Marks 4

- Consider the following reaction.

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

An equilibrium mixture in a 1.00 L vessel was found to contain $[SO_2(g)] = 0.800$ M, $[NO_2(g)] = 0.100$ M, $[SO_3(g)] = 0.600$ M and $[NO(g)] = 0.400$ M. If the volume and temperature are kept constant, what amount (in mol) of NO(g) needs to be added to the reaction vessel to give an equilibrium concentration of $NO_2(g)$ of 0.300 M?

From the chemical equation,

$$K_{eq} = \frac{[SO_3(g)][NO(g)]}{[SO_2(g)][NO_2(g)]}$$

As the original mixture is at equilibrium:

$$K_{eq} = \frac{[SO_3(g)][NO(g)]}{[SO_2(g)][NO_2(g)]} = \frac{(0.600)(0.400)}{(0.800)(0.100)} = 3.00$$

This equilibrium is now disturbed by the addition of *x* M of NO(g). To re-establish equilibrium, the reaction will shift to the left by an unknown amount *y*. The reaction table for this is:

	$SO_2(g)$	$NO_2(g)$		$SO_3(g)$	$NO(g)$
initial	**0.800**	**0.100**	$\rightleftharpoons$	**0.600**	$0.400 + x$
change	$+y$	$+y$		$-y$	$-y$
equilibrium	$0.800 + y$	$0.100 + y$		$0.600 - y$	$0.400 + x - y$

As $[NO_2(g)] = 0.300$ M at the new equilibrium, $y = (0.300 - 0.100)$ M $= 0.200$ M. Hence, the new equilibrium concentrations are:

$[SO_2(g)] = (0.800 + 0.200)$ M $= 1.000$ M
$[NO_2(g)] = 0.300$ M
$[SO_3(g)] = (0.600 - 0.200)$ M $= 0.400$ M
$[NO(g)] = (0.400 + x - 0.200)$ M $= (0.200 + x)$ M

As the system is at equilibrium,

$$K_{eq} = \frac{[SO_3(g)][NO(g)]}{[SO_2(g)][NO_2(g)]} = \frac{(0.400)(0.200+x)}{(1.000)(0.300)} = 3.00$$

Solving this gives $x = 2.05$ M. As the reaction is carried out in a 1.00 L container, this is also the number of moles required.

Answer: **2.05 mol**