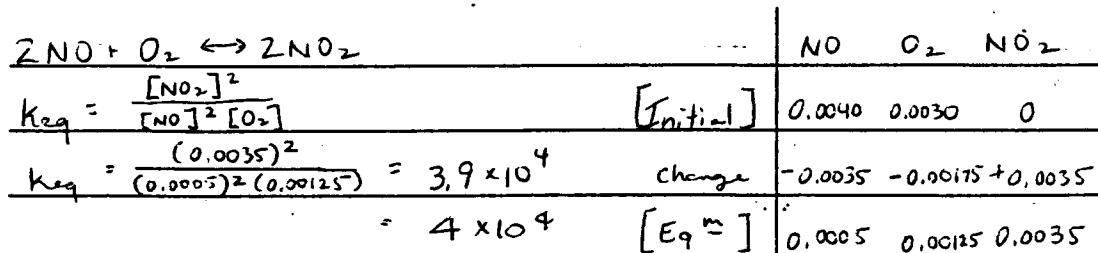
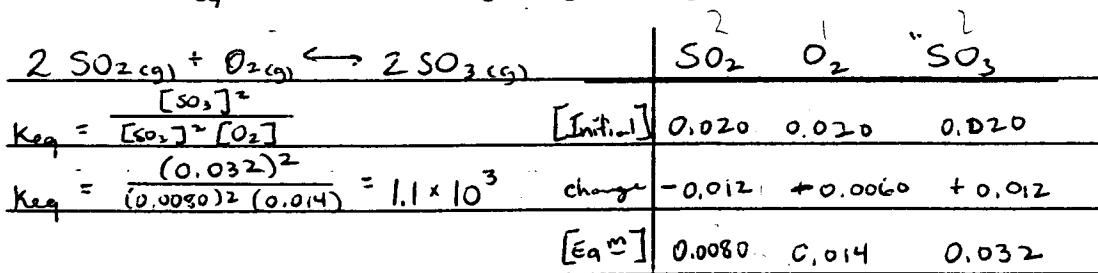


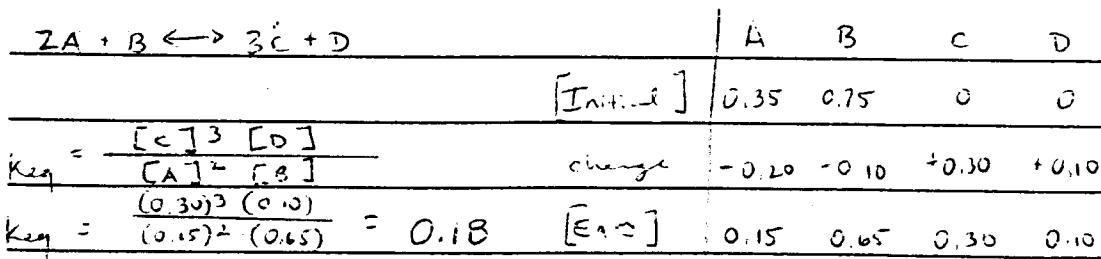
1. 0.0040 mol of NO and 0.0030 mol of O₂ are introduced into a 1.0 L flask, and the reaction $2\text{NO(g)} + \text{O}_2\text{(g)} \rightleftharpoons 2\text{NO}_2\text{(g)}$ occurs. At equilibrium, it is determined that [NO₂] = 0.0035 mol/L. What is the value of K_{eq} for the reaction?



2. 0.020 mol of each of SO₂, O₂, and SO₃ is placed in a 1.0 L flask and allowed to come to equilibrium. The equilibrium [SO₃] is found to be 0.0080 mol/L. What is the value of K_{eq} for the reaction $2\text{SO}_2\text{(g)} + \text{O}_2\text{(g)} \rightleftharpoons 2\text{SO}_3\text{(g)}$?



3. A solution initially contains 0.35M of A and 0.75M of B. A reaction occurs according to the equation $2\text{A} + \text{B} \rightleftharpoons 3\text{C} + \text{D}$. At equilibrium, [D] is found to be 0.10M. What is the value of K_{eq}?



4. For the equilibrium reaction $\text{CO}(g) + \text{H}_2\text{O}(g) \rightleftharpoons \text{CO}_2(g) + \text{H}_2(g)$, the K_{eq} value at 690°C is 10.0. A mixture of 0.300 mol of CO, 0.300 mol of H₂O, 0.500 mol of CO₂, and 0.500 mol of H₂ is placed in a 1.0 L flask.

Key.

- Show that the reaction is not at equilibrium.
- Determine the direction in which the reaction will shift to reach equilibrium.
- Calculate the equilibrium concentrations of all four species.

a) trial product = $\frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]} = \frac{(0.500)(0.500)}{(0.300)(0.300)} = 2.8 \neq 10.0 \therefore \text{not at eqm}$

b) trial product < K_{eq} ∴ shifts to products

$K_{eq} = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}$	CO	H ₂ O	CO ₂	H ₂
$K_{eq} = \frac{(0.500+x)^2}{(0.300-x)^2} = 10.0$	[Initial]	0.300	0.300	0.500
	Change	-x	-x	+x
$x = 0.108 \text{ M}$	[Eqm]	0.300-x	0.300-x	0.500+x

$$[\text{CO}] = [\text{H}_2\text{O}] = 0.300 - 0.108 = 0.192 \text{ M} \quad [\text{CO}_2] = [\text{H}_2] = 0.500 + 0.108 = 0.608 \text{ M}$$

5. The K_{eq} for the reaction $2\text{HI}(g) \rightleftharpoons \text{H}_2(g) + \text{I}_2(g)$ has a value of 1.85×10^{-2} at 425°C. If 0.18 mol of HI is placed in a 2.0 L flask and allowed to come to equilibrium at this temperature, what will be the equilibrium [I₂]?

$K_{eq} = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = 1.85 \times 10^{-2}$	HI	H ₂	I ₂
$[\text{Initial}]$	$\frac{0.18 \text{ mol}}{2.0 \text{ L}}$	0	0
$= \frac{x^2}{(0.090-2x)^2} = 1.85 \times 10^{-2}$	Change	-2x	+x
$\sqrt{\text{both sides}}$	[Eqm]	0.090-2x	x
$x = 9.6 \times 10^{-3} \text{ M}$			

$$[\text{I}_2] = 9.6 \times 10^{-3} \text{ M}$$

6. For the reaction $2\text{A}(g) \rightleftharpoons 2\text{B}(g) + \text{C(s)}$, the value of K_{eq} is known to be 6.8×10^2 . If 0.42 mol of A is placed in a 3.0 L container and allowed to reach equilibrium, what is the equilibrium [B]?

$K_{eq} = \frac{[\text{B}]^2}{[\text{A}]^2} = 6.8 \times 10^2$	A	B
$\frac{0.14^2}{3.0}$	0.14	0
$K_{eq} = \frac{x^2}{(0.14-x)^2} = 6.8 \times 10^2$	Change	-x
$x = 0.13 \text{ M}$		0.14-x

$$[\text{B}]_{eq} = 0.13 \text{ M}$$