14.82, p 618

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

ethylene ethane

Predict the optimal conditions (T and P) for maximum conversion of ethylene to ethane.

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

* At warm temps, we have MORE HEAT AVAILABLE, which will cause the equilibrium to shift away from the side with heat and towards the ethylene side

* Cooler temperatures cause the equilibrium to shift to the right, making more ethane. (This is what we want!)

Pressure:

There are more moles of gas on the reactant side, so this equilibrium WILL respond to pressure change. (The answer is NOT that it doesn't matter what pressure you use.)

At increased pressure, the equilibrium will shift to reduce pressure (to the side with less moles of gas - the product side). This produces more ethane, which is what we want.

OPTIMUM CONDITIONS: Low temperature, high pressure.

14.77, p 628

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

methanol

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

Exothermic reaction. View heat as a product. Increased temperature increases the amount of heat available, forcing the equlibrium to shift towards reactants. This LOWERS the fraction of methanol in the mixture.

Would increased PRESSURE increase the fraction of methanol at equilibrium?

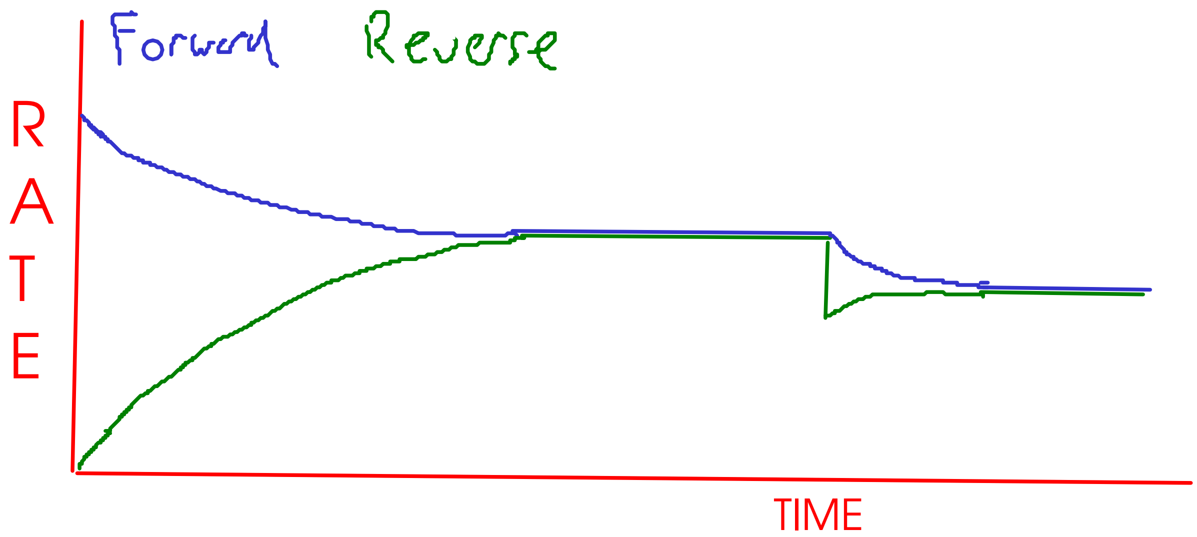
There are three moles of gas on the reactant side and only one mole on the product side, so increasing pressure would cause the equilibrium to shift to the side with less moles of gas (product), INCREASING the fraction of methanol.

14.73

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

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 overall equilibrium will shift to the RIGHT (towards products) as a new equilibrium is achieved.



14.67

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

A 5.0L vessel initially contains 0.0015 moles of each of the reactants. Find equilibrium composition of the mixture in the vessel.

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

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

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

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

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

Easiest to solve by taking square root of both sides.

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

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

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

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

$$x = \frac{3\times10^{-4}}{\left(1+\frac{2}{\sqrt{120}}\right)} = 2.5\times10^{-4}$$

$$[I_2] = 3\times10^{-4} - 2.5\times10^{-4} = \boxed{0.5\times10^{-4}} \rightarrow 0.00025\text{ mol}$$

$$[Br_2] = 3\times10^{-4} - 2.5\times10^{-4} = \boxed{0.5\times10^{-4}} \rightarrow 0.00025\text{ mol}$$

$$[IBr] = \boxed{5.0\times10^{-4}} \times 5.0\text{ L} = 0.0025\text{ mol}$$

(Textbook wanted molar concentrations...)