Ex 2: The following gas phase reaction has an equilibrium constant, K_c, of 4.00×10^2 at 20°C. The reaction is started with 1.000 M of each product. What are the equilibrium concentrations of all species?

$$2\,AB \rightleftharpoons A_2 + B_2$$

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	$2\,AB$	$\rightleftharpoons$	A_2	+	B_2
I	0		1.000		1.000
C	+ 2x		- x		- x
E	2x		1.000 - x		1.000 - x

$$K_c = \frac{[A_2]\,[B_2]}{[AB]^2} = \frac{(1.000 - x)^2}{(2x)^2} = 4.00 \times 10^2$$

Even though the reaction is going from right to left (in the reverse direction) the equil. constant given in the problem is for the reaction as written and for the form of K I've given based on the reaction.

Since the K is big in the forward direction, this reaction isn't going very far to the left. That means the "x" is small compared to 1.000 so you might be able to ignore it in the numerator. However, you don't have to. The left side is a perfect square.

Take the square root of both sides:

$$\frac{(1.000 - x)}{2x} = 20.0$$

$$1.000 - x = (40.0)\,x$$

$$x = 0.024\underline{\mathbf{3}}9$$

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$[A_2] = [B_2] = 1.000 - x = 1.000 - 0.024\underline{\mathbf{3}}9 = 0.976$ M
$[AB] = 2x = 2(0.024\underline{\mathbf{3}}9) = 0.0488$ M

What if you don't like having the "-" on the right and going from right to left? You can reverse the reaction and then use K_{rev}.

	A_2	+	B_2	$\rightleftharpoons$	2 AB
I	1.000		1.000		0
C	- x		- x		+ 2x
E	1.000 - x		1.000 - x		2x

$$K_{c,rev} = \frac{[AB]^2}{[A_2]\,[B_2]} = \frac{(2x)^2}{(1.000 - x)^2} = \frac{1}{K_{c,for}} = \frac{1}{4.00 \times 10^2}$$

$$K_{c,rev} = \frac{(2x)^2}{(1.000 - x)^2} = 2.50 \times 10^{-3}$$

Take the square root of both sides:

$$K_{c,rev} = \frac{(2x)}{(1.000 - x)} = 5.00 \times 10^{-2}$$

$$1.000 - x = (40.0)\,x$$

$x = 0.024\underline{\mathbf{3}}9$ Same as above