

# Solving Problems with a Small $K_{eq}$

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**\* Some equilibrium expressions may require you to use the quadratic equation to solve**

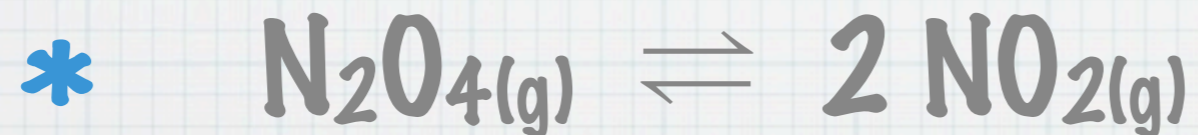
# Using Quadratic Equation

\* Reminder

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

# Example

\* If initial concentration of  $\text{N}_2\text{O}_4$  is 0.50 M, what are the equilibrium concentrations if  $K_{\text{eq}}$  is 0.0059?



# Solution

- \* Fill out an ICE table. Write a chemical equation and add know values.



Initial	0.50	0
Change	-x	+2x
Equilibrium	0.50 - x	+2x

# Solution

\* Now use the  $K_{eq}$  expression to solve

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$
$$0.0059 = \frac{(2x)^2}{(0.5 - x)}$$

Since the equation doesn't involve a perfect square you must use the quadratic formula

# Solution

\* Rearrange into the quadratic equation

$$0.0059 (0.5 - x) = (2x)^2$$

$$0.00295 - 0.0059x = 4x^2$$

$$-4x^2 - 0.0059x + 0.00295 = 0$$

$$-1(4x^2 + 0.0059x - 0.00295) = 0$$

This mean  $a = 4$ ,  $b = 0.0058$ ,  $c = -0.00295$

# Solution

\* Now substitute into the quadratic equation:

$$x = \frac{-0.0058 \pm \sqrt{[(0.0058)^2 - 4(4)(-0.00295)]}}{2(4)}$$

$$x = 0.026$$

OR

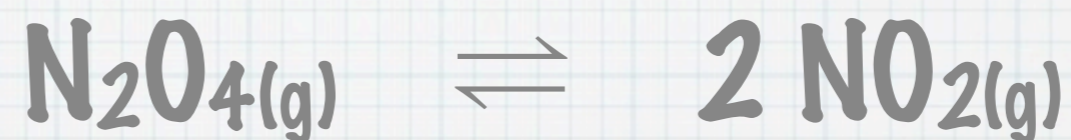
$$x = -0.0029$$

Since  $x$  can't be negative,  $x = 0.026$



# Solution

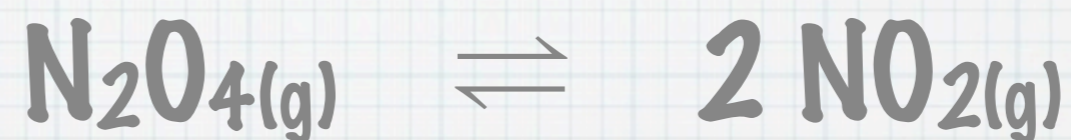
- \* Fill out an ICE table. Write a chemical equation and add know values.



Initial	0.50	0
Change	-x	+2x
Equilibrium	0.50 - 0.026	+2 (0.026)

# Solution

- \* Fill out an ICE table. Write a chemical equation and add know values.



Initial	0.50	0
Change	-x	+2x
Equilibrium	0.474	0.052

# Approximating Small Quantities

- \* When dealing with small concentration changes, you can sometimes assume that initial and final concentrations are the same.

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Example:

$$[I]_i = 0.065$$

$$\text{Change is } [I] = -0.000032$$

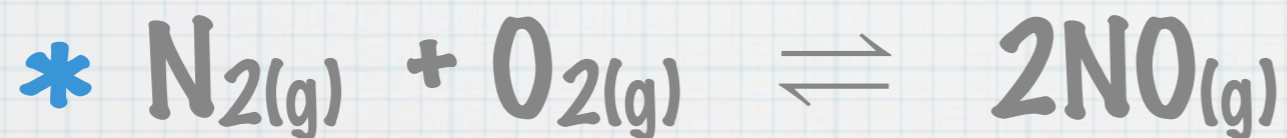
$$[I]_{eq} = 0.065 - 0.000032 = 0.064968$$

# Approximating Small Quantities

- \* “hundred rule” is an assumption made to simplify problems
- \* If the ratio of [initial concentration of reactant]/ $K > 100$ , then  $x$  is very small compared to initial concentration and may be discarded from some calculations.

# Example

\* In a reaction  $K$  is  $4.2 \times 10^{-8}$  in the equilibrium constant. A chemist puts 0.085 mol of nitrogen and 0.038 mol into a 1.5 L container. Determine the equilibrium concentration of NO.



# Solution

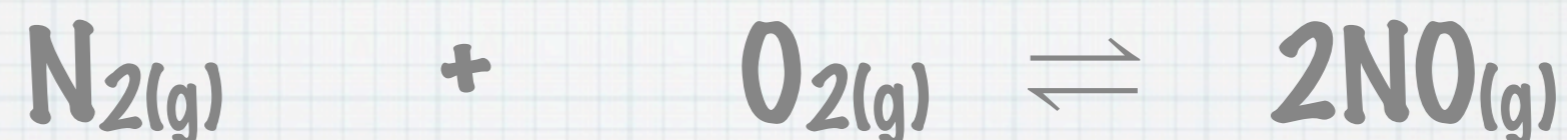
\* First calculate the concentration of each product and reactant:

$$[\text{N}_3] = \frac{n}{V} = \frac{0.085 \text{ mol}}{1.5\text{L}} = 0.057 \text{ mol/L}$$

$$[\text{O}_2] = \frac{n}{V} = \frac{0.038 \text{ mol}}{1.5\text{L}} = 0.025 \text{ mol/L}$$

# Solution

- \* Set up an ICE table and record information



Initial	0.057	0.025	0
Change	-x	-x	+2x
Equilibrium	0.057-x	0.025 - x	2x



# Solution

- \* Check to see if “hundred rule applies”
- \*  $[I]_i/K > 100$
- \*  $0.57/(4.2 \times 10^{-8}) = >100$

This means you can assume the concentrations will remain approximately the same

# Solution

\* Now substitute and solve

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

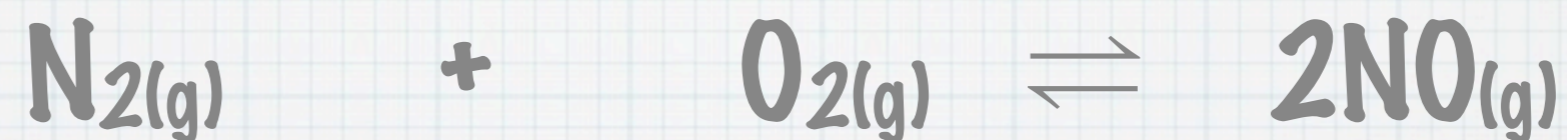
$$4.2 \times 10^{-8} = \frac{(2x)^2}{(0.057)(0.025)}$$

$$4x^2 = 5.985 \times 10^{-11}$$

$$x = 3.9 \times 10^{-6}$$

# Solution

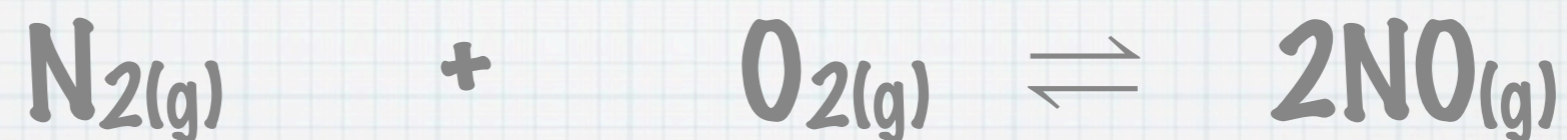
- \* Set up an ICE table and record information



Initial	0.057	0.025	0
Change	-x	-x	+2x
Equilibrium	0.057-x	0.025 - x	2(3.9 x 10 <sup>-6</sup> )

# Solution

- \* Set up an ICE table and record information



Initial	0.057	0.025	0
Change	-x	-x	+2x
Equilibrium	0.057-x	0.025 - x	$7.8 \times 10^{-6}$

# Homework

\* p 457 #71, 72