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

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

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

Let's make a chart to reduce the number of variables ...

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

Let "x" equal the change in barium ion concentration

Plug the expressions back into the Kc equation.

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

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

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

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?

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

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

We already have an equilibrium mixture, so all we need to do is plug the mixture concentrations into this Kc equation.

$$[HI] = \frac{2.250\,mol}{6.00\,L} = 0.375\,M\ HI$$

$$[H_2] = \frac{0.488\,mol}{6.00\,L} = 0.081333333\,M\ H_2$$

$$[I_2] = \frac{0.206\,mol}{6.00\,L} = 0.034333333\,M\ I_2$$

$$K_c = \frac{(0.375)^2}{(0.081333333)(0.034333333)} = \boxed{50.4 = K_c}$$

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?

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g);\ K_c = 4.17\times10^{-2}$$

Write an equation for Q, the reaction quotient.

$$Q = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = ?$$

Plug in the current concentrations, then compare Q to Kc...

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

$$Q > K_c$$

$$40 > 4.17\times10^{-2}$$

Since Q > Kc, the reaction proceeds to the LEFT, forming more sulfur dioxide and oxygen gas.

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.

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

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

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

Let "x" equal the change in iodine concentration

Plug expressions back into Kc equation...

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

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

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

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

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

$$\frac{2x}{0.00030-x} = 10.9544515$$

$$2x = 10.9544515(0.00030-x)$$

$$2x = 0.0032863353 - 10.9544515x$$

$$12.9544515x = 0.0032863353$$

$$x = 0.000254$$

$[I_2] = 0.00030 - x = 0.0000463\ M\ I_2$

$[B_2] = 0.00030 - x = 0.0000463\ M\ Br_2$

$[IBr] = 2x = 0.000507\ M\ IBr$

Looks reasonable given Kc = 120 (more products at equilibrium)

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 gas mixture through water acts to REMOVE carbon dioxide, lowering its concentration. Le Chateleir's Principle suggests that the equilibrium will shift to counteract that - meaning that the equilibrium will try to make more carbon dioxide to replace the carbon dioxide that was lost.

The reaction should proceed TO THE RIGHT, making more carbon dioxide and iron.

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

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

ethylene — ethane

FIrst, delta H being less than zero (negative), means the reaction is EXOTHERMIC.

If the reaction is EXOTHERMIC, we can treat the heat as a PRODUCT:

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

Increased temperature should shift the reaction to the left (to relieve the increase).

So, we should run the reaction at LOW TEMPERATURE, since that woudl cause the equilibrium to shift towards the ethane side.

Pressure? COMPRESSING the gas should result in the equilibrium shifiting towards the side with less gas. Here, that's the ethane (product) side.

$$C_2H_4 + H_2 \rightleftharpoons C_2H_6$$

(2 mol gas) — (1 mol gas)

The optimal conditions should be LOW TEMPERATURE and HIGH PRESSURE.