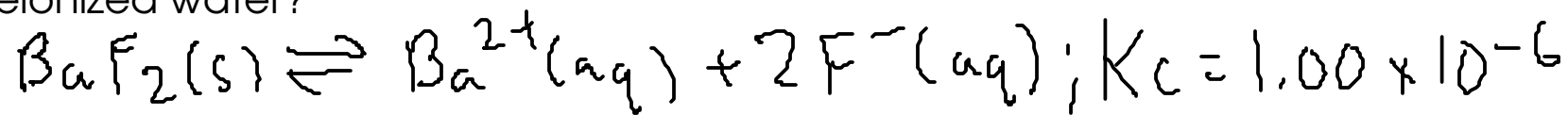What is the concentration of barium ion at equilibrium if solid barium fluoride is mixed with deionized water?

$$BaF_2(s) \rightleftharpoons Ba^{2+}(aq) + 2F^-(aq); \quad K_c = 1.00 \times 10^{-6}$$

$$K_c = [Ba^{2+}][F^-]^2 = 1.00 \times 10^{-6}$$

Species	[Initial]	Δ	[Equilibrium]
Ba^{2+}	0	$+X$	X
F^-	0	$+2x$	$2x$

Define "x" as the change in barium ion concentration

$$[Ba^{2+}][F^-]^2 = \boxed{(x)(2x)^2 = 1.00 \times 10^{-6}}$$

$$4x^3 = 1.00 \times 10^{-6}$$

$$x^3 = 2.50 \times 10^{-7}$$

$$x = 0.0062996052 = [Ba^{2+}]$$

$$\boxed{[Ba^{2+}] = 0.00630 \ M}$$

A 6.00 L reaction vessel contains 0.488 mol hydrogen gas, 0.206 mol iodine vapor, and 2.250 mol HI at equilibrium at 491 C. . What is the value of Kc at 491 C?

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

We're being asked to find Kc ... and we're given equilibrium conditions. So all we need to do is plug MOLAR CONCENTRATION into the equilibrium expression.
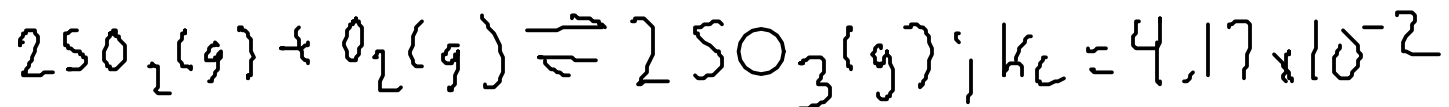
$$[HI] = \frac{2.250 \text{ mol } HI}{6.00 L} = 0.375 \text{ M } HI$$

$$[H_2] = \frac{0.488 \text{ mol } H_2}{6.00 L} = 0.0813333333 \text{ M } H_2$$

$$[I_2] = \frac{0.206 \text{ mol } I_2}{6.00 L} = 0.0343333333 \text{ M } I_2$$

$$K_c = \frac{(0.375)^2}{(0.0813333333)(0.0343333333)} = \boxed{50.4 = K_c}$$

What is the direction of reaction when a mixture of 0.20 M sulfur dioxide, 0.10 M oxygen gas, and 0.40 M sulfur trioxide approaches equilibrium?

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g); \quad k_c = 4.17 \times 10^{-2}$$

$$Q = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(0.40)^2}{(0.20)^2(0.10)} = 40$$

Calculate the reaction quotient, Q!

Compare Q with Kc to find the direction of reaction..

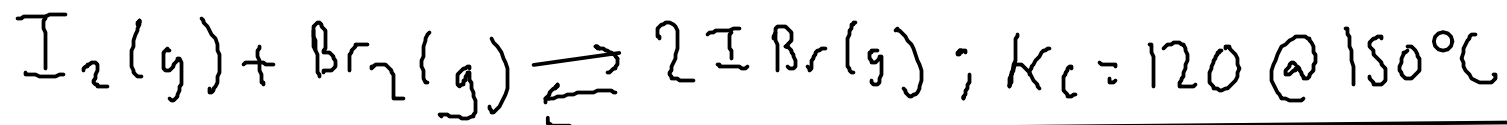$$40 \quad > \quad 4.17 \times 10^{-2}$$

$$Q \quad > \quad k_c$$

Q is much larger than Kc ...

Since Q > Kc, the reaction will proceed to the LEFT (or the "reverse" direction) and produce more sulfur dioxide and oxygen at the expense of sulfur trioxide.

A 5.0 L vessel initially contains 0.0015 mol of each reactant. Find the equilibrium concentrations of all species in the vessel at equilibrium at 150 C.

$$I_2(g) + Br_2(g) \rightleftharpoons 2IBr(g) \; ; \; K_c = 120 @ 150°C$$

$$K_c = \frac{[IBr]^2}{[I_2][Br_2]} = 120$$

Define "x" as the change in iodine concentration

Species	[Initial]	Δ	[Equilibrium]
I_2	$\frac{0.0015 \, mol}{5.0 L} = 0.00030$	$-x$	$0.00030 - x$
Br_2	$\frac{0.0015 \, mol}{5.0 L} = 0.00030$	$-x$	$0.00030 - x$
IBr	0	$+2x$	$2x$

$$\frac{[IBr]^2}{[I_2][Br_2]} = \frac{(2x)^2}{(0.00030 - x)(0.00030 - x)} = 120$$

$$\frac{(2x)^2}{(0.00030 - x)^2} = 120$$

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$$\frac{(2x)^2}{(0.00030-x)^2} = 120$$

We can solve either by using the quadratic formula or by taking the square root of both sides.

$$\sqrt{\frac{(2x)^2}{(0.00030-x)^2}} = \sqrt{120}$$

$$\frac{2x}{0.00030-x} = 10.9544511S$$

$$2x = 10.9544511S\,(0.00030-x)$$

$$0.1825741858x = 0.00030-x$$

$$1.1825741858\,x = 0.00030$$

$$x = 2.54 \times 10^{-4}$$

Species	[Equilibrium]
I_2	$0.00030-x$
Br_2	$0.00030-x$
IBr	$2x$

Plug in value of "x" to find concentrations...

$$[I_2] = 0.00030-x =$$
$$[Br_2] = 0.00030-x =$$
$$[IBr] = \quad 2x \qquad =$$

$$\boxed{\begin{array}{l} 4.6 \times 10^{-5}\,M\ I_2 \\ 4.6 \times 10^{-5}\,M\ Br_2 \\ 5.1 \times 10^{-4}\,M\ IBr \end{array}}$$

When carbon dioxide is removed from the equilibrium mixture by passing the gases through water (which preferentially absorbs carbon DIOXIDE), what is the direction of net reaction as a new equilibrium is achieved?

$$FeO (s) + CO (g) \rightleftharpoons Fe (s) + CO_2 (g)$$
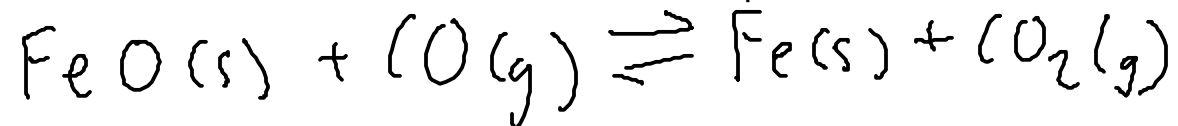
Removing carbon dioxide from the reaction mixture will cause the equilbrium to shift to the RIGHT to compensate for the loss ... the reaction attempts to replace the missing carbon dioxide. More iron(II) oxide and carbon monoxide will be consumed, and more iron will be produced.

Le Chateleir's Principle!

Predict the optimal conditions (temperature and pressure) for maximum conversion of ethylene to ethane.

$$C_2H_4 (g) + H_2 (g) \rightleftharpoons C_2H_6 (g) \; ; \; \Delta H° < 0$$

ethylene   ethane
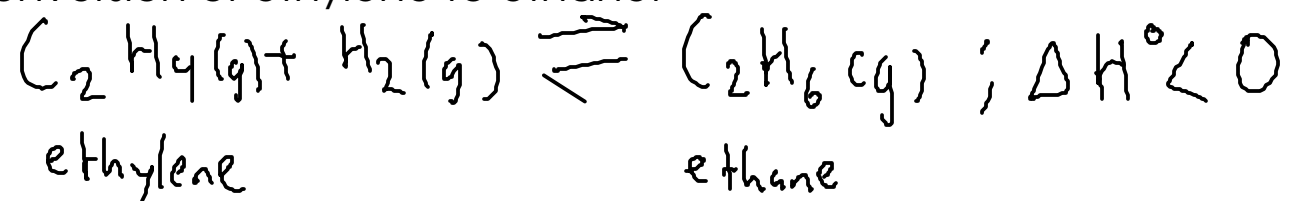
Since the enthalpy change for the reaction is NEGATIVE, this is an EXOTHERMIC process. View heat as a PRODUCT ...

$$C_2H_4 + H_2 \rightleftharpoons C_2H_6 + heat$$

Raising the temperature pushes the reaction to the left, which is what we DON'T want! Instead, RUN THE REACTION AT A LOWER TEMPERATURE (reduce the amount of heat) to cause it to shift right and produce more ethane.

Pressure? Compress or decompress the mixture? For pressure to affect a gas-phase equiilbrium, there must be a way for the equilibrium to respond by changing the number of moles of gas. There are two moles of gas on the left and only one on the right. INCREASING THE PRESSURE would cause the equilibrium to shift to the side with less moles of gas ... the right in this case. This is what we want ... more ethane. So, RUN THE REACTION AT HIGH PRESSURE.