

**EXAMPLE C**

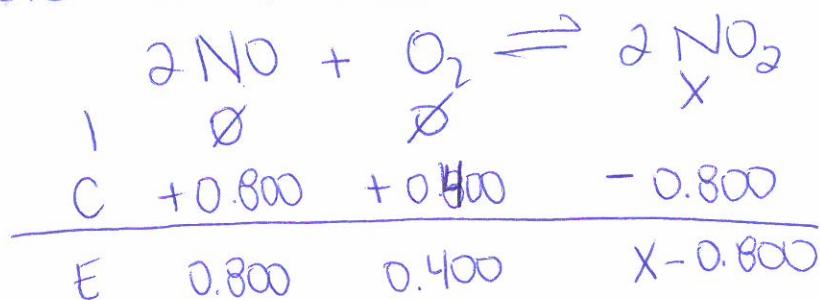
A certain amount of  $\text{NO}_2(g)$  was introduced into a 5.00 L bulb. When equilibrium was attained according to the equation



the concentration of  $\text{NO}(g)$  was 0.800 M. If  $K_{\text{eq}}$  has a value of 24.0, how many moles of  $\text{NO}_2$  were originally put into the bulb?

1. Write out  $K_{\text{eq}}$  expression :  $K_{\text{eq}} = \frac{[\text{NO}_2]^2}{[\text{NO}]^2[\text{O}_2]}$

2. Blc "TIME PASSES" AN ICE TABLE IS NEEDED!



3. Plug into  $K_{\text{eq}}$  and solve for "x"

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

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

$$2.479 + 0.800 = \frac{x-0.800 + 0.800}{13.279 = x}$$

h/e #mols

$\text{O}_2$	2 mol
$\text{NO}_2$	16.4 mol

**EXAMPLE D**

$$K_{\text{eq}} = 49 \text{ for } 2 \text{NO}(g) + \text{O}_2(g) \rightleftharpoons 2 \text{NO}_2(g)$$

If we introduce 2.0 mol of  $\text{NO}(g)$ , 0.20 mol of  $\text{O}_2(g)$  and 0.40 mol of  $\text{NO}_2(g)$  into a 2.0 L bulb, which way will the reaction shift in order to reach equilibrium?

TRIAL K<sub>eq</sub> Question: " $Q = \frac{[\text{NO}_2]^2}{[\text{NO}]^2[\text{O}_2]}$ "

If  $Q = K_{\text{eq}}$  system is in equilibrium

$Q > K_{\text{eq}}$  system must shift to REACTANTS to reach equil.

$Q < K_{\text{eq}}$  system must shift to PRODUCTS to reach equil.

Blc  $Q < K_{\text{eq}}$

$$0.40 < 49$$

The rxn must shift to PRODUCTS to reach equilibrium!

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

PLO's: F5, F6 and F7