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

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. (find # moles)

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

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	$+2x$	$2x$

Let's define 'x' to be the change in concentration of iodine...

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

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

Looks to be solvable by taking the square root of both sides...

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

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

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

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

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

$$x = 2.54 \times 10^{-4} = 0.000254$$

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

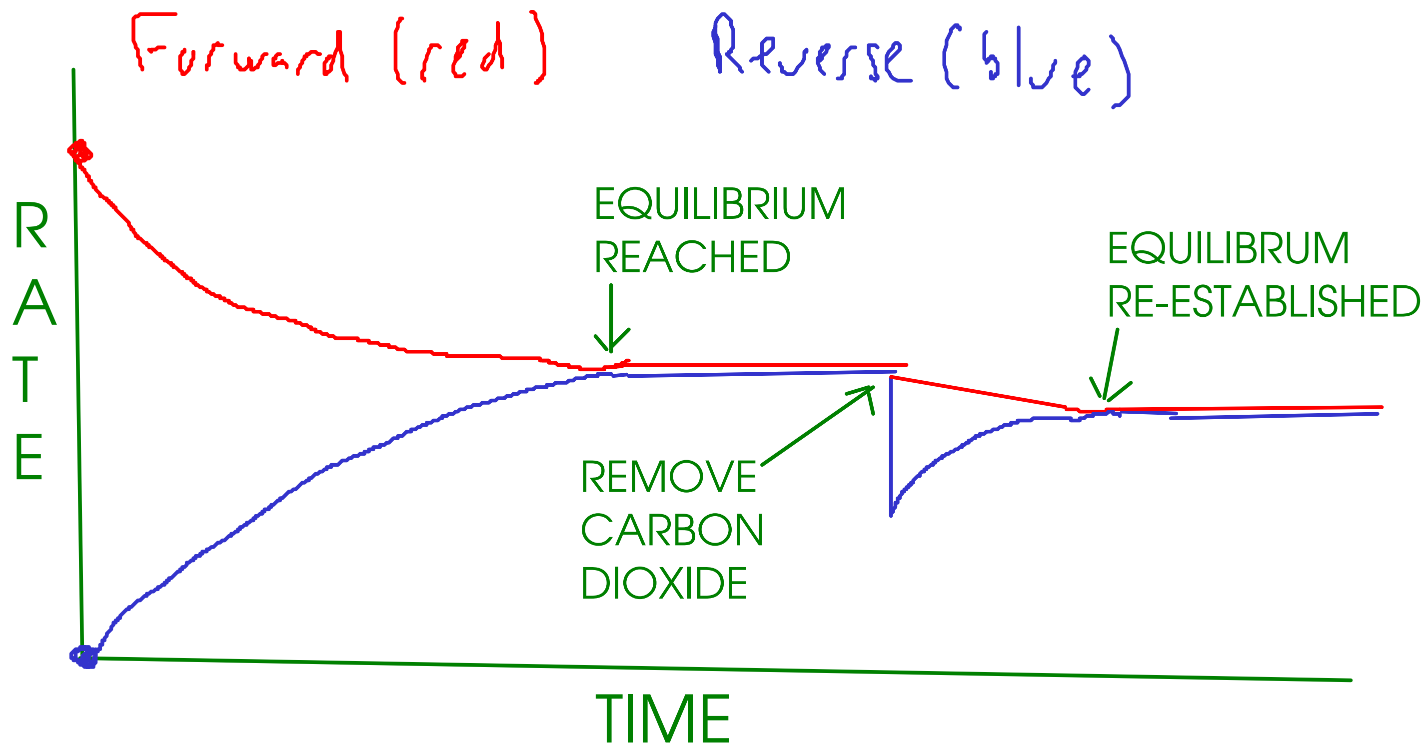
$[I_2]$: $0.00030 - 0.00025 = 0.00005\,M \times 5.0\,L = 0.00025\ mol\ I_2$

$[Br_2]$: $0.00030 - 0.00025 = 0.00005\,M \times 5.0\,L = 0.00025\ mol\ Br_2$

$[IBr]$ $2x$ $= 0.00050\,\underline{M} \times 5.0\,L = 0.0025\ mol\ IBr$

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

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?



Removing carbon dioxide affects reverse reaction.

$$Rate = k_r [CO_2]^q \downarrow$$

Since the reverse reaction is slowed, the forward rate will be greater than the reverse rate until equilibrium is re-established. This meanse the equilibrium will shift to the right and make more PRODUCTS!

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

(methanol)

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

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

This is an exothermic reaction, so we can view HEAT as a product. Increasing temperature effectively increases the amount of heat, and the equiilbrium should shift to counter that disturbance. This means equilibrium shifts to the left, lowering the methanol fraction. (If we want a higher fraction of methanol, we should actually LOWER the temperature.)

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

Since this equilibrium has three moles of gas on the reactant side and only one mole of gas on the product side, we would expect this equilibtrium to respond to compression.

The equilibrium will shift towards products (less moles of gas), so compressing the mixture will increase the methanol fraction.

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

ethylene ethane

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

Temperature: This is an exothermic process. We view heat as if it's a PRODUCT of the reaction:

* At high temeratures, we have more heat available, and this will drive the equilibrium to the left (ethylene side).
* At low temeratures, we have less heat available, and this will drive the equilibrium to the right (ethane side).

Pressure: This reaction has two moles of gas on the reactant side, and one mole on the product side.

* Compressing this equilibrium will cause it to shift to the side with fewer moles of gas (the ethane side).

So the optimum conditions would be:
Low temperature
High pressure!