14,67, p 618: A 5.0 L vessel initially contains 0.0015 mol of each reactant. Find equilibrium composition of the mixture.

$$I_2(g) + Br_2(g) \rightleftarrows 2\,IBr(g)\ ;\ K_c = 120 \text{ @ } 150\,^\circ C$$

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

We need to express all of these in terms of one variable

Species	[Init]	Δ	[Equil]
I_2	$\frac{0.0015\text{ mol}}{5.0\text{ L}} = 3\times10^{-4}$	$-x$	$3\times10^{-4} - x$
Br_2	$\frac{0.0015\text{ mol}}{5.0\text{ L}} = 3\times10^{-4}$	$-x$	$3\times10^{-4} - x$
IBr	0	$+2x$	$2x$

$$K_c = \frac{[IBr]^2}{[I_2][Br_2]} = 120 = \frac{(2x)^2}{(3\times10^{-4} - x)(3\times10^{-4} - x)}$$

$$120 = \frac{(2x)^2}{(3\times10^{-4}-x)(3\times10^{-4}-x)}$$

$$120 = \frac{(2x)^2}{(3\times10^{-4}-x)^2}$$

$$\sqrt{120} = \frac{2x}{3\times10^{-4}-x}$$

$$\sqrt{120}\,(3\times10^{-4}-x) = 2x$$

$$0.003286335 - \sqrt{120}\,x = 2x$$

$$0.003286335 = (2+\sqrt{120})\,x$$

$$2.5\times10^{-4} = x$$

Species	[Equil]
I_2	$3\times10^{-4}-x$
Br_2	$3\times10^{-4}-x$
IBr	$2x$

$[I_2] = 0.5\times10^{-4}\ (5\times10^{-5})$ M

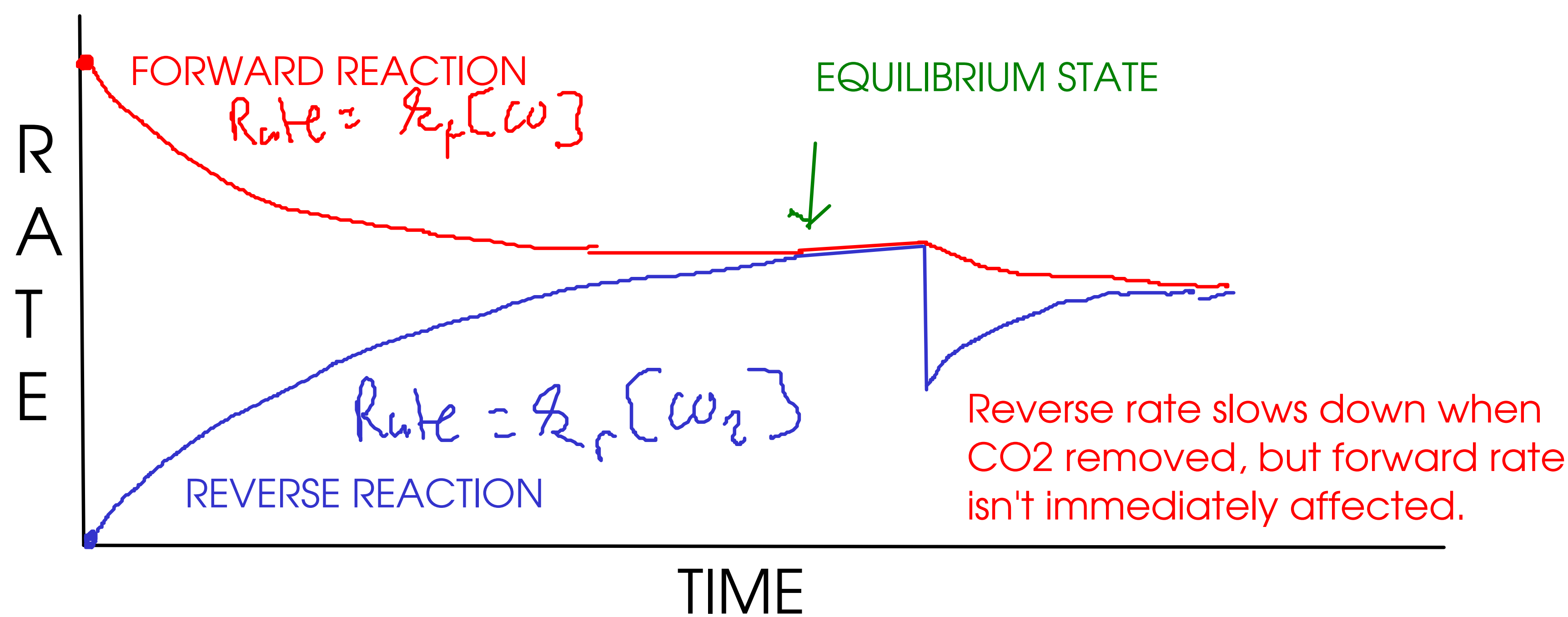
$[Br_2] = 0.5\times10^{-4}\ (5\times10^{-5})$ M

$[IBr] = 5.0\times10^{-4}$ M

Quick check ... Kc = 120, so we expect products to dominate at equilibrium ... and they do!

14.73, p 618: If carbon dioxide is removed from the equilibrium mixture by passing the gas through water, what is the direction of net reaction as the new equilibrium is achieved?

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



So, the net reaction proceeds to the RIGHT, making more iron and carbon dioxide at the expense of iron(II) oxide and carbon monoxide.

14,77, p 618: Would the fraction of methanol at equilibrium be increased by raising temperature? (Additional question: What about pressure?)

$$CO(g) + 2H_2(g) \rightleftarrows CH_3OH(g)\ ;\ \Delta H^\circ = -21.7\ kcal$$

methanol

What about temperature response?

This is an exothermic reaction (heat is evolved):

$$CO(g) + 2H_2(g) \rightleftarrows CH_3OH(g) + heat$$

If you view heat as a product (since it's being released), then Le Chateleir's principle applies. This means that increased temperature shifts the equilibrium to the left, and the FRACTION OF METHANOL GOES DOWN.

What about increased PRESSURE?

There are three moles of gas on the left, and one on the right. If we increase pressure (by compression), the equilibrium will shift to the side with fewer moles of gas - to lower pressure. Since that's the methanol side of the equilibrium, the FREACTION OF METHANOL WILL INCREASE.

14.82, p 618: Predict the optimal conditions for conversion of ethylene to ethane.

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

ethylene ethane

Temperature:

$$C_2H_4(g) + H_2(g) \rightleftarrows C_2H_6(g) + \text{heat}$$

At warm temperatures, the equilibrium would shift to the left, which is the opposite of what we want - less ethane and more ethylene.

Cooler temperatures (less heat) would shift the equilibrium to the right and make more ethane. This is the desired outcome!

Pressure:

There are two moles of gas on the left side and only one mole on the right, so this equilibrium WILL respond to pressure changes.

At increased pressure, the equilibrium will shift towards the side of the reaction with fewer moles of gas - the ethane side. This is what we want.

Optimum conditions are low temperature and high pressure.