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

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

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

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

let 'x' be the increase in barium ion concentration

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

Solve for 'x'

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

$$x = 6.30 \times 10^{-3}$$

Since $x = [Ba^{2+}]$, $[Ba^{2+}] = 6.30 \times 10^{-3}\ M$

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 are given the equilibrium state and need to find Kc. To do that, we'll need to calculate the equilibrium concentrations from the information provided.

Calculate concentrations:

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

$$[H_2] = \frac{.488\text{ mol}}{6.00\text{ L}} = 0.081333333\,M\ H_2$$

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

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

$$K_c = 50.4$$

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?

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

Use reaction quotient Q to figure out what diresction the reaction will proceed:

$$Q = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(0.40)^2}{(.20)^2(.10)} = 40$$

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

$$Q > K_c$$

So, reaction proceeds to LEFT

(makes more $SO_2 + O_2$)

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.

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

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

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

We defined 'x' as the decrease in concentration of iodine vapor.

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

Solve for 'x'

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

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

You may use the quadratic equation here OR take the square root of both sides.

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

$$\frac{2x}{0.00030-x} = 10.95445115$$

$$2x = 10.95445115(0.00030-x)$$

$$0.182574185x = 0.00030 - x$$

$$1.182574186x = 0.00030$$

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

$[I_2] = 0.00030 - x = 0.000046\ M\ I_2$

$[Br_2] = 0.00030 - x = 0.000046\ M\ Br_2$

$[IBr] = 2x = 0.00051\ M\ IBr$

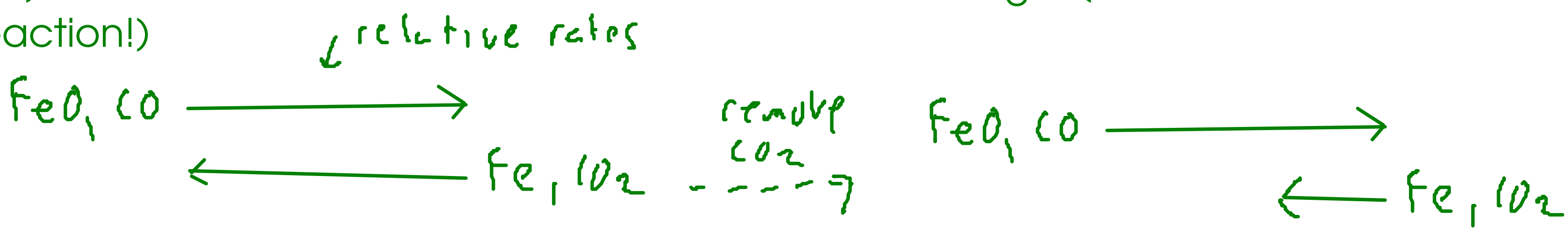
Given that Kc=120, we expect to have more IBr at equilibrium ... and we do!

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?

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

If we remove carbon dioxide from the equilibrium, we expect that the equilibrium will shift to compensate. For the equilibrium to make more carbon dioxide, it must shift to the RIGHT.

An alternate way to think about this is ... removing carbon dioxide lowers the rate of the REVERSE reaction while leaving the forward rate unchanged (at least at first). This means the net direction of reaction is to the right (same as the forward reaction!)



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

Temperature? This is an EXOTHERMIC PROCESS. Rewritten, it looks like this:

$$C_2H_4 + H_2 \rightleftharpoons C_2H_6 + \text{heat}$$

* An increase in temperature (more heat) should cause this equilibrium to shift left (back to reactants). This is the opposite of what we want!

→ * At low temperatures, we have the opposite ... more products should be produced.

Pressure? This reaction is gas-phase and has a different number of moles of gas on each side. It SHOULD respond to pressure changes.

→ * Compressing the mixture (increasing pressure by decreasing volume) should cause the equilibrium to shift towards the side with less moles of gas ... since that will decrease the overall pressure. That's the ethylene side, which is what we want!

To sum up, the optmum conditions are LOW TEMPERATURE and HIGH PRESSURE.