Marks 5

- Consider the following reaction.

$$SO_2(g) + NO_2(g) \rightleftharpoons NO(g) + SO_3(g)$$

At 460 °C this reaction has a value of K_c = 85.0. Suppose 0.100 mol of SO_2, 0.0600 mol of NO_2, 0.0800 mol of NO and 0.120 mol of SO_3 are placed in a 10.0 L container at this temperature. What are the concentrations of all of the gases when the system reaches equilibrium?

The initial concentrations are:

$[SO_2(g)]$ = number of moles / volume = (0.100 mol) / (10.0 L) = 0.0100 M

$[NO_2(g)]$ = (0.0600 mol) / (10.0 L) = 0.00600 M

[NO(g)] = (0.0800 mol) / (10.0 L) = 0.00800 M

$[SO_3(g)]$ = (0.120 mol) / (10.0 L) = 0.0120 M

The reaction quotient can be used to predict the direction that the reaction will shift:

$$Q = \frac{[NO(g)][SO_3(g)]}{[SO_2(g)][NO_2(g)]} = \frac{(0.0120)(0.00800)}{(0.0100)(0.00600)} = 1.6$$

As $Q < K$, the reaction will shift to the right – to increase the amount of products and decrease the amount of reactants. The reaction table is then:

	$SO_2(g)$	$NO_2(g)$	$\rightleftharpoons$	NO(g)	$SO_3(g)$
initial	**0.0100**	**0.00600**		**0.00800**	**0.0120**
change	$-x$	$-x$		$+x$	$+x$
equilibrium	$0.0100 - x$	$0.00600 - x$		$0.00800 + x$	$0.0120 + x$

Hence,

$$K = \frac{(0.00800+x)(0.0120+x)}{(0.0100-x)(0.00600-x)} = 85.0$$

$$85.0(x^2 - 0.01600x + 0.0000600) = x^2 + 0.02000x + 0.000096$$

$$84.0x^2 - 1.38x + 0.005004 = 0$$

Solving this quadratic equation gives x = 0.0054 and 0.011. The second root is not possible, as it leads to negative concentrations for the reactants.

Using x = 0.0054 M gives,

$[SO_2(g)]$ = (0.0100 – 0.0054) M = 0.00460 M

$[NO_2(g)]$ = (0.00600 – 0.0054) M = 0.000597 M

[NO(g)] = (0.00800 + 0.0054) M = 0.0134 M

$[SO_3(g)]$ = (0.0120 + 0.0054) M = 0.0174 M

$[SO_2(g)]$ = **0.00460 M**	$[NO_2(g)]$ = **0.000597 M**
$[SO_3(g)]$ = **0.0174 M**	[NO(g)] = **0.0134 M**