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

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.

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

We need to express each of these concentrations in terms of a single variable.

Species	[Initial]	Δ	[Equilibrium]
I_2	$\frac{0.0015 \text{ mol}}{5.00 L} = 3\times10^{-4}$	$-x$	$0.0003 - x$
Br_2	$\frac{0.0015 \text{ mol}}{5.00 L} = 3\times10^{-4}$	$-x$	$0.0003 - x$
IBr	0	$+2x$	$2x$

$$\frac{[IBr]^2}{[I_2][Br_2]} = \boxed{120 = \frac{(2x)^2}{(0.0003-x)(0.0003-x)}}$$

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

Let's try taking the square root of both sides!

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

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

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

$$0.0032863353 - 10.9544511x = 2x$$

$$0.0032863353 = 12.9544511x$$

$$2.536838734 \times 10^{-4} = x$$

$$[I_2] = 0.0003 - x = 4.6 \times 10^{-5}\ M \quad (0.46 \times 10^{-4})$$

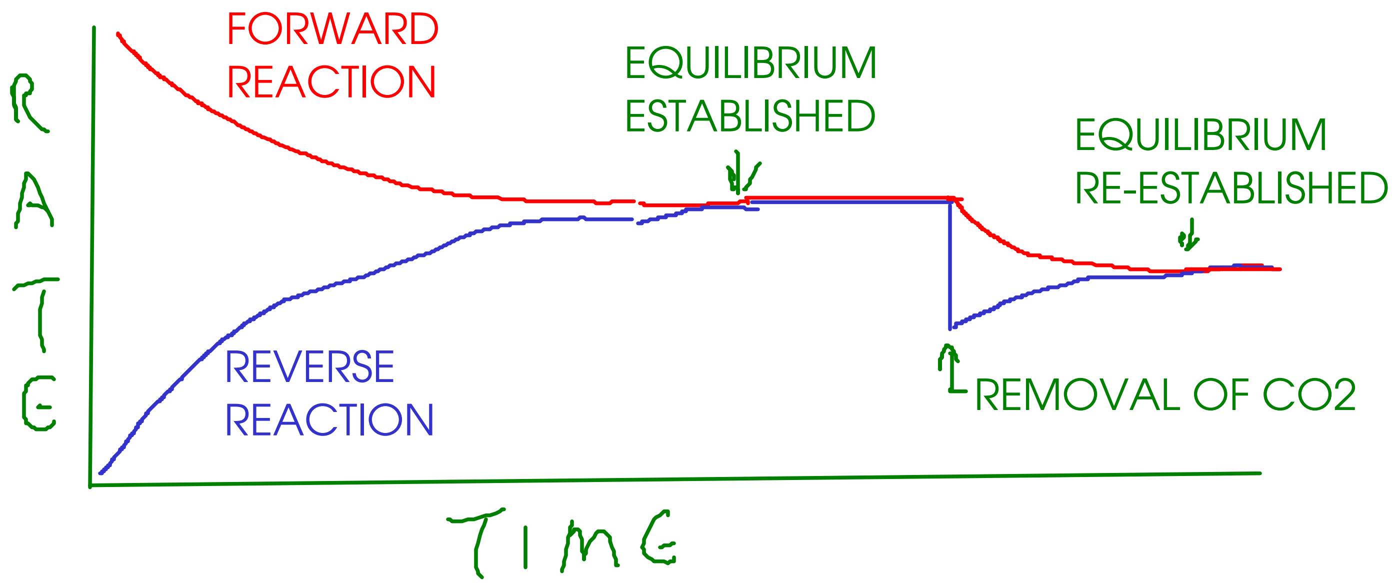
$$[Br_2] = 0.0003 - x = 4.6 \times 10^{-5}\ M$$

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

Since Kc=120, we aren't surprised by
a product-dominated equilibrium state!

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

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?



Lowering the concentration of product lowers the rate of the REVERSE reaction. This means that the new equilibrium state will have more IRON than the old one, and the equilibrium will shift in the direction of PRODUCTS as it re-establishes itself.

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g);\ \Delta H^\circ = -21.7\ \text{kcal}$$
(methanol)

Would the fraction of methalnol at equilibrium be increased by raising the temperature? Why (or why not)?

Exothermic reaction (negative enthalpy change): Increased temperature increases the amount of heat available, forcing the equilibrium to shift towards reactants - LOWERING the fraction of methanol at equilibrium.

$$CO(\ ) + 2H_2(g) \rightleftharpoons CH_3OH(g) + \text{heat}$$

View heat as a PRODUCT in an exothermic reaction!

What about PRESSURE? Would COMPRESSING the mixture (increasing pressure by decreasing volume) increase the methanol fraction?

There are three moles of gas on the left, but only one mole of gas on the right. The equiilibrium will shift towards the right to relieve the pressure, and this will increase the methanol fraction!

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

ethylene ethane

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

Temperature: This is an EXOTHERMIC process. We can view heat as a PRODUCT

* Higher termperatures would cause the equilibrium to shift LEFT, forming more ETHYLENE. This is not what we want.
* Lower temperatures would cause the equilibrium to shift RIGHT, forming more ETHANE. This IS what we want!

Pressure: There are more moles of gas on the left side of this reaction, so the equilibrium does respond to compression.

* At increased pressure, the equilibriun will shift towards the right, making more ETHANE and reducing pressure. This is what we want!
* At reduced pressure, the equilibrium will shift towards the left, making more ETHYLENE and increasing pressure. This is not what we want.

Optimally, we'll use low temperature and high pressure to run the reaction.