CHAPTER 16	Using Stoichiometry to Calculate K_c	BLM 16.3.2
OVERHEAD		

Problem

Hydrogen gas is considered to be a non-polluting, sustainable fuel source. It can be used in internal combustion engines and in fuel cells to produce heat and electricity. Hydrogen gas can be produced in several different ways. The following reaction illustrates one method using water and carbon monoxide:

$CO(g) + H_2O(g) \leftrightarrow H_2(g) + CO_2(g)$

The reaction has been studied at different temperatures to find the optimum conditions. At 700 K, the equilibrium constant is 0.83. Suppose that you start with 1.0 mol of CO(g) and 1.0 mol of $H_2O(g)$ in a 5.0 L container. What amount of each substance will be present in the container when the gases are in equilibrium at 700 K?

What Is Required?

You need to find the amount (in moles) of CO(g), $H_2O(g)$, $H_2(g)$, and $CO_2(g)$ at equilibrium.

What Is Given?

You have the chemical equation. You know the initial amount of each gas, the volume of the container, and the equilibrium constant.

Plan Your Strategy

Step 1 Calculate the initial concentrations.

Step 2 Set up an ICE table. Record the initial concentrations you calculated in Step 1 in your ICE table. Let the change in molar concentrations of the reactants be *x*. Use the stoichiometry of the chemical equation to write and record expressions for the equilibrium concentrations.

Step 3 Write the equilibrium expression. Substitute the expressions for the equilibrium concentrations into the expression for K_c. Solve the equilibrium expression for *x*.

Step 4 Calculate the equilibrium concentration of each gas. Then use the volume of the container to find the amount of each gas present.

Act on Your Strategy

Step 1 The initial amount of CO(g) is equal to the initial amount of $H_2O(g)$:

$$[CO(g)] = [H_2O(g)] = \frac{1.0\text{ mol}}{5.0\text{ L}} = 0.20\text{ mol/L}$$

Step 2 Set up an ICE table.

$CO(g) + H_2O(g) \leftrightarrow H_2(g) + CO_2(g)$

	[CO(g)] (mol/L)	[$H_2O(g)$] (mol/L)	[$H_2(g)$] (mol/L)	[$CO_2(g)$] (mol/L)
Initial	0.20	0.20	0	0
Change	$-x$	$-x$	$+x$	$+x$
Equilibrium	$0.20 - x$	$0.20 - x$	x	x

Note that change in concentration for the reactants is negative because they are being used up in the reaction, while change in concentration for the products is positive because products are being created.

Using Stoichiometry to Calculate K_c (continued)

Step 3 $K_c = \dfrac{[H_2][CO_2]}{[CO][H_2O]}$

$$0.83 = \frac{(x)(x)}{(0.20 - x)(0.20 - x)} = \frac{(x)^2}{(0.20 - x)^2}$$

$$0.9110 = \frac{x}{0.20 - x}$$

$$x = 0.095\ 34$$

Step 4 The concentrations of the reactants at equilibrium are as follows:

$[H_2(g)] = [CO_2(g)] \approx 0.095\ 34$ mol/L
$[CO(g)] = [H_2O(g)] \approx 0.2000 - 0.095\ 34 = 0.104\ 66$ mol/L

Round to two significant digits:
$[H_2(g)] = [CO_2(g)] \approx 0.095$ mol/L
$[CO(g)] = [H_2O(g)] \approx 0.10$ mol/L

To find the amount of each gas, multiply the concentration of each gas by the volume of the container (5.0 L).
Amount of $H_2(g) = CO_2(g) = 0.48$ mol
Amount of $CO(g) = H_2O(g) = 0.50$ mol

Check Your Solution

The equilibrium expression has product concentrations in the numerator and reactant concentrations in the denominator. The concentration of each chemical present at equilibrium is given in mol/L. Check K_c by substituting these concentrations back into the equation.

$$K_c = \frac{(0.095\ 343)^2}{(0.104\ 66)^2} = 0.8298 \approx 0.83$$

MathTip
FYI

In this problem, the right side of the equilibrium expression is a perfect square in Step 3. Noticing perfect squares, then taking the square root of both sides makes solving the equation easier. Also, when you use an answer to one part of a problem as data in another calculation, always use the unrounded value. If you round numbers at every step, the error will become quite large. For example, if you had rounded every step in the solution to this Sample Problem to two significant figures, your answer for K_c would have been 0.75, which is an error of about 10%.