The equilibrium constant, K_c, for the reaction:

 $CO_2(g) <---> CO(g) + 1/2 O_2(g) \text{ is } 6.7 \times 10^{-12} \text{ at } 1000 \text{ K.}$ **1.** Calculate K_c for the reaction: $2 CO(g) + O_2(g) <---> 2 CO_2(g)$

1. 6.7 x 10 ⁻¹² 2. 1.3 x 10 ⁻¹¹	For the first reaction $K_1 = \frac{[CO(g)][O_2(g)]^{1/2}}{[CO(g)]} = 6.7 \times 10^{-12}$
3. 2.5×10^{-6}	Very reactant (CO ₂) favored.
4. 3.9×10^{-9} 5. 1.5×10^{-11}	
6. 3.0×10^{11}	For the second reaction $K_2 = \frac{[CO_2(g)]^2}{[CO_2(x)]^2 + Q(x)]^2} = (1/K_1)^2 = (1/6.7 \times 10^{-12})^2$
7. 2.2 x 10 ²²	$[CO(g)]^{-}[O_{2}(g)]$
	= 2.2×10^{22} very product (CO ₂) favored.

Note: The reaction doesn't care which way we write it. The chemistry is the same, CO_2 is the substance in abundance.

2. This reaction is: 1. Reactant favored 2. Product favored at equilibrium.

3. Consider the nitrogen dioxide equilibrium: $2 \text{ NO}_2(\mathbf{g}) \iff N_2 O_4(\mathbf{g})$ Write the equilibrium constant for this reaction in terms of the equilibrium constants, Ka and Kb, for reactions **a** and **b** below: $N_2(g) + 2 O_2(g) < ----> N_2O_4(g)$ Ka a.) b.) $1/2 N_2(g) + O_2(g) < > NO_2(g)$ Kb 1. $K = K_A K_B$ $N_2(g) + 2 O_2(g) < ----> N_2O_4(g)$ Rxn 1: Use as is Ka 2. $K = (K_A/K_B)$ $2NO_2(g) < > N_2(g) + 2O_2(g) (1/K_b)^2$ Rxn 2: Reverse, muliply by 2 3. $K = 2 (K_A/K_B)$ $2 \text{ NO}_2(g) \iff N_2O_4(g)$ Add (multiply K's) 4. $K = 2(K_B/K_A)$ $K = K_a (1/K_b)^2 = K_a / (K_b)^2$ 5. $K = (K_A/K_B)^2$ 6. $K = K_A / (K_B)^2$ 7. $K = (K_B/K_A)^{1/2}$ 8. $K = (1/2)(K_A/K_B)^{1/2}$ 9. K = $(1/2)(K_A)/(K_B)^2$

4. The formation of ammonia is an extremely important reaction worldwide for the production of fertilizers and explosives. At 25° C the equilibrium constant, Kc, is 3.5 x 10^{8} .

 $N_2(g) + 3H_2(g) < ----> 2NH_3(g)$

A. If for the NH₃ reaction above at 25°C, the reaction quotient is equal to 3.5, we can say that:

1. More $NH_3(g)$ must form in order to reach equilibrium. Q < K rxn runs --->

- 2. More $N_2(g)$ must form in order to reach equilibrium.
- 3. More $H_2(g)$ must form in order to reach equilibrium.
- 4. The reaction is at equilibrium.
- 5. Cannot tell from the information given.

In another experiment the reaction has run for a while and the concentrations are found to be:

Nitrogen	$7x10^{-2}$ M	Hydrog	gen $9 \times 10^{-3} M$	Ammonia 2x10 ⁻⁴ M
	<i>Q</i> =	$= \frac{[NH_3]^2}{[N_2][H_2]^3} =$	$=\frac{(2x10^{-4})^2}{(7x10^{-2})(9x10^{-3})^2}$	$\frac{1}{3} = 0.78 < K$

B.	In order to reach equilibrium the reaction must:	C. Why ?
	1. Run in the forward direction.	1. $\mathbf{Q} > \mathbf{K}$
	2. Run in the reverse direction	2. $Q = K$
	3. The reaction is at equilibrium	3. Q < K
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How does ΔG relate to ΔG^0 and Q ?

$$\Delta G = \Delta G^0 + RT \ln Q = RT \ln \left(\frac{Q}{K}\right)$$

What does ΔG tell us about a chemical reaction ?

 ΔG tells the direction a reaction must run in order to reach equilibrium.

 $\Delta G < 0$ Reaction spontaneous in forward direction.

 $\Delta G = 0$ Reaction at equilibrium. No further driving force for change.

 $\Delta G > 0$ Reaction NONspontaneous in forward direction Reaction spontaneous in REVERSE direction What does ΔG^0 tell us about a chemical reaction ?

 $\Lambda G^0 = -RT \ln K$

 ΔG^0 tells how far the equilibrium postion lies toward reactants or products.

How does ΔG^0 relate to K?

- $\Delta G^0 < 0$ Equilibrium lies closer to products. PRODUCT-favored $\Delta G^0 > 0$ Equilibrium lies closer to reactants.
- REACTANT-favored
- $\Delta G^0 = 0$ Equilibrium lies about midway between. (This is rare)