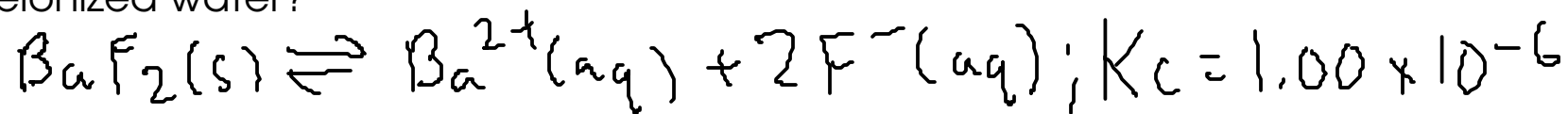


What is the concentration of barium ion at equilibrium if solid barium fluoride is mixed with deionized water?



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

Define 'x' as the change in barium ion concentration...

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

Substitute equilibrium conditions into the equilibrium expression...

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

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

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

$$x = 0.00630 = [\text{Ba}^{2+}]$$

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

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 K_c at 491 C?



$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = ?$$

Calculate molar concentrations of hydrogen, iodine, and HI ...

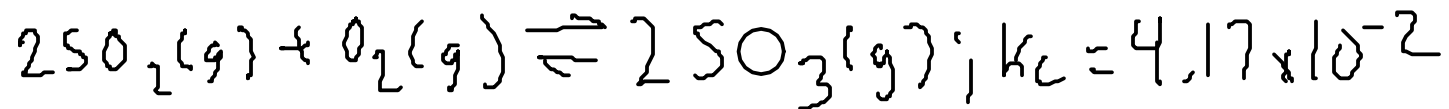
$$[\text{HI}] = \frac{2.250 \text{ mol HI}}{6.00 \text{ L}} = 0.375 \text{ M HI}$$

$$[\text{H}_2] = \frac{0.488 \text{ mol H}_2}{6.00 \text{ L}} = 0.0813333333 \text{ M H}_2$$

$$[\text{I}_2] = \frac{0.206 \text{ mol I}_2}{6.00 \text{ L}} = 0.0343333333 \text{ M I}_2$$

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

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?

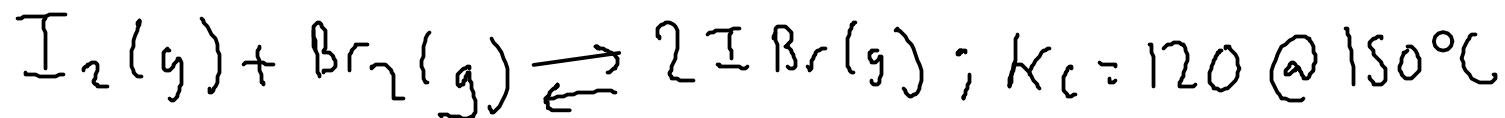


$$Q = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]} = \frac{(0.40)^2}{(0.20)^2 (0.10)} = 40$$

$$Q > K_c$$

Since $Q > K_c$, the reaction proceeds to the LEFT SIDE to form 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.



$$K_c = \frac{[\text{IBr}]^2}{[\text{I}_2][\text{Br}_2]} = 120$$

Define "x" as the change in iodine concentration

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	$+2x$	$2x$

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

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

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

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

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

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

$$\frac{2}{\sqrt{120}}x + x = 0.00030$$

$$x\left(\frac{2}{\sqrt{120}} + 1\right) = 0.00030$$

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

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

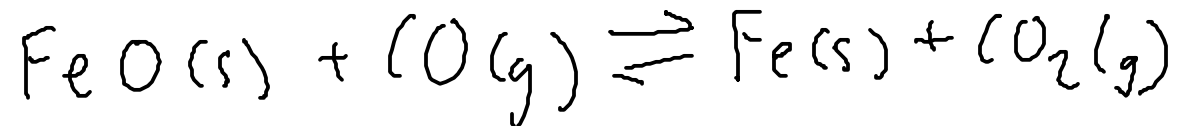
Now plug our value of "x" into the expressions for equilibrium concentrations...

$$[I_2] = 0.00030 - 2.54 \times 10^{-4} = 4.6 \times 10^{-5} M$$

$$[Br_2] = 0.00030 - 2.54 \times 10^{-4} = 4.6 \times 10^{-5} M$$

$$[IBr] = 2(2.54 \times 10^{-4}) = 5.1 \times 10^{-4} M$$

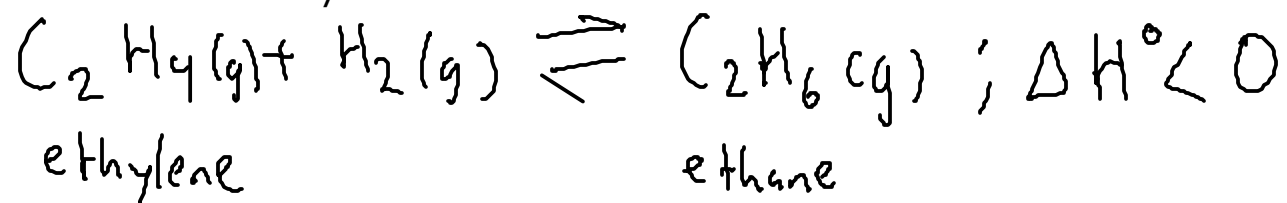
When carbon dioxide is removed from the equilibrium mixture by passing the gases through water (which preferentially absorbs carbon DIOXIDE), what is the direction of net reaction as a new equilibrium is achieved?



Removing carbon dioxide from the equilibrium mixture lowers the concentration of carbon dioxide (a product). This causes the equilibrium to try to replace the lost carbon dioxide and shift to the RIGHT.

(Le Chateleur's Principle)

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



Since the enthalpy change of reaction is negative (less than zero), this is an EXOTHERMIC process - the reaction releases energy as it proceeds, and we can imagine heat as a PRODUCT of the reaction:



Raising the temperature ("turning up the heat") will cause the equilibrium to shift LEFT to reduce the amount of heat. This is the OPPOSITE of what we want! Instead, we should run the reaction at a LOWER temperature ... which will cause the equilibrium to shift RIGHT

Pressure? For pressure to effect a gas-phase equilibrium, we must have more moles of gas on one side than on the other. We do .. there are more moles of gas on the left side (ethylene and hydrogen).

If we compress the reaction mixture (increasing pressure), we would expect the equilibrium to shift towards the side with less gas ... in this case, the ethane side. This is exactly what we want!

Optimum conditions: LOW temperature, HIGH pressure.