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

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

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

Make an equilibrium chart to reduce the number of variables.

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

Assigned "x" to be the change in barium ion concentration

Plug back into the equilibrium expression...

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

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

$$x = 0.00629960052 = [Ba^{2+}]$$

So, $[Ba^{2+}] =$ **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?

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

We're being asked to find Kc for a mixture already at equilibrium. To do that, we'll just plug equilibrium concentrations into the Kc expression.

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

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

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

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

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(0.375)^2}{(0.081333333)(0.034333333)} = \boxed{50.4}$$

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 the reaction quotient Q to find out which direction this reaction goes.

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

$$Q > K_c$$

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

Since Q>Kc, the reaction is not at equilibrium and will proceed to the LEFT to form more sulfur dioxide and oxygen.

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

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

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

Set up an equilibrium chart to reduce the number of variables.

Species	[Initial]	Δ	[Equilibrium]
I_2	$\frac{0.0015\,mol}{5.00\,L} = 0.0003$	$-x$	$0.0003 - x$
Br_2	$\frac{0.0015\,mol}{5.00\,L} = 0.0003$	$-x$	$0.0003 - x$
IBr	0	$+2x$	$2x$

Assigned "x" to be the change in iodine concentration

$$\frac{(2x)^2}{(0.0003 - x)(0.0003 - x)} = 120$$

$$\frac{(2x)^2}{(0.0003 - x)^2} = 120$$

$$\frac{(2x)^2}{(0.0003-x)^2} = 120$$

We can solve this with the quadratic equation OR we can take the square root of both sides!

Species	[Equilibrium]
I_2	0.0003 - x
Br_2	0.0003 - x
IBr	2x

$$\sqrt{\frac{(2x)^2}{(0.0003-x)^2}} = \sqrt{120}$$

$$\frac{2x}{0.0003-x} = 10.9544511 5$$

$$2x = 10.9544511 5(0.0003-x)$$

$$0.182574185 8x = 0.0003 - x$$

$$1.182574185 8x = 0.0003$$

$$x = 0.00025368387 3$$

Plug the value of "x" we calculated back into the expression for IBr concentration.

$$[IBr] = 2x = 0.00051 \text{ M IBr}$$

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?

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

Since the water absorbs carbon dioxide, it removes it from the gas and LOWERS ITS CONCENTRATION.

Le Chateleir's principle says that an equilibrium will shift to counteract a disturbance. If we lower a concentration, then the equilbrium will try to make more of the substance we lowered the concentration of!

The equilibrium will shift to the RIGHT to make more iron and carbon dioxide.

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

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

ethylene — ethane

EXOTHERMIC!

Since the enthalpy change is less than zero, the reaction is EXOTHERMIC.

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

(ethane)

Temperature: Will a high or low temperature make more ethane?

If we run at LOWER TEMPERATURES, the equilibrium should shift to the right for more ETHANE.

Pressure: Is this equilibrium affected by pressure?

$$C_2H_4(g) + H_2(g) \rightleftharpoons C_2H_6(g)$$

(2 mol gas) | (1 mol gas)

Different number of moles of gas on each side, so pressure affects this eq!

Run reaction at HIGH PRESSURE, then the equilibrium will shift to the side of the reaction with fewer moles of gas ... the ETHANE side.

Optimal conditions: LOW TEMPERATURE and HIGH PRESSURE.