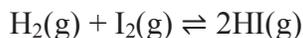


**Using Stoichiometry and K_{eq} to Calculate Equilibrium Concentrations
(Student textbook page 451)**

51. At 1100 K, hydrogen, $\text{H}_2(\text{g})$, and iodine, $\text{I}_2(\text{g})$, combine to form hydrogen iodide:



At equilibrium in a 1.0 L reaction vessel, the mixture of gases contained 0.30 mol of hydrogen, 1.3 mol of iodine, and 3.4 mol of hydrogen iodide. Determine the value of K_{eq} .

What Is Required?

You need to determine the value of K_{eq} .

What Is Given?

You know the balanced chemical equation for the equilibrium reaction.

You know the volume of the reaction vessel.

You know the initial amount in moles of each gas:

$$n_{\text{H}_2} = 0.30 \text{ mol}$$

$$n_{\text{I}_2} = 1.3 \text{ mol}$$

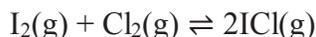
$$n_{\text{HI}} = 3.4 \text{ mol}$$

Plan Your Strategy	Act on Your Strategy
Express the equilibrium concentration of the components in mol/L.	$[\text{H}_2] = 0.30 \text{ mol/L}$ $[\text{I}_2] = 1.3 \text{ mol/L}$ $[\text{HI}] = 3.4 \text{ mol/L}$
Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$
Substitute the equilibrium concentrations into the expression and solve for K_{eq} .	$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$ $= \frac{(3.4)^2}{(0.30)(1.3)}$ $= 29.641$ $= 3.0 \times 10^1$ The equilibrium constant is 3.0×10^1 .

Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactants in the denominator each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. The answer is reasonable and has the correct number of significant digits.

52. At 25°C, the following reaction takes place:



K_{eq} for the reaction is 82. If 0.83 mol of iodine gas, $\text{I}_2(\text{g})$, and 0.83 mol of chlorine gas, $\text{Cl}_2(\text{g})$, are placed in a 10 L container at 25°C, what are the concentrations of the various gases at equilibrium?

What Is Required?

You need to determine the equilibrium concentrations of each component.

What Is Given?

You know the balanced chemical equation for the reaction.

You know the value of the equilibrium constant, K_{eq} : 82

You know the volume of the container: 10 L

You know the initial amount in moles of the iodine gas and the chlorine gas:

$$n_{\text{I}_2} = 0.83 \text{ mol}$$

$$n_{\text{Cl}_2} = 0.83 \text{ mol}$$

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \frac{n}{V}$ to express the initial concentration of the components in mol/L	$[\text{I}_2] = \frac{0.83 \text{ mol}}{10 \text{ L}}$ $= 0.083 \text{ mol/L}$ $[\text{Cl}_2] = \frac{0.83 \text{ mol}}{10 \text{ L}}$ $= 0.083 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent the decrease in concentration in mol/L of $\text{I}_2(\text{g})$ when the system reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

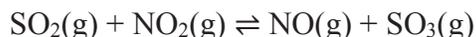
	$\text{I}_2(\text{g})$	+	$\text{Cl}_2(\text{g})$	\rightleftharpoons	$2\text{ICl}(\text{g})$
	$[\text{I}_2] \text{ (mol/L)}$		$[\text{Cl}_2] \text{ (mol/L)}$		$[\text{ICl}] \text{ (mol/L)}$
I	0.083		0.083		0
C	$-x$		$-x$		$+2x$
E	$0.083 - x$		$0.083 - x$		$2x$

Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$
Substitute the equilibrium concentrations into the expression for K_{eq} and solve for x .	$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$ $82 = \frac{(2x)^2}{(0.083 - x)(0.083 - x)}$ $\sqrt{82} = \sqrt{\frac{(2x)^2}{(0.083 - x)(0.083 - x)}}$ $9.05538 = \frac{2x}{0.083 - x}$ $x = 0.06798 \text{ mol/L}$ $= 0.068 \text{ mol/L}$
Determine the equilibrium concentrations of each component.	$[\text{H}_2] = [\text{I}_2]$ $= 0.083 \text{ mol/L} - 0.068 \text{ mol/L}$ $= 0.015 \text{ mol/L}$ $[\text{HI}] = 2(0.068 \text{ mol/L})$ $= 0.136 \text{ mol/L}$ $= 0.14 \text{ mol/L}$

Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 87$. Considering the rounding for the final answers, this approximates the given K_{eq} . The answers are reasonable.

53. A chemist is studying the following reaction at a certain temperature:



In a 1.0 L container, the chemist adds 0.17 mol of sulfur dioxide, $\text{SO}_2(\text{g})$, to 0.11 mol of nitrogen dioxide, $\text{NO}_2(\text{g})$. At equilibrium, the concentration of sulfur trioxide, $\text{SO}_3(\text{g})$, is 0.089 mol/L. What is the value of K_{eq} for the reaction at this temperature?

What Is Required?

You need to determine the value of K_{eq} for the reaction.

What Is Given?

You know the balanced chemical equation for the reaction.

You know the volume of the container: 1.0 L

You know the initial amounts in moles of sulfur dioxide and nitrogen dioxide:

$$n_{\text{SO}_2} = 0.17 \text{ mol}$$

$$n_{\text{NO}_2} = 0.11 \text{ mol}$$

You know the concentration of sulfur trioxide at equilibrium: $[\text{SO}_3(\text{g})] = 0.089 \text{ mol/L}$

Plan Your Strategy	Act on Your Strategy
Express the initial concentrations of $\text{SO}_2(\text{g})$ and $\text{NO}_2(\text{g})$ in mol/L.	$[\text{SO}_2(\text{g})] = 0.17 \text{ mol/L}$ $[\text{NO}_2(\text{g})] = 0.11 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations of each component. Use the stoichiometry of the reaction and the given equilibrium concentration of $\text{SO}_3(\text{g})$ to determine the equilibrium concentrations of each component.	See the ICE table below.

	$\text{SO}_2(\text{g})$	+	$\text{NO}_2(\text{g})$	\rightleftharpoons	$\text{NO}(\text{g})$	+	$\text{SO}_3(\text{g})$
	$[\text{SO}_2] \text{ (mol/L)}$		$[\text{NO}_2] \text{ (mol/L)}$		$[\text{NO}] \text{ (mol/L)}$		$[\text{SO}_3] \text{ (mol/L)}$
I	0.170		0.110		0		0
C	$0.170 - 0.089$		$0.110 - 0.089$		0.089		0.089
E	0.081		0.021		0.089		0.089

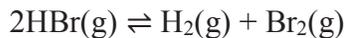
Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{NO}][\text{SO}_3]}{[\text{SO}_2][\text{NO}_2]}$
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<p>Substitute the equilibrium concentrations into the expression and solve for K_{eq}.</p>	$K_{\text{eq}} = \frac{[\text{NO}][\text{SO}_3]}{[\text{SO}_2][\text{NO}_2]}$ $= \frac{(0.089)(0.089)}{(0.081)(0.021)}$ $= 4.6566$ $= 4.7$ <p>The equilibrium constant for the reaction is 4.7.</p>
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Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. The answer is reasonable and has the correct number of significant digits.

54. Hydrogen bromide, HBr(g) , decomposes at 700 K, as represented by the following reaction:



K_{eq} is 4.2×10^{-9} for this reaction. If 0.090 mol of hydrogen bromide is placed into a 2.0 L reaction vessel and heated to 700 K, what is the equilibrium concentration of each gas?

What Is Required?

You need to calculate the equilibrium concentration of each gas?

What Is Given?

You know the balanced chemical equation for the reaction.

You know the value of the equilibrium constant, K_{eq} : 4.2×10^{-9}

You know the volume, V , of the reaction vessel: 2.0 L

You know the initial amount in moles of hydrogen bromide: $n_{\text{HBr}} = 0.090 \text{ mol}$

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \frac{n}{V}$ to express the initial concentration, c , of HBr in mol/L.	$[\text{HBr}] = \frac{0.090 \text{ mol}}{2.0 \text{ L}} = 0.045 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent the concentration, in mol/L, of $\text{Br}_2\text{(g)}$ when the system reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$2\text{HBr(g)} \rightleftharpoons$	$\text{H}_2\text{(g)}$	$+$	$\text{Br}_2\text{(g)}$
	$[\text{HBr}](\text{mol/L})$	$[\text{H}_2](\text{mol/L})$		$[\text{Br}_2](\text{mol/L})$
I	0.045	0		0
C	$-2x$	$+x$		$+x$
E	$0.045 - 2x$	x		x

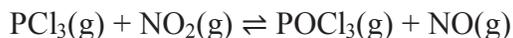
Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{H}_2][\text{Br}_2]}{[\text{HBr}]^2}$
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<p>Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.</p>	$K_{\text{eq}} = \frac{[\text{H}_2][\text{Br}_2]}{[\text{HBr}]^2}$ $4.2 \times 10^{-9} = \frac{(x)(x)}{(0.045 - 2x)^2}$ $\sqrt{4.2 \times 10^{-9}} = \sqrt{\frac{(x)^2}{(0.045 - 2x)^2}}$ $6.4807 \times 10^{-5} = \frac{(x)}{0.045 - 2x}$ $x = 2.9 \times 10^{-6} \text{ mol/L}$
<p>Determine the equilibrium concentration of each component.</p>	$[\text{H}_2] = [\text{Br}_2]$ $= 2.9 \times 10^{-6} \text{ mol/L}$ $[\text{HBr}] = 0.045 - 2(2.9 \times 10^{-6} \text{ mol/L})$ $= 0.045 \text{ mol/L}$

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactant in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 4.2 \times 10^{-9}$. This closely approximates the given K_{eq} . The answers are reasonable.

55. For the reaction of phosphorus trichloride, PCl_3 , and nitrogen dioxide, NO_2 , an equilibrium is established according to the following equation:



Initially, 1.24 mol of each reactant is placed into a 1.0 L container and equilibrium is established. If $K_{\text{eq}} = 3.77$ for the reaction, what are the equilibrium concentrations of all reactants and products?

What Is Required?

You need to calculate the equilibrium concentrations of all reactants and products.

What Is Given?

You know the balanced chemical equation for the reaction.

You know the initial concentration of each reactant in the equilibrium mixture: 1.24 mol/L

You know the value of the equilibrium constant, K_{eq} : 3.77

Plan Your Strategy	Act on Your Strategy
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $\text{PCl}_3(\text{g})$ had decreased when the system reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$\text{PCl}_3(\text{g})$	+	$\text{NO}_2(\text{g})$	\rightleftharpoons	$\text{POCl}_3(\text{g})$	+	$\text{NO}(\text{g})$
	$[\text{PCl}_3](\text{mol/L})$		$[\text{NO}_2](\text{mol/L})$		$[\text{POCl}_3](\text{mol/L})$		$[\text{NO}](\text{mol/L})$
I	1.24		1.24		0		0
C	$-x$		$-x$		$+x$		$+x$
E	$1.24 - x$		$1.24 - x$		x		x

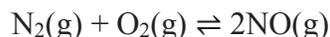
Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{POCl}_3][\text{NO}]}{[\text{PCl}_3][\text{NO}_2]}$
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<p>Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.</p>	$K_{\text{eq}} = \frac{[\text{POCl}_3][\text{NO}]}{[\text{PCl}_3][\text{NO}_2]}$ $3.77 = \frac{(x)(x)}{(1.24 - x)(1.24 - x)}$ $\sqrt{3.77} = \sqrt{\frac{(x)(x)}{(1.24 - x)(1.24 - x)}}$ $1.9416 = \frac{x}{1.24 - x}$ $x = 0.818$
<p>Determine the equilibrium concentration of each component.</p>	$[\text{POCl}_3] = [\text{NO}]$ $= 0.82 \text{ mol/L}$ $[\text{PCl}_3] = [\text{NO}_2]$ $= 1.24 \text{ mol/L} - 0.82 \text{ mol/L}$ $= 0.42 \text{ mol/L}$

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 3.81$. This closely approximates the given K_{eq} . The answers are reasonable.

56. When nitrogen gas reacts with oxygen gas, nitrogen monoxide gas forms:



Initially, 0.30 mol of each reactant is placed into a 2.0 L container and equilibrium is established. If $K_{\text{eq}} = 52.1$ for the reaction, what are the equilibrium concentrations of all reactants and products?

What Is Required?

You need to determine the equilibrium concentrations of the reactants and products.

What Is Given?

You know the balanced chemical equation for the reaction.

You know the value of the equilibrium constant, K_{eq} : 52.1

You know the volume, V , of the container: 2.0 L

You know the initial amount in moles of each of $\text{N}_2(\text{g})$ and $\text{O}_2(\text{g})$: $n = 0.30$ mol

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \frac{n}{V}$ to express the initial concentrations, c , of $\text{N}_2(\text{g})$ and $\text{O}_2(\text{g})$ in mol/L.	$[\text{N}_2] = [\text{O}_2]$ $= \frac{0.30 \text{ mol}}{2.0 \text{ L}}$ $= 0.15 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $\text{N}_2(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

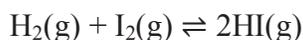
	$\text{N}_2(\text{g})$	+	$\text{O}_2(\text{g})$	\rightleftharpoons	$2\text{NO}(\text{g})$
	$[\text{N}_2] \text{ (mol/L)}$		$[\text{O}_2] \text{ (mol/L)}$		$[\text{NO}] \text{ (mol/L)}$
I	0.15		0.15		0
C	$-x$		$-x$		$+2x$
E	$0.15 - x$		$0.15 - x$		$2x$

Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x .	$K_{\text{eq}} = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]}$ $52.1 = \frac{(2x)^2}{(0.15 - x)(0.15 - x)}$ $\sqrt{52.1} = \sqrt{\frac{(2x)^2}{(0.15 - x)(0.15 - x)}}$ $7.2180 = \frac{2x}{(0.15 - x)}$ $x = 0.11745 \text{ mol/L}$
Determine the equilibrium concentration of each component.	$[\text{NO}] = 2x$ $= 2(0.11745 \text{ mol/L})$ $= 0.2349 \text{ mol/L}$ $= 0.23 \text{ mol/L}$ $[\text{N}_2] = [\text{O}_2]$ $= 0.15 - 0.11745$ $= 3.3 \times 10^{-2} \text{ mol/L}$

Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 48.6$. This approximates the given K_{eq} . The answers are reasonable.

57. In the reaction of hydrogen gas reacting with iodine gas to form hydrogen iodide gas, $K_{\text{eq}} = 49.6$ at 730 K. The reaction is as follows:



Initially, 3.51 mol of each reactant is placed into a 3.0 L container and equilibrium is established. Determine the amount in moles of each substance that will be present when the system reaches equilibrium at 730 K.

What Is Required?

You need to determine the amount in moles of each substance at equilibrium.

What Is Given?

You know the balanced chemical equation for the reaction.

You know the value of the equilibrium constant, K_{eq} : 49.6

You know the volume, V , of the container: 3.0 L

You know the initial amount in moles of each of $\text{H}_2(\text{g})$ and $\text{I}_2(\text{g})$: $n = 3.51$ mol

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \frac{n}{V}$ to express the initial concentrations, c , in mol/L.	$[\text{H}_2] = \frac{[\text{I}_2]}{3.51 \text{ mol}} = \frac{3.51 \text{ mol}}{3.0 \text{ L}} = 1.17 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $\text{H}_2(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$\text{H}_2(\text{g})$	+	$\text{I}_2(\text{g})$	\rightleftharpoons	$2\text{HI}(\text{g})$
	$[\text{H}_2] \text{ (mol/L)}$		$[\text{I}_2] \text{ (mol/L)}$		$[\text{HI}] \text{ (mol/L)}$
I	1.17		1.17		0
C	$-x$		$-x$		$+2x$
E	$1.17 - x$		$1.17 - x$		$2x$

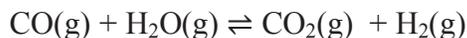
Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$
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<p>Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.</p>	$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$ $49.6 = \frac{(2x)^2}{(1.17-x)(1.17-x)}$ $\sqrt{49.6} = \sqrt{\frac{(2x)^2}{(1.17-x)(1.17-x)}}$ $7.0427 = \frac{2x}{1.17-x}$ $x = 0.91123 \text{ mol/L}$ $= 0.911 \text{ mol/L}$
<p>Determine the equilibrium concentration of each component.</p>	$[\text{H}_2] = [\text{I}_2]$ $= 1.17 \text{ mol/L} - 0.911 \text{ mol/L}$ $= 0.259 \text{ mol/L}$ $= 0.26 \text{ mol/L}$ $[\text{HI}] = 2(0.911 \text{ mol/L})$ $= 1.822 \text{ mol/L}$ $= 1.8 \text{ mol/L}$
<p>Use the formula $n = cV$ to determine the amount in moles, n, of each component.</p>	$n_{\text{H}_2} = cV$ $= (0.26 \text{ mol/L})(3.0 \text{ L})$ $= 0.78 \text{ mol}$ $n_{\text{I}_2} = cV$ $= (0.26 \text{ mol/L})(3.0 \text{ L})$ $= 0.78 \text{ mol}$ $n_{\text{HI}} = cV$ $= (1.8 \text{ mol/L})(3.0 \text{ L})$ $= 5.4 \text{ mol}$

Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression, gives $K_{\text{eq}} = 47.9$. This approximates the given K_{eq} . The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

58. Initially, 2.25 mol of carbon monoxide gas, CO(g), and 2.25 mol of water vapour, H₂O(g), are placed into a 1.50 L container and the following equilibrium is established, with a K_{eq} value of 4.2:



What is the amount in moles of each reactant and product at equilibrium?

What Is Required?

You need to determine the amount in moles of each substance at equilibrium.

What Is Given?

You know the balanced chemical equation for the reaction.

You know the value of the equilibrium constant, K_{eq} : 4.2

You know the volume, V , of the container = 1.50 L

You know initial amount in moles of CO(g) and H₂O(g): $n = 2.25$ mol

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \frac{n}{V}$ to express the initial concentrations, c , in mol/L.	$[\text{CO}] = [\text{H}_2\text{O}]$ $= \frac{2.25 \text{ mol}}{1.50 \text{ L}}$ $= 1.50 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of CO(g) had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	CO(g)	+	H ₂ O(g)	\rightleftharpoons	CO ₂ (g)	+	H ₂ (g)
	[CO] (mol/L)		[H ₂ O] (mol/L)		[CO ₂] (mol/L)		[H ₂] (mol/L)
I	1.50		1.50		0		0
C	- x		- x		+ x		+ x
E	1.50 - x		1.50 - x		x		x

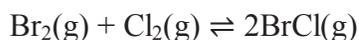
Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}$
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<p>Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.</p>	$K_{\text{eq}} = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}$ $4.2 = \frac{(x)(x)}{(1.50 - x)(1.5 - x)}$ $\sqrt{4.2} = \sqrt{\frac{(x)(x)}{(1.50 - x)(1.5 - x)}}$ $2.0494 = \frac{x}{1.50 - x}$ $x = 1.0081 \text{ mol/L}$
<p>Determine the equilibrium concentration of each component.</p>	$[\text{CO}] = [\text{H}_2\text{O}]$ $= 1.50 \text{ mol/L} - 1.0081 \text{ mol/L}$ $= 0.4919 \text{ mol/L}$ $[\text{CO}_2] = [\text{H}_2]$ $= 1.0081 \text{ mol/L}$
<p>Use the formula $n = cV$ to determine the amount in moles, n, of each component.</p>	$n_{\text{CO}} = cV$ $= (0.4919 \text{ mol/L})(1.5 \text{ L})$ $= 0.73785 \text{ mol}$ $= 0.74 \text{ mol}$ $n_{\text{H}_2\text{O}} = n_{\text{CO}}$ $= 0.74 \text{ mol}$ $n_{\text{CO}_2} = cV$ $= (1.0081 \text{ mol/L})(1.5 \text{ L})$ $= 1.5125 \text{ mol}$ $= 1.5 \text{ mol}$ $n_{\text{H}_2} = n_{\text{CO}_2}$ $= 1.5 \text{ mol}$

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 4.2$. This agrees with the given K_{eq} . The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

59. Bromine gas and chlorine gas establish equilibrium when they react to form bromine chloride gas. The K_{eq} for this reaction at 450 K is 28.8.



When 8.70 mol of each reactant is placed into a 5.00 L container, what is the amount in moles of each reactant and product at equilibrium?

What Is Required?

You need to determine the amount in moles of each substance at equilibrium.

What Is Given?

You know the balanced chemical equation for the reaction.

You know the value of the equilibrium constant, K_{eq} : 28.8

You know the volume, V , of the container: 5.00 L

You know the initial amount in moles of $\text{Br}_2(\text{g})$ and $(\text{Cl}_2(\text{g}))$: $n = 8.70 \text{ mol}$

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \frac{n}{V}$ to express the initial concentrations, c , in mol/L.	$[\text{Br}_2] = [\text{Cl}_2]$ $= \frac{8.70 \text{ mol}}{5.00 \text{ L}}$ $= 1.740 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent amount by which the concentration of $\text{Br}_2(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$\text{Br}_2(\text{g})$	+	$\text{Cl}_2(\text{g})$	\rightleftharpoons	$2\text{BrCl}(\text{g})$
	$[\text{Br}_2] \text{ (mol/L)}$		$[\text{Cl}_2] \text{ (mol/L)}$		$[\text{BrCl}] \text{ (mol/L)}$
I	1.740		1.740		0
C	$-x$		$-x$		$2x$
E					

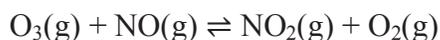
Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{BrCl}]^2}{[\text{Br}_2][\text{Cl}_2]}$
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<p>Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.</p>	$K_{\text{eq}} = \frac{[\text{BrCl}]^2}{[\text{Br}_2][\text{Cl}_2]}$ $28.8 = \frac{(2x)^2}{(1.740 - x)(1.740 - x)}$ $\sqrt{28.8} = \sqrt{\frac{(2x)^2}{(1.740 - x)(1.740 - x)}}$ $5.3665 = \frac{2x}{1.740 - x}$ $x = 1.2676 \text{ mol/L}$
<p>Determine the equilibrium concentration of each component.</p>	$[\text{Br}_2] = [\text{Cl}_2]$ $= 1.740 \text{ mol/L} - 1.2676 \text{ mol/L}$ $= 0.4724 \text{ mol/L}$ $= 0.472 \text{ mol/L}$ $[\text{BrCl}] = 2(1.2676 \text{ mol/L})$ $= 2.5352 \text{ mol/L}$
<p>Use the formula $n = cV$ to determine the amount in moles, n, of each component.</p>	$n_{\text{Br}_2} = cV$ $= (0.472 \text{ mol/L})(5.00 \text{ L})$ $= 2.36 \text{ mol}$ $n_{\text{Cl}_2} = n_{\text{Br}_2}$ $= 2.36 \text{ mol}$ $n_{\text{BrCl}} = cV$ $= (2.5352 \text{ mol/L})(5.00 \text{ L})$ $= 12.676 \text{ mol}$ $= 12.7 \text{ mol}$

Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 28.9$. This approximates the given K_{eq} . The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

60. Ozone, $\text{O}_3(\text{g})$, reacts with nitrogen monoxide, $\text{NO}(\text{g})$, and forms nitrogen dioxide $\text{NO}_2(\text{g})$, and oxygen, $\text{O}_2(\text{g})$. At 325 K, K_{eq} for the reaction is 6.70. The reaction is:



If 1.62 mol of each reactant is placed into a 2.00 L container, what is the amount in moles of each reactant and product at equilibrium?

What Is Required?

You need to determine the amount in moles of each substance at equilibrium.

What Is Given?

You know the balanced chemical equation for the reaction.

You know the value of the equilibrium constant, K_{eq} : 6.70

You know the volume, V , of the container: 2.00 L

You know the initial amount in moles of $\text{O}_3(\text{g})$ and $\text{NO}(\text{g})$: $n = 1.62 \text{ mol}$

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \frac{n}{V}$ to express the initial concentrations, c , in mol/L.	$[\text{O}_3] = [\text{NO}]$ $= \frac{1.62 \text{ mol}}{2.00 \text{ L}}$ $= 0.810 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent the the amount by which the concentration of $\text{O}_3(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$\text{O}_3(\text{g})$	+	$\text{NO}(\text{g})$	\rightleftharpoons	$\text{NO}_2(\text{g})$	+	$\text{O}_2(\text{g})$
	$[\text{O}_3] \text{ (mol/L)}$		$[\text{NO}] \text{ (mol/L)}$		$[\text{NO}_2] \text{ (mol/L)}$		$[\text{O}_2] \text{ (mol/L)}$
I	0.810		0.810		0		0
C	$-x$		$-x$		$+x$		$+x$
E	$0.810 - x$		$0.810 - x$		x		x

Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{NO}_2][\text{O}_2]}{[\text{O}_3][\text{NO}]}$
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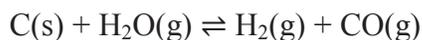
<p>Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.</p>	$K_{\text{eq}} = \frac{[\text{NO}_2][\text{O}_2]}{[\text{O}_3][\text{NO}]}$ $6.70 = \frac{(x)(x)}{(0.810 - x)(0.810 - x)}$ $\sqrt{6.70} = \sqrt{\frac{(x)(x)}{(0.810 - x)(0.810 - x)}}$ $2.5884 = \frac{x}{0.810 - x}$ $x = 0.58427 \text{ mol/L}$ $= 0.584 \text{ mol/L}$
<p>Determine the equilibrium concentration of each component.</p>	$[\text{O}_3] = [\text{NO}]$ $= 0.810 \text{ mol/L} - 0.584 \text{ mol/L}$ $= 0.226 \text{ mol/L}$ $[\text{NO}_2] = [\text{O}_2]$ $= 0.584 \text{ mol/L}$
<p>Use the formula $n = cV$ to determine the amount in moles, n, of each component.</p>	$n_{\text{O}_3} = cV$ $= (0.226 \text{ mol/L})(2.00 \text{ L})$ $= 0.452 \text{ mol}$ $n_{\text{NO}} = n_{\text{O}_3}$ $= 0.452 \text{ mol}$ $n_{\text{NO}_2} = cV$ $= (0.584 \text{ mol/L})(2.00 \text{ L})$ $= 1.168 \text{ mol}$ $= 1.17 \text{ mol}$ $n_{\text{O}_2} = n_{\text{NO}_2}$ $= 1.17 \text{ mol}$

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 6.67$. This approximates the given K_{eq} . The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

Using the Quadratic Equation to Determine Equilibrium Information
(Student textbook page 454)

61. When solid carbon reacts with water vapour, hydrogen gas and carbon dioxide gas form according to the following equation:



When 2.25 mol of water vapour is reacted with solid carbon in a 1.0 L container, and equilibrium is established, what concentration of each reactant and product in the gas phase will be present at equilibrium if K_{eq} equals 23.4?

What Is Required?

You need to calculate the concentration of each reactant and product in the gas phase at equilibrium.

What Is Given?

You know the balanced equation for the reaction and the state of each component in the reaction.

You know the volume of the container: 1.0 L

You know the initial amount in moles of $\text{H}_2\text{O(g)}$: $n = 2.25$ mol

You know the value of the equilibrium constant, K_{eq} : 23.4

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $\text{H}_2\text{O(g)}$ in mol/L.	$[\text{H}_2\text{O}] = 2.25$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration $\text{H}_2\text{O(g)}$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component in the gas phase.	See the ICE table below.

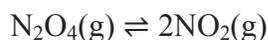
	C(s)	+	$\text{H}_2\text{O(g)}$	\rightleftharpoons	$\text{H}_2\text{(g)}$	+	CO(g)
			$[\text{H}_2\text{O}]$ (mol/L)		$[\text{H}_2]$ (mol/L)		$[\text{CO}]$ (mol/L)
I			2.25		0		0
C			$-x$		$+x$		$+x$
E			$2.25 - x$		x		x

Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{H}_2][\text{CO}]}{[\text{H}_2\text{O}]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} .	$K_{\text{eq}} = \frac{[\text{H}_2][\text{CO}]}{[\text{H}_2\text{O}]}$ $23.4 = \frac{(x)(x)}{(2.25 - x)}$
Re-arrange the expression into a quadratic equation.	$x^2 + 23.4x - 52.65 = 0$
Use the quadratic formula to solve for x . $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \frac{-23.4 \pm \sqrt{(23.4)^2 - (4)(-52.65)}}{2(1)}$ $= \frac{-23.4 \pm 27.535}{2}$ $x = 2.0674 \text{ or } x = -25.4674$ Because it is impossible to have a negative concentration, $x = 2.0674$.
Determine the equilibrium concentration of each component in the gas phase.	$[\text{H}_2\text{O}] = 2.25 \text{ mol/L} - 2.0674 \text{ mol/L}$ $= 0.1826 \text{ mol/L}$ $= 0.18 \text{ mol/L}$ $[\text{H}_2] = [\text{CO}]$ $= 2.0674$ $= 2.1 \text{ mol/L}$

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactant in the denominator each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 6.67$. This approximates the given K_{eq} . The answers are reasonable and are expressed to the correct number of significant digits.

62. Dinitrogen tetroxide, $\text{N}_2\text{O}_4(\text{g})$, reversibly decomposes into nitrogen dioxide, $\text{NO}_2(\text{g})$, according to the following equation:



This reaction has resulted in one of the most important rocket propellant processes that researchers have developed. At 325 K, the K_{eq} value is 0.91. Initially, 0.34 mol $\text{N}_2\text{O}_4(\text{g})$ is placed into a 1.00 L container and equilibrium is established. Determine the equilibrium concentration of each gas in the container.

What Is Required?

You need to determine the equilibrium concentration of each gas in the container.

What Is Given?

You know the balanced equation for the reaction and the state of each component in the reaction.

You know the volume of the container: 1.00 L

You know the initial amount in moles of $\text{N}_2\text{O}_4(\text{g})$: $n = 0.34$ mol

You know the value of the equilibrium constant, K_{eq} : 0.91

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $\text{N}_2\text{O}_4(\text{g})$ in mol/L.	$[\text{N}_2\text{O}_4] = 0.34$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $\text{N}_2\text{O}_4(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component in the gas phase.	See the ICE table below.

$\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$		
	$[\text{N}_2\text{O}_4]$ (mol/L)	$[\text{NO}_2]$ (mol/L)
I	0.34	0
C	$-x$	$+2x$
E	$0.34 - x$	$2x$

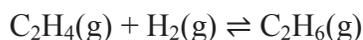
Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$
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Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} .	$K_{\text{eq}} = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$ $0.91 = \frac{(2x)^2}{(0.34 - x)}$
Re-arrange the expression into a quadratic equation.	$4x^2 + 0.91x - 0.3094 = 0$
Use the quadratic formula to solve for x . $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \frac{-0.91 \pm \sqrt{(0.91)^2 - (4)(4)(-0.3094)}}{2(4)}$ $= \frac{-0.91 \pm 2.4038}{2(4)}$ $x = 0.18673 \text{ or } x = -0.4142$ <p>Because it is impossible to have a negative concentration, $x = 0.18673$.</p>
Determine the equilibrium concentration of each component in the gas phase.	$[\text{N}_2\text{O}_4] = 0.34 \text{ mol/L} - 0.18673 \text{ mol/L}$ $= 0.15327 \text{ mol/L}$ $= 0.15 \text{ mol/L}$ $[\text{NO}_2] = 2(0.18673) \text{ mol/L}$ $= 0.37346 \text{ mol/L}$ $= 0.37 \text{ mol/L}$

Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactant in the denominator each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 0.91$. This agrees with the given K_{eq} . The answers are reasonable and are expressed to the correct number of significant digits.

63. In an equilibrium process, ethene, $\text{C}_2\text{H}_4(\text{g})$, is reacted with hydrogen gas to produce ethane, $\text{C}_2\text{H}_6(\text{g})$. At a given temperature, the K_{eq} value is 1.04 for the following reaction:



Initially, 0.34 mol $\text{C}_2\text{H}_4(\text{g})$ and 0.53 mol $\text{H}_2(\text{g})$ are placed into a 1.00 L container and equilibrium is established. Determine the equilibrium concentration of each gas in the container.

What Is Required?

You need to determine the equilibrium concentration of each gas in the container.

What Is Given?

You know the balanced equation for the reaction and the state of each component in the reaction.

You know the volume of the container: 1.00 L

You know the initial amount in moles, n , of $\text{C}_2\text{H}_4(\text{g})$: 0.34 mol

You know the initial amount in moles, n , of $\text{H}_2(\text{g})$: 0.53 mol

You know the value of the equilibrium constant, K_{eq} : 1.04

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $\text{C}_2\text{H}_4(\text{g})$ and $\text{H}_2(\text{g})$ in mol/L.	$[\text{N}_2\text{O}_4] = 0.34 \text{ mol/L}$ $[\text{H}_2] = 0.53 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $\text{C}_2\text{H}_4(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component in the gas phase.	See the ICE table below.

	$\text{C}_2\text{H}_4(\text{g})$	+	$\text{H}_2(\text{g})$	\rightleftharpoons	$\text{C}_2\text{H}_6(\text{g})$
	$[\text{C}_2\text{H}_4] \text{ (mol/L)}$		$[\text{H}_2] \text{ (mol/L)}$		$[\text{C}_2\text{H}_6] \text{ (mol/L)}$
I	0.34		0.53		0
C	$-x$		$-x$		$+x$
E	$0.34 - x$		$0.53 - x$		x

Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{C}_2\text{H}_6]}{[\text{C}_2\text{H}_4][\text{H}_2]}$
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Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} .	$K_{\text{eq}} = \frac{[\text{C}_2\text{H}_6]}{[\text{C}_2\text{H}_4][\text{H}_2]}$ $1.04 = \frac{x}{(0.34 - x)(0.53 - x)}$
Re-arrange the expression into a quadratic equation.	$1.04(0.1802 - 0.87x + x^2) = x$ $1.04x^2 - 1.9048x + 0.1874 = 0$
Use the quadratic formula to solve for x . $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \frac{-(-1.9048) \pm \sqrt{(-1.9048)^2 - (4)(1.04)(0.1874)}}{2(1.04)}$ $\frac{-(-1.9048) \pm 1.6878}{2.08}$ $x = 0.1047 \text{ or } x = 1.7272$ <p>Because it is impossible to have a negative concentration, x cannot be greater than 0.34. Therefore, $x = 0.1047$.</p>
Determine the equilibrium concentration of each component in the gas phase.	$[\text{C}_2\text{H}_4] = 0.34 \text{ mol/L} - 0.1047 \text{ mol/L}$ $= 0.2353 \text{ mol/L}$ $= 0.24 \text{ mol/L}$ $[\text{H}_2] = 0.53 \text{ mol/L} - 0.1047 \text{ mol/L}$ $= 0.4253 \text{ mol/L}$ $= 0.43 \text{ mol/L}$ $[\text{C}_2\text{H}_6] = 0.1047 \text{ mol/L}$ $= 0.10 \text{ mol/L}$

Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 0.97$. This approximates the given K_{eq} . The answers are reasonable and are expressed to the correct number of significant digits.

64. When 1.88 mol of hydrogen gas and 2.86 mol of iodine gas are placed in a 2.00 L container, an equilibrium is reached with hydrogen iodide gas such that the K_{eq} value is 55.3 at 700 K. Determine the amount in moles of each reactant and product in the container when equilibrium is established.

What Is Required?

You need to determine the amount in moles of each reactant and product at equilibrium.

What Is Given?

You know the volume, V , of the container: 2.00 L

You know the initial amount in moles of $\text{H}_2(\text{g})$: $n = 1.88 \text{ mol}$

You know the initial amount in moles of $\text{I}_2(\text{g})$: $n = 2.86 \text{ mol}$

You know the value of the equilibrium constant, K_{eq} : 55.3

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the equilibrium reaction.	$\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$
Use $c = \frac{n}{V}$ to express the initial concentration of $\text{H}_2(\text{g})$ and $\text{I}_2(\text{g})$ in mol/L.	$[\text{H}_2] = \frac{1.88 \text{ mol}}{2.00 \text{ L}}$ $= 0.940 \text{ mol/L}$ $[\text{I}_2] = \frac{2.86 \text{ mol}}{2.00 \text{ L}}$ $= 1.43 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $\text{H}_2(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component in the gas phase.	See the ICE table below.

	$\text{H}_2(\text{g})$	+	$\text{I}_2(\text{g})$	\rightleftharpoons	$2\text{HI}(\text{g})$
	$[\text{H}_2] \text{ (mol/L)}$		$[\text{I}_2] \text{ (mol/L)}$		$[\text{HI}] \text{ (mol/L)}$
I	0.940		1.43		0
C	$-x$		$-x$		$+2x$
E	$0.940 - x$		$1.43 - x$		$2x$

Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} .	$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$ $55.3 = \frac{(2x)^2}{(0.940 - x)(1.43 - x)}$
Re-arrange the expression into a quadratic equation.	$55.3(1.3442 - 2.37x + x^2) = 4x^2$ $51.4x^2 - 131.061x + 74.3232 = 0$
Use the quadratic formula to solve for x . $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \frac{-(-131.061) \pm \sqrt{(-131.061)^2 - (4)(51.3)(74.3232)}}{2(51.3)}$ $= \frac{131.061 \pm 43.885}{102.6}$ $x = 1.7051 \text{ or } x = 0.8497$ Because it is impossible to have a negative concentration, x cannot be greater than 0.940. Therefore, $x = 0.8497$.
Determine the equilibrium concentration of each component in the gas phase.	$[\text{H}_2] = 0.940 \text{ mol/L} - 0.8497 \text{ mol/L}$ $= 0.0903 \text{ mol/L}$ $[\text{I}_2] = 1.43 \text{ mol/L} - 0.8497 \text{ mol/L}$ $= 0.5803 \text{ mol/L}$ $[\text{HI}] = 2(0.8497 \text{ mol/L})$ $= 1.6994 \text{ mol/L}$
Use the formula $n = cV$ to calculate the amount in moles, n , of each component.	$n_{\text{H}_2} = cV$ $= (0.0903 \text{ mol/L})(2.00 \text{ L})$ $= 0.181 \text{ mol}$ $n_{\text{I}_2} = cV$ $= (0.5803 \text{ mol/L})(2.00 \text{ L})$ $= 1.16 \text{ mol}$ $n_{\text{HI}} = cV$ $= (1.6994 \text{ mol/L})(2.00 \text{ L})$ $= 3.40 \text{ mol}$

Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression $K_{\text{eq}} = 55.1$. This is very close to the given K_{eq} . The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

65. Carbonyl chloride gas, also called phosgene, $\text{COCl}_2(\text{g})$, was used during World War I as a chemical weapon. Now, it is used in a process to manufacture plastics such as the ones used to make lenses in eyeglasses. It is formed by reacting carbon monoxide gas with chlorine gas in the following equilibrium system:



The K_{eq} for this reaction at 1000 K is 24.0. If 2.35 mol $\text{CO}(\text{g})$ and 1.14 mol $\text{Cl}_2(\text{g})$ are placed into a 1.00 L container, what concentration of each gas will exist at equilibrium at 1000 K?

What Is Required?

You need to determine the equilibrium concentration of each gas in the container.

What Is Given?

You know the balanced equation for the reaction.

You know the state of each component in the reaction.

You know the volume of the container: 1.00 L

You know the initial amount in moles of $\text{CO}(\text{g})$: $n = 2.35$ mol

You know the initial amount in moles of $\text{Cl}_2(\text{g})$: $n = 1.14$ mol

You know the value of the equilibrium constant, K_{eq} : 24.0

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $\text{CO}(\text{g})$ and $\text{Cl}_2(\text{g})$ in mol/L.	$[\text{CO}] = 2.35$ mol/L $[\text{Cl}_2] = 1.14$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $\text{CO}(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

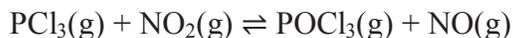
	$\text{CO}(\text{g})$	+	$\text{Cl}_2(\text{g})$	\rightleftharpoons	$\text{COCl}_2(\text{g})$
	$[\text{CO}]$ (mol/L)		$[\text{Cl}_2]$ (mol/L)		$[\text{COCl}_2]$ (mol/L)
I	2.35		1.14		0
C	$-x$		$-x$		$+x$
E	$2.35 - x$		$1.14 - x$		x

Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} .	$K_{\text{eq}} = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]}$ $24.0 = \frac{x}{(2.35 - x)(1.14 - x)}$
Re-arrange the expression into a quadratic equation.	$24.0(2.679 - 3.49x + x^2) = x$ $24.0x^2 - 84.76x + 64.296 = 0$
Use the quadratic formula to solve for x . $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \frac{-(-84.76) \pm \sqrt{(-84.76)^2 - (4)(24.0)(64.296)}}{2(24.0)}$ $= \frac{84.76 \pm 31.809}{48.0}$ $x = 2.4285 \text{ or } x = 1.103$ Because it is impossible to have a negative concentration, x cannot be greater than 1.14. Therefore, $x = 1.1031$.
Determine the equilibrium concentration of each component in the gas phase.	The equilibrium concentrations are: $[\text{CO}] = 2.35 \text{ mol/L} - 1.1031 \text{ mol/L}$ $= 1.2469 \text{ mol/L}$ $= 1.25 \text{ mol/L}$ $[\text{Cl}_2] = 1.14 \text{ mol/L} - 1.1031 \text{ mol/L}$ $= 0.0370 \text{ mol/L}$ $[\text{COCl}_2] = 1.1031 \text{ mol/L}$ $= 1.10 \text{ mol/L}$

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 23.8$. This agrees closely with the given K_{eq} . The answers are reasonable and are expressed to the correct number of significant digits.

66. Phosphoryl chloride, $\text{POCl}_3(\text{g})$, is used in the manufacturing of flame retardants. It is manufactured in an equilibrium process in which phosphorus trichloride reacts with nitrogen dioxide to form $\text{POCl}_3(\text{g})$ and $\text{NO}(\text{g})$ according to the following equation:



The K_{eq} for this reaction at 800 K is 1.82. If 1.86 mol $\text{PCl}_3(\text{g})$ and 1.64 mol $\text{NO}_2(\text{g})$ are placed into a 2.00 L container, what is the amount in moles of each gas at equilibrium at 800 K?

What Is Required?

You need to determine the amount in moles of each reactant and product at equilibrium.

What Is Given?

You know the volume of the container: 2.00 L

You know the initial amount in moles of $\text{PCl}_3(\text{g})$: $n = 1.86$ mol

You know the initial amount in moles of $\text{NO}_2(\text{g})$: $n = 1.64$ mol

You know the value of the equilibrium constant, K_{eq} : 1.82

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $\text{PCl}_3(\text{g})$ and $\text{NO}_2(\text{g})$ in mol/L.	$[\text{PCl}_3] = \frac{1.86 \text{ mol}}{2.00 \text{ L}}$ $= 0.930 \text{ mol/L}$ $[\text{NO}_2] = \frac{1.64 \text{ mol}}{2.00 \text{ L}}$ $= 0.820 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $\text{PCl}_3(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component in the gas phase.	See the ICE table below.

	$\text{PCl}_3(\text{g})$	+	$\text{NO}_2(\text{g})$	\rightleftharpoons	$\text{POCl}_3(\text{g})$	+	$\text{NO}(\text{g})$
	$[\text{PCl}_3] \text{ (mol/L)}$		$[\text{NO}_2] \text{ (mol/L)}$		$[\text{POCl}_3] \text{ (mol/L)}$		$[\text{NO}] \text{ (mol/L)}$
I	0.930		0.820		0		0
C	$-x$		$-x$		$+x$		$+x$
E	$0.930 - x$		$0.820 - x$		x		x

Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{POCl}_3][\text{NO}]}{[\text{PCl}_3][\text{NO}_2]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} .	$K_{\text{eq}} = \frac{[\text{POCl}_3][\text{NO}]}{[\text{PCl}_3][\text{NO}_2]}$ $1.82 = \frac{(x)(x)}{(0.930 - x)(0.820 - x)}$
Re-arrange the expression into a quadratic equation.	$1.82(0.7626 - 1.75x + x^2) = x^2$ $0.820x^2 - 3.185x + 1.3879 = 0$
Use the quadratic formula to solve for x . $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \frac{-(-3.185) \pm \sqrt{(-3.185)^2 - (4)(0.820)(1.3879)}}{2(0.820)}$ $= \frac{3.185 \pm 2.3647}{1.64}$ $x = 3.3836 \text{ or } x = 0.5005$ <p>Because it is impossible to have a negative concentration, x cannot be greater than 0.820. Therefore, $x = 0.5005$.</p>
Determine the equilibrium concentration of each component in the gas phase.	At equilibrium: $[\text{PCl}_3] = 0.930 \text{ mol/L} - 0.5005 \text{ mol/L}$ $= 0.430 \text{ mol/L}$ $[\text{NO}_2] = 0.820 \text{ mol/L} - 0.5005 \text{ mol/L}$ $= 0.320 \text{ mol/L}$ $[\text{POCl}_3] = [\text{NO}]$ $= 0.5005 \text{ mol/L}$
Use the formula $n = cV$ to calculate the amount in moles, n , of each component.	$n_{\text{PCl}_3} = cV$ $= (0.430 \text{ mol/L})(2.00 \text{ L})$ $= 0.860 \text{ mol}$ $n_{\text{NO}_2} = cV$ $= (0.320 \text{ mol/L})(2.00 \text{ L})$ $= 0.640 \text{ mol}$ $n_{\text{POCl}_3} = cV$ $= (0.5005 \text{ mol/L})(2.00 \text{ L})$ $= 1.00 \text{ mol}$

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression $K_{\text{eq}} = 1.82$. This agrees with the given K_{eq} . The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

67. In the reaction $\text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons \text{PCl}_5(\text{g})$, the K_{eq} value at 305 K is 14.15. Initially, 0.81 mol $\text{PCl}_3(\text{g})$ and 1.37 mol $\text{Cl}_2(\text{g})$ are placed into a 1.00 L container and allowed to reach equilibrium. What are the equilibrium concentrations of all these substances?

What Is Required?

You need to determine the equilibrium concentration of each gas in the container.

What Is Given?

You know the balanced equation for the reaction.

You know the state of each component in the reaction.

You know the volume of the container: 1.00 L

You know the initial amount in moles of $\text{PCl}_3(\text{g})$: $n = 0.81$ mol

You know the initial amount in moles of $\text{Cl}_2(\text{g})$: $n = 1.37$ mol

You know the value of the equilibrium constant, K_{eq} : 14.15

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $\text{PCl}_3(\text{g})$ and $\text{Cl}_2(\text{g})$ in mol/L.	$[\text{PCl}_3] = 0.81$ mol/L $[\text{Cl}_2] = 1.37$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $\text{PCl}_3(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

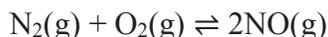
	$\text{PCl}_3(\text{g})$	+	$\text{Cl}_2(\text{g})$	\rightleftharpoons	$\text{PCl}_5(\text{g})$
	$[\text{PCl}_3]$ (mol/L)		$[\text{Cl}_2]$ (mol/L)		$[\text{PCl}_5]$ (mol/L)
I	0.81		1.37		0
C	$-x$		$-x$		$+x$
E	$0.81 - x$		$1.37 - x$		x

Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{PCl}_5]}{[\text{PCl}_3][\text{Cl}_2]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} .	$K_{\text{eq}} = \frac{[\text{PCl}_5]}{[\text{PCl}_3][\text{Cl}_2]}$ $14.15 = \frac{x}{(0.81-x)(1.37-x)}$
Re-arrange the expression into a quadratic equation.	$14.15(1.1087 - 2.18x + x^2) = x$ $14.15x^2 - 31.847x + 15.7022 = 0$
Use the quadratic formula to solve for x . $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \frac{-(-31.847) \pm \sqrt{(-31.847)^2 - (4)(14.15)(15.7022)}}{2(14.15)}$ $= \frac{31.847 \pm 11.2021}{28.3}$ $x = 0.7295 \text{ or } x = 1.521$ Because it is impossible to have a negative concentration, x cannot be greater than 0.81. Therefore, $x = 0.7295$
Determine the equilibrium concentration of each component in the gas phase.	At equilibrium: $[\text{PCl}_3] = 0.81 \text{ mol/L} - 0.7295 \text{ mol/L}$ $= 0.080 \text{ mol/L}$ $[\text{Cl}_2] = 1.37 \text{ mol/L} - 0.7295 \text{ mol/L}$ $= 0.64 \text{ mol/L}$ $[\text{PCl}_5] = 0.7295 \text{ mol/L}$ $= 0.73 \text{ mol/L}$

Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactants in the denominator each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 14.3$. This agrees closely with the given K_{eq} . The answers are reasonable and are expressed to the correct number of significant digits.

68. The synthesis of nitrogen monoxide, $\text{NO}(\text{g})$, is an equilibrium reaction, as shown below:



Initially, 1.33 mol of nitrogen gas and 2.80 mol of oxygen gas are placed into a 2.50 L container and allowed to reach equilibrium. What is the amount in moles of each gas at equilibrium at 200 K, if K_{eq} is 8.68 at this temperature?

What Is Required?

You need to determine the amount in moles of each reactant and product at equilibrium.

What Is Given?

You know the volume of the container: 2.50 L

You know the initial amount in moles of $\text{N}_2(\text{g})$: $n = 1.33$ mol

You know the initial amount in moles of $\text{O}_2(\text{g})$: $n = 2.80$ mol

You know the value of the equilibrium constant, K_{eq} : 8.68

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $\text{N}_2(\text{g})$ and $\text{O}_2(\text{g})$ in mol/L.	$[\text{N}_2] = \frac{1.33 \text{ mol}}{2.50 \text{ L}}$ $= 0.532 \text{ mol/L}$ $[\text{O}_2] = \frac{2.80 \text{ mol}}{2.50 \text{ L}}$ $= 1.12 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $\text{N}_2(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$\text{N}_2(\text{g})$	+	$\text{O}_2(\text{g})$	\rightleftharpoons	$2\text{NO}(\text{g})$
	$[\text{N}_2] \text{ (mol/L)}$		$[\text{O}_2] \text{ (mol/L)}$		$[\text{NO}] \text{ (mol/L)}$
I	0.532		1.12		0
C	$-x$		$-x$		$+2x$
E	$0.532 - x$		$1.12 - x$		$2x$

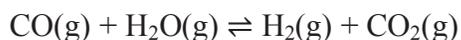
Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} .	$K_{\text{eq}} = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]}$ $8.68 = \frac{(2x)^2}{(0.532 - x)(1.12 - x)}$
Re-arrange the expression into a quadratic equation.	$8.68(0.59584 - 1.652x + x^2) = 4x^2$ $4.68x^2 - 14.339x + 5.1719 = 0$
Use the quadratic formula to solve for x . $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \frac{-(-14.339) \pm \sqrt{(-14.339)^2 - (4)(4.68)(5.1719)}}{2(4.68)}$ $= \frac{14.339 \pm 10.4302}{9.36}$ $x = 2.646 \text{ or } x = 0.4176$ Because it is impossible to have a negative concentration, x cannot be greater than 0.532. Therefore, $x = 0.4176$.
Determine the equilibrium concentration of each component in the gas phase.	At equilibrium: $[\text{N}_2] = 0.532 \text{ mol/L} - 0.4176 \text{ mol/L}$ $= 0.114 \text{ mol/L}$ $[\text{O}_2] = 1.12 \text{ mol/L} - 0.4176 \text{ mol/L}$ $= 0.70 \text{ mol/L}$ $[\text{NO}] = 2(0.4176)$ $= 0.8352 \text{ mol/L}$

<p>Use the formula $n = cV$ to calculate the amount in moles of each component.</p>	$n_{\text{N}_2} = cV$ $= (0.114 \text{ mol/L})(2.50 \text{ L})$ $= 0.286 \text{ mol}$ $n_{\text{O}_2} = cV$ $= (0.70 \text{ mol/L})(2.50 \text{ L})$ $= 1.76 \text{ mol}$ $n_{\text{NO}} = cV$ $= (0.8352 \text{ mol/L})(2.50 \text{ L})$ $= 2.088 \text{ mol}$ $= 2.09 \text{ mol}$
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Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression, $K_{\text{eq}} = 8.68$. This agrees with the given K_{eq} . The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

69. Carbon monoxide reaches equilibrium with water vapour as they react to form hydrogen gas and carbon dioxide gas. At 300 K, K_{eq} is equal to 0.52 for the reaction:



When 1.45 mol of carbon monoxide and 2.22 mol of water vapour are combined in a 1.00 L flask, equilibrium is established. What concentration of each reactant and product will be present at equilibrium?

What Is Required?

You need to determine the concentration of each reactant and product at equilibrium.

What Is Given?

You know the balanced chemical equation for the reaction.

You know the state of each component in the reaction.

You know the volume of the container: 1.00 L

You know the initial amount in moles of CO(g): $n = 1.45$ mol

You know the initial amount in moles of H₂O(g): $n = 2.22$ mol

You know the value of the equilibrium constant, K_{eq} : 0.52

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of CO(g) and H ₂ O(g) in mol/L.	[CO] = 1.45 mol/L [H ₂ O] = 2.22 mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration in of CO(g) had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

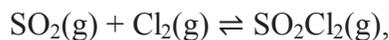
	CO(g)	+	H ₂ O(g)	\rightleftharpoons	H ₂ (g)	+	CO ₂ (g)
	[CO] (mol/L)		[H ₂ O] (mol/L)		[H ₂] (mol/L)		[CO ₂] (mol/L)
I	1.45		2.22		0		0
C	- x		- x		+ x		+ x
E	1.45- x		2.22- x		x		x

Write the expression for the equilibrium constant, K_{eq} .	$K_{\text{eq}} = \frac{[\text{H}_2][\text{CO}_2]}{[\text{CO}][\text{H}_2\text{O}]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} .	$K_{\text{eq}} = \frac{[\text{H}_2][\text{CO}_2]}{[\text{CO}][\text{H}_2\text{O}]}$ $0.52 = \frac{(x)(x)}{(1.45 - x)(2.22 - x)}$
Re-arrange the expression into a quadratic equation.	$0.52(3.219 - 3.67x + x^2) = x^2$ $0.48x^2 + 1.9084x - 1.6734 = 0$
Use the quadratic formula to solve for x . $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \frac{-(1.9084) \pm \sqrt{(1.9084)^2 - (4)(0.48)(-1.6734)}}{2(0.48)}$ $= \frac{-1.9084 \pm 2.6182}{0.96}$ $x = 0.7394 \text{ or } x = -4.7152$ Because it is impossible to have a negative concentration, $x = 0.7394$.
Determine the equilibrium concentration of each component in the gas phase.	At equilibrium: $[\text{H}_2] = 1.45 \text{ mol/L} - 0.7394 \text{ mol/L}$ $= 0.7106 \text{ mol/L}$ $= 0.71 \text{ mol/L}$ $[\text{CO}_2] = 2.22 \text{ mol/L} - 0.7394 \text{ mol/L}$ $= 1.4806 \text{ mol/L}$ $= 1.5 \text{ mol/L}$ $[\text{CO}] = [\text{H}_2\text{O}]$ $= 0.7394 \text{ mol/L}$ $= 0.74 \text{ mol/L}$

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{\text{eq}} = 0.52$. This agrees with the given K_{eq} . The answers are reasonable and are expressed to the correct number of significant digits.

70. Sulfuryl chloride, $\text{SO}_2\text{Cl}_2(\text{g})$ is used in a process to produce pesticides. It is also used as a convenient source of chlorine gas, because it can be cooled to form a pourable liquid that is easier to store and dispense. To generate this compound, the following reaction can be used, with a K_{eq} value of 0.45 at 650 K:



Initially, 3.15 mol $\text{SO}_2(\text{g})$ and 2.14 mol $\text{Cl}_2(\text{g})$ are placed into a 1.00 L container and allowed to reach equilibrium. What are the equilibrium concentrations of all these substances?

What Is Required?

You need to determine the concentration of each reactant and product at equilibrium.

What Is Given?

You know the balanced equation for the reaction.

You know the state of each component in the reaction.

You know the volume of the container: 1.00 L

You know the initial amount in moles of $\text{SO}_2(\text{g})$: $n = 3.15$ mol

You know the initial amount in moles of $\text{Cl}_2(\text{g})$: $n = 2.14$ mol

You know the value of the equilibrium constant, K_{eq} : 0.45

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $\text{SO}_2(\text{g})$ and $\text{Cl}_2(\text{g})$ in mol/L.	$[\text{SO}_2] = 3.15$ mol/L $[\text{Cl}_2] = 2.14$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $\text{SO}_2(\text{g})$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentration of each component.	See the ICE table below.

	$\text{SO}_2(\text{g})$	+	$\text{Cl}_2(\text{g})$	\rightleftharpoons	$\text{SO}_2\text{Cl}_2(\text{g})$ (mol/L)
	$[\text{SO}_2]$ (mol/L)		$[\text{Cl}_2]$ (mol/L)		$[\text{SO}_2\text{Cl}_2]$
I	3.15		2.14		0
C	$-x$		$-x$		$+x$
E	$3.15 - x$		$2.14 - x$		x