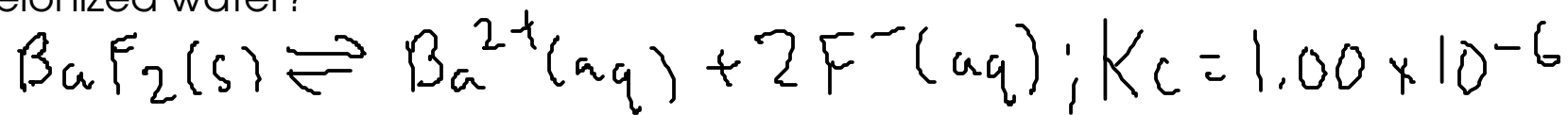What is the concentration of barium ion at equilibrium if solid barium fluoride is mixed with deionized water?

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

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

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

Let "x" equal the change in barium ion cocnentration ...

Plug (Equilibrium) back into the Kc expression ...

$$(x)(2x)^2 = 1.00 \times 10^{-6}$$

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

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

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

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?

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

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

FInd the concentrations of each substance from the information given, then just plug them into the Kc expression ...

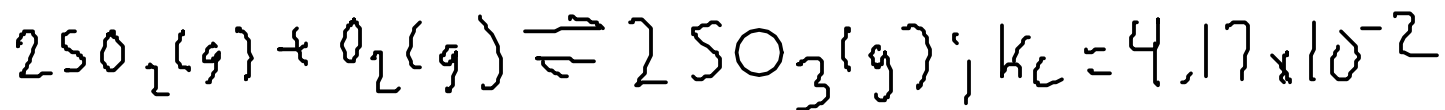$$[H_2] = \frac{0.488 \text{ mol } H_2}{6.00 \text{ L}} = 0.0813333333 \text{ m } H_2$$

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

$$[HI] = \frac{2.250 \text{ mol } HI}{6.00 \text{ L}} = 0.375 \text{ m } HI$$

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

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?

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

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = 4.17 \times 10^{-2}$$

Find Q, the reaction quotient, to determine if this mixture is at equilibrium. Same expression as Kc!

$$Q = \frac{(0.40)^2}{(0.20)^2(0.10)} = 40$$

$Q > K_c$, so the reaction proceeds to the left.

— More SO_2 and O_2 will be made, and SO_3 will be consumed.

A 5.0 L vessel initially contains 0.0015 mol of each reactant. Find the equilibrium concentration of IBr in the vessel at equilibrium at 150 C.

$$I_2(g) + Br_2(g) \rightleftharpoons 2IBr(g) ; K_c = 120 @ 150°C$$
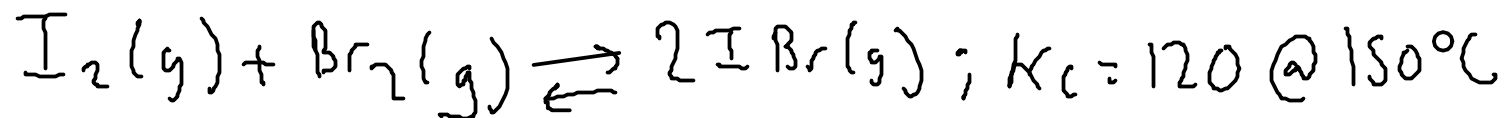
$$K_c = \frac{[IBr]^2}{[I_2][Br_2]} = 120$$

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

Let "x" equal the loss of iodine concentration

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

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

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

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

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

$$0.1825741858 x = 0.00030 - x$$

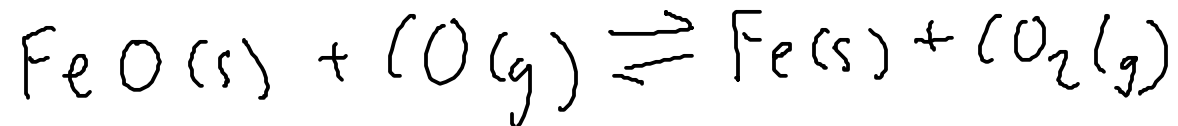$$1.1825741858 x = 0.00030$$

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

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

Use chart to find (IBr) ...

$$[IBr] = 2x = \boxed{5.1 \times 10^{-4} M \; IBr}$$

$$(0.00051 \, m)$$

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?

$$FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g)$$

Passing the gases through water is said to remove carbon dioxide from the system (making its concentration go down). By LE CHATELEIR'S PRINCIPLE, lowering the concentration should cause the equilibrium to shift TO THE RIGHT in an attempt to regenerate some of the lost carbon dioxide.

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
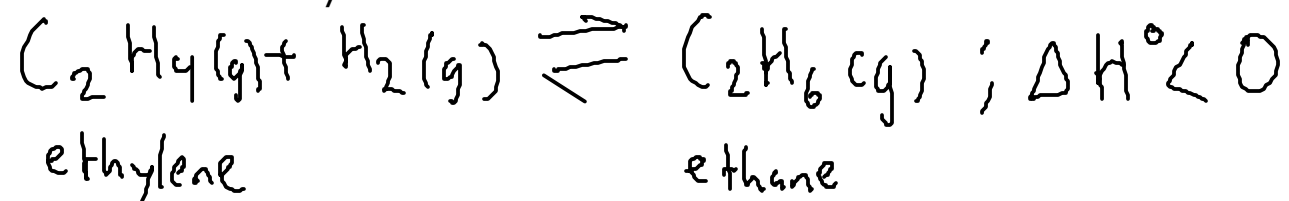
Temperature? The reaction is EXOTHERMIC (negative enthapy change), so ..

$$C_2H_4 + H_2 \rightleftharpoons C_2H_6 + HEAT$$

DECREASING temperature sholuld cause an equilibrium shift to the right, so we want to run at LOW TEMPERATURE.

Pressure? An equilibrium can react to a pressure change if one side of the reaction has fewer moles of gas than the other. Increased pressure should cause THIS equilibrium to shift towards the ethane side (1 mole of gas compared to two).

So ... run the reaction at HIGH PRESSURE.