14.82, page 618

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

ethylene ethane

Predict the optimal conditions to maximize conversion of ethylene to ethane. (T & P)

Temperature: This is an exothermic process. We can treat heat as if it;s a PRODUCT.

At higher temperatures, we have MORE HEAT AVAILABLE. This will cause the equilibrium to shift away from the side with heat (the ethane side) and towards the ethylene side.

* Cooler temperatures cause the equilibrim to shift towards the side with heat (the ethane side). This is what we want!

Pressure:

There are more moles of gas on the ethylene side of the reaction (reactants) than the ethane side (products). So, pressure affects this equilibrium and the answer is NOT that it doesn't matter what pressure we use!

At increased pressure, the equilibrium will shift to the side with fewer moles of gas (the ethane side), which will relieve the pressure increase. This is what we want!

So, the OPTIMUM CONDITIONS are: Low temperature, high pressure.

14.77, page 618

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

methanol

Would the fraction of methanol obtained at equilibrium be increased by raising temperature? Why or why not?

Exothermic reaction: We should view heat as a PRODUCT. So, increased temperature would cause equilibrium to shift away from the side with the heat

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g) + HEAT$$

$$\longleftarrow$$

This REDUCES the fraction of methanol in the mixture, because the equilibrium shifts away from the methanol side.

Would increased PRESSURE increase the fraction of methanol at equilibrium?

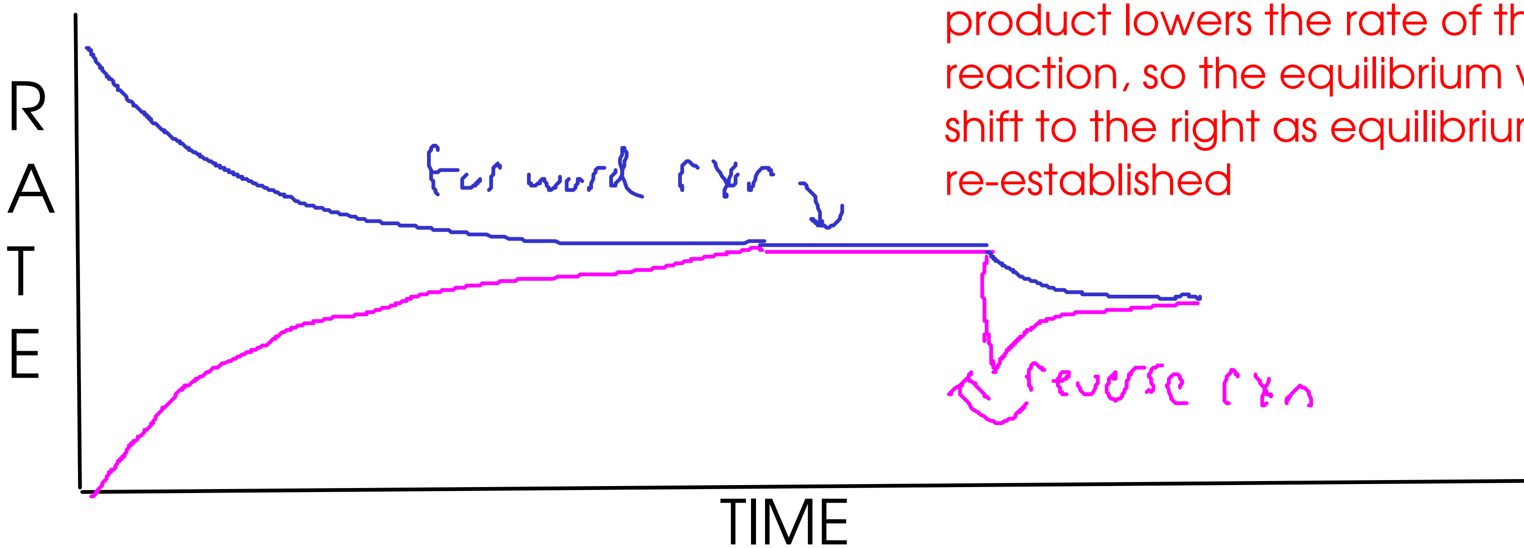
Since there are three moles of gas on the reactant side, and only one mole on the product (methanol) side, an increase in pressure would cause the equilibrium to shift to the right (relieving the pressure) and making more methanol - INCREASING the fraction of methanol in the mixture.

14.73, page 618

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

When carbon dioxide is removed from the equilibrium mixture, what is the direction of reaction as the equilibrium is re-established? (The way this is done is to spray the mixture with water - since carbon dioxide dissolves in water much better than carbon monoxide does)

Lowering the concenration of a product lowers the rate of the reverse reaction, so the equilibrium will shift to the right as equilibrium is re-established



14.67, p618

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

What are the equilibrium concentrations if a 5.0L reaction vessel iniitally contains 0.0015 moles of iodine and 0.0015 moles of bromine.

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

We need to solve this expression, but it's got too many variables. We must express these concentrations in terms of a single variable.

Species	[Initial]	Δ	[Equilibrium]
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	$+2x$	$2x$

We've assigned 'x' to be the change in concentration of iodine

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

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

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

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

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

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

$$3\times10^{-4} = \left(\frac{2}{\sqrt{120}}+1\right)x$$

$$\frac{3\times10^{-4}}{\left(\frac{2}{\sqrt{120}}+1\right)} = x$$

$$x = 2.5366\times10^{-4}$$

$$[I_2] = 3\times10^{-4} - x = 4.6\times10^{-5}\ \underline{M}$$

$$[Br_2] = 3\times10^{-4} - x = 4.6\times10^{-5}\ \underline{M}$$

$$[IBr] = 2x = 5.1\times10^{-4}\ \underline{M}$$