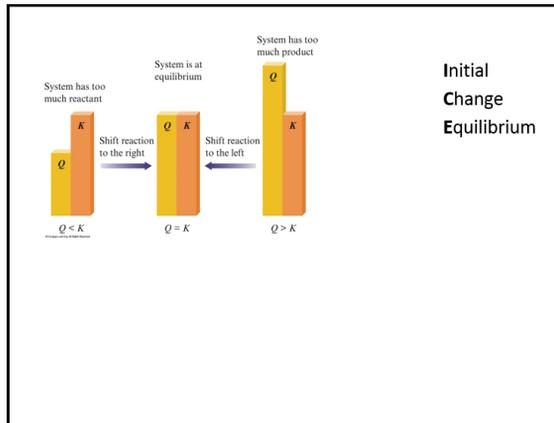


**Using ICE Charts to Calculate Equilibrium Quantities and Constants**

Solving Equilibrium Problems

- 1) Write the balanced equation for the reaction.
- 2) Write the equilibrium expression using the law of mass action.
- 3) List the initial concentrations.
- 4) Calculate Q, and determine the direction of the shift to equilibrium.
- 5) Define the change needed to reach equilibrium, and define the equilibrium concentrations by applying the change to the initial concentrations.
- 6) Substitute the equilibrium concentrations into the equilibrium expression, and solve for the unknown.
- 7) Check your calculated equilibrium concentrations by making sure they give the correct value of K.



Initial  
Change  
Equilibrium

Example 1. 0.500 mol SO<sub>2</sub> and 0.300 mol O<sub>2</sub> are placed into an evacuated 5.0 L vessel at 398 K. Once equilibrium is established, 0.100 mol of SO<sub>3</sub> are in the vessel.

a) What are the equilibrium concentrations of SO<sub>2</sub> and O<sub>2</sub>?

$$2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \leftrightarrow 2 \text{SO}_3(\text{g}) \quad K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]}$$

I	0.500 mol	0.300 mol	0 mol
C	-2x	-x	+2x
E	0.500 mol - 2x	0.300 mol - x	2x = 0.100 mol

∴ x = 0.050 mol

0.500 mol - 0.100 mol = 0.400 mol SO<sub>2</sub>  
0.300 mol - 0.050 mol = 0.250 mol O<sub>2</sub>

$[\text{SO}_2] = \frac{0.400 \text{ mol SO}_2}{5.0 \text{ L}} = 0.0800 \text{ M SO}_2$   
 $[\text{O}_2] = \frac{0.250 \text{ mol O}_2}{5.0 \text{ L}} = 0.0500 \text{ M O}_2$

Example 1. 0.500 mol SO<sub>2</sub> and 0.300 mol O<sub>2</sub> are placed into an evacuated 5.0 L vessel at 398 K. Once equilibrium is established, 0.100 mol of SO<sub>3</sub> are in the vessel.

b) What is the value of K<sub>c</sub> at 398 K?

$$K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]} = \frac{(0.100 \text{ mol} / 5.0 \text{ L})^2}{(0.0800)^2 (0.0500)} = 1.25$$

c) What is the value of K<sub>p</sub> at 398 K?

$$K_p = K_c (RT)^{\Delta n}$$

$$K_p = (1.25) (0.0821 \times 398 \text{ K})^{2-3} = 0.0383$$

Example 2. A reaction starts with 5.00 mol each of SbCl<sub>3</sub> and Cl<sub>2</sub> in a 2.00 L flask at 350 °C. When equilibrium is established at 350 °C, the moles of SbCl<sub>5</sub> present are 1.67 mol. What is the K<sub>c</sub> for this rxn?

$$\text{SbCl}_3(\text{g}) + \text{Cl}_2(\text{g}) \leftrightarrow \text{SbCl}_5(\text{g}) \quad K_c = \frac{[\text{SbCl}_5]}{[\text{SbCl}_3][\text{Cl}_2]}$$

I	5.00 mol	5.00 mol	0 mol
C	-x	-x	+x
E	5.00 mol - x	5.00 mol - x	x = 1.67 mol

∴ x = 1.67 mol

5.00 mol - 1.67 mol = 3.33 mol SbCl<sub>3</sub>  
5.00 mol - 1.67 mol = 3.33 mol Cl<sub>2</sub>

$[\text{SbCl}_3] = [\text{Cl}_2] = \frac{3.33 \text{ mol}}{2.00 \text{ L}} = 1.665 \text{ M SbCl}_3 \text{ \& Cl}_2$   
 $[\text{SbCl}_5] = \frac{1.67 \text{ mol}}{2.00 \text{ L}} = 0.835 \text{ M SbCl}_5$

$$K_c = \frac{0.835}{(1.665)(1.665)} = 0.301$$

Example 3. H<sub>2</sub>(g), I<sub>2</sub>(g), and HI(g) are placed into a 5.25 L container at 698 K. The initial partial pressure of H<sub>2</sub>(g) is 3.27 atm, I<sub>2</sub> is 3.27 atm and HI is 32.7 atm. The K<sub>c</sub> value of the reaction is 54.3 at 698 K.

$$\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \leftrightarrow 2 \text{HI}(\text{g}) \quad K_c = 54.3$$

a) What is the value of K<sub>p</sub> at 698K?

$$K_p = K_c (RT)^{\Delta n} = 54.3 (0.0821 \times 698 \text{ K})^{2-2} = 54.3$$

b) What is the equilibrium partial pressure of each gas in the reaction vessel?

$$\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \leftrightarrow 2 \text{HI}(\text{g})$$

I	3.27 atm	3.27 atm	32.7 atm
C	+x	+x	-2x
E	3.27 atm + x	3.27 atm + x	32.7 atm - 2x

$Q = (P_{\text{HI}})^2 / (P_{\text{H}_2})(P_{\text{I}_2}) = (32.7)^2 / (3.27)(3.27) = 54.3 = K_c$

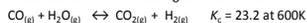
$K_p = (P_{\text{HI}})^2 / (P_{\text{H}_2})(P_{\text{I}_2}) = 54.3$

$3.27 \text{ atm} + 0.918 \text{ atm} = 4.19 \text{ atm} = P_{\text{H}_2} = P_{\text{I}_2}$   
 $P_{\text{HI}} = 30.4 \text{ atm}$

$24.096 \text{ atm} + 7.369 \text{ atm} = 32.7 \text{ atm} < Q > K_p$   
∴ Reverse shift

$X = 0.918 \text{ atm}$

**Example 4.** 0.100 moles of each reactant and product species are placed into a 5.00 L flask, equilibrium is established in the following reaction at 600 K:



a) What is the equilibrium concentration of  $\text{H}_2$ ?

	$\text{CO}_{(g)}$	$+$	$\text{H}_2\text{O}_{(g)}$	$\rightleftharpoons$	$\text{CO}_{2(g)}$	$+$	$\text{H}_2_{(g)}$	
I	0.100 mol		0.100 mol		0.100 mol		0.100 mol	
C	-x		-x		+x		+x	
E	0.100 mol - x		0.100 mol - x		0.100 + x		0.100 mol + x	

$Q = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}$   
 $Q = \frac{(0.100/5)(0.100/5)}{(0.100/5)(0.100/5)}$

$$K_c = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}$$

$$23.2 = \frac{(0.100+x)^2}{(0.100-x)^2}$$

$$4.817 = \frac{0.100+x}{0.100-x}$$

$$0.4817 - 4.817x = 0.100 + x \quad Q = 1$$

$$0.3817 = 5.817x \quad Q < K_c$$

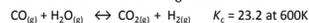
$\therefore$  forward shift

$$x = 0.0656 \text{ mol}$$

$\therefore [\text{CO}] = [\text{H}_2\text{O}] = \frac{0.100 \text{ mol} - 0.0656 \text{ mol}}{5.00 \text{ L}} = 6.88 \times 10^{-3} \text{ M}$

$[\text{CO}_2] = [\text{H}_2] = \frac{0.100 \text{ mol} + 0.0656 \text{ mol}}{5.00 \text{ L}} = 3.31 \times 10^{-2} \text{ M}$

**Example 4.** 0.100 moles of each reactant and product species are placed into a 5.00 L flask, equilibrium is established in the following reaction at 600 K:



b) Once equilibrium has been established, 0.0150 moles of both  $\text{CO}_2(g)$  and  $\text{H}_2(g)$  are added to the reaction vessel.

- How will the concentration of  $\text{CO}(g)$  change in response to this addition? Explain your reasoning.
- Calculate the concentration of  $\text{CO}(g)$  once equilibrium has been reestablished after this addition.

## GE WS Part II

- Carbon dioxide gas reacts with solid carbon to give carbon monoxide. At 700. °C and equilibrium a 5.0 L flask contains 0.27 moles of carbon monoxide, 0.58 moles of carbon dioxide, and 0.44 moles of carbon.
  - Write a balanced equation for the reaction of one mole of carbon dioxide with carbon.
  - Calculate  $K_c$  for the reaction at 700. °C.
  - Calculate  $K_p$  for the reaction at 700. °C.