EXAMPLE C

A certain amount of $NO_2(g)$ was introduced into a 5.00 L bulb. When equilibrium was attained according to the equation

$2\ NO(g) + O_2(g) \rightleftharpoons 2\ NO_2(g)$,

the concentration of NO(g) was 0.800 M. If $\boxed{K_{EQ} \text{ has a value of } 24.0}$, how many moles of NO_2 were originally put into the bulb?

1. Write out Keq expression : $Keq = \dfrac{[NO_2]^2}{[NO]^2[O_2]}$

2. B/c "TIME PASSES" AN ICE TABLE IS NEEDED!

	2 NO	+ O_2	$\rightleftharpoons$ 2 NO_2
I	Ø	Ø	X
C	+0.800	+0.400	−0.800
E	0.800	0.400	X−0.800

3. Plug into keq and solve for "X"

$$24.0 = \frac{(X-0.800)^2}{(0.800)^2(0.400)} \times (0.800)^2(0.400)$$

$$\sqrt{(0.800)^2(0.400)24.0} = \sqrt{(X-0.800)^2}$$

$$2.479 + 0.800 = X - 0.800 + 0.800$$

$$\boxed{3.279 = X}$$ n/e # moles

3.279 mol / L × 5.00 L = $\boxed{16.4 \text{ mol } NO_2}$

EXAMPLE D

$K_{EQ} = 49$ for $2\ NO(g) + O_2(g) \rightleftharpoons 2\ NO_2(g)$.

If we introduce 2.0 mol of NO(g), 0.20 mol of $O_2(g)$ and 0.40 mol of $NO_2(g)$ into a 2.0 L bulb, which way will the reaction shift in order to reach equilibrium?

"TRIAL Keq"

TRIAL Keq Question: "$Q = \dfrac{[NO_2]^2}{[NO]^2[O_2]}$"

If $Q = keq$ system is in equilibrium

$Q > keq$ system must shift to REACTANTS to reach equilibrium

$Q < keq$ system must shift to PRODUCTS

$$Q = \frac{[0.20]^2}{[1.0]^2[0.10]} = 0.40$$

B/c $Q < keq$

$0.40 < 49$

The rxn must shift to PRODUCTS to reach Equilibrium!

SEAT WORK/HOMEWORK: Exercises 47- 54 pgs 70-71

PLO's: F5,F6 and F7

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47−57 (except 55)