.00; however, in choice “a” both reactants and products went down from their initial pressures (reactants and products cannot both go down, one goes up and one goes down); in choice “c”, the product did not change but the reactant increased (again, one must go up and one must go down); in choice “d” the reactant went down the same amount that the product went up which makes sense: 1) one goes up and the other goes down, and 2) the stoichiometry of the reaction (1 : 1 ratio) was present when the reactant lost 0.42 atm and the product gained 0.42 atm; finally the \(K_P\) of “d” equals 5.00 (\(K_P = \frac{P_B}{P_A} = \frac{10.42}{2.08} = 5.01\))

11. 2.9 mol A; 2.1 mol B; 2.1 mol C \{[A] = 5.0 mol/2.0 L = 2.5 M; write ICE Table (non-bold values given in problem; bold values added after); \(K_c = \frac{[B][C]}{[A]} = 0.75 = \frac{(x)(x)}{(2.5 - x)}\);
can’t solve easily; can’t take square root \(\rightarrow\) use quadratic equation;
cross multiply: \(x^2 = -0.75x + (2.5)(0.75)\); collect terms: \(x^2 = -0.75x + 1.875\);
x\(^2\) + 0.75x - 1.875 = 0; solve for \(x\) using quadratic equation; \(a = 1; b = 0.75; c = -1.875;\)
\[
x = \frac{-0.75 \pm \sqrt{(0.75)^2 - 4(1)(-1.875)}}{2(1)} = \frac{-0.75 \pm \sqrt{0.5625 + 7.5}}{2} = \frac{-0.75 \pm \sqrt{8.0625}}{2} = \frac{-0.75 \pm 2.839}{2}; x = \frac{-0.75 \pm 2.839}{2}; x = \frac{2.089}{2}; x = 1.045; [A] = 2.5-x = 1.455M; mol = M x L = (1.455)(2.0) = 2.91 mol A; [B] = [C] = x = 1.045M; mol = M x L = (1.045)(2) = 2.09 mol B and 2.09 mol C\}

12. \([N_2O_4] = 5.94 \times 10^{-4} \text{M}; [N_2O_4] = 2.83 \times 10^{-6} \text{M}\)
\{\([N_2O_4] = \frac{0.0015\text{mol}}{5.0\text{L}} = 0.00030\text{M}; write ICE Table (non-bold values given in problem; bold values added after); \(K_c = \frac{[N_2O_2]^2}{[N_2O_4]^2}; 0.125 = \frac{(2x)^2}{(0.00030 - x)}\); can’t solve easily and cannot take square root; use quadratic equation; cross multiply: \(-0.125x + 0.000375 = 4x^2;\)
Collect terms: \(4x^2 + 0.125x - 0.000375 = 0;\) solve for \(x\) using the quadratic equation; \(a = 4; b = 0.125; c = -0.000375;\)
\[
x = \frac{-0.125 \pm \sqrt{(0.125)^2 - 4(4)(-0.000375)}}{8} = \frac{-0.125 \pm \sqrt{0.015625 + 0.00006}}{8} = \frac{-0.125 \pm \sqrt{0.015625}}{8}; x = \frac{-0.125 \pm 0.12738}{8}; x = \frac{0.002377}{8}; x = 0.0002972M; [N_2O_4] = 0.00030 - x = 0.00030 - 0.0002972 = 2.83 \times 10^{-6} \text{M};\)
\([N_2O_2] = 2x = 2(0.0002972) = 0.000594M = 5.94 \times 10^{-4}\text{M}\)
13. d  
{first determine the value of Q;  \( \text{Q} = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} \);  \( \text{Q} = \frac{(4.5)^2}{(2.5)(7.0)} = 1.16 \);  \( Q < K_c \) \{reaction will shift right and increase Q\}

14. I.  \( Q > K \); not at equilibrium; shifts left  
\[ Q = \frac{[\text{CH}_3\text{OH}]}{[\text{CO}][\text{H}_2]^2} \];  
\( Q = \frac{(1.33)}{(0.00135)(0.00226)^2} = 1.93 \times 10^8 \);  
\( Q > K_c \) \{1.93 \times 10^8 > 10.5 \} \{shifts left\}

II.  \( Q < K \); not at equilibrium; shifts right  
\[ Q = \frac{[\text{CH}_4][\text{H}_2\text{O}]}{[\text{CO}][\text{H}_2]^3} \];  
\( Q = \frac{(1.2)(0.30)}{(0.020)(0.60)^3} = 83.3 \);

\( Q < K_c \) \{83.3 > 4.8 \times 10^3 \} \{shifts right\}

III.  \( Q < K \), not at equilibrium; shifts right  
\[ \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2} \];  
\( \frac{(0.50)}{(0.50)^2} = 2.0 \);  
\( Q < K \) \{2.0 < 8 \};  
\( Q < K \) \{shifts right\}

15. a.  \( K_c = 8.0 \)  
{write K expression:  \( K_c = \frac{[\text{AB}]^2}{[A_2][B_2]} \);  plug-in equilibrium values:  \( K_c = \frac{(4.0)^2}{(1.0)(2.0)} = 8.0 \)}

b.  \([A_2] = 1.657\text{M}; [B_2] = 1.657; [AB] = 4.686\text{M} \)  
{write ICE Table (non-bold}

values given in problem; bold values added after):  
\( K_c = \frac{[\text{AB}]^2}{[A_2][B_2]} \);  
\( 8 = \frac{(4+2x)^2}{(2-x)(2-x)} \);

cross multiply:  \( 5.656 - 2.828x = 4 + 2x \);  
simplify:  \( 4.828x = 1.656 \);  
\( x = 0.343 \);  
\( [A_2] = [B_2] = 2.0 - x = 2.0 - 0.343 = 1.657\text{M}; \)

\( [AB] = 4.0 + 2x = 4.0 + 2(0.343) = 4.686\text{M} \)

c.  \( K_c = \frac{[\text{AB}]^2}{[A_2][B_2]} \);  
\( K_c = \frac{(4.686)^2}{(1.657)(1.657)} = \frac{21.959}{2.746} = 7.998 = 8 \)  
(note how the value of K returns to its original value)