Chemical Equilibrium

1. Write the equilibrium law for each of the following chemical reaction equations.

   a) $2\ SO_2(g)\ +\ O_2(g)\ \rightleftharpoons\ 2\ SO_3(g)$

   $$k_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

   b) $2\ NO_2(g)\ \rightleftharpoons\ 2\ NO(g)\ +\ O_2(g)$

   $$k_c = \frac{[NO]^2[O_2]}{[NO_2]^2}$$

   c) $N_2(g)\ +\ 3\ H_2(g)\ \rightleftharpoons\ 2\ NH_3(g)$

   $$k_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

   d) $Na_2SO_4(aq)\ +\ CaCl_2(aq)\ \rightleftharpoons\ 2\ NaCl(aq)\ +\ CaSO_4(s)$

   $$k_c = \frac{1}{[SO_4^{2-}][Ca^{2+}]}$$

2. In an experiment at 200°C, 0.500 mol/L of hydrogen bromide gas is placed in a sealed container and it decomposes into hydrogen gas and bromine gas.

   a) Write the equilibrium equation and law for this reaction.

   $$k_c = \frac{[H_2][Br_2]}{[HBr]^2} \qquad 2HBr_{(g)} \rightleftharpoons H_{2(g)} + Br_{2(g)}$$

   b) The equilibrium concentrations in this system are: $[HBr(g)] = 0.240$ mol/L, $[H_2(g)] = [Br_2(g)] = 0.130$ mol/L. Calculate the equilibrium constant.

   $$k_c = \frac{(0.130\ mol/L)(0.130\ mol/L)}{(0.240\ mol/L)^2} = \boxed{0.293}$$

ICE Tables

1. 1.00 mol of hydrogen gas and 1.00 mol of iodine gas are sealed in a 1.00 L reaction vessel and allowed to react at 450°C. At equilibrium, 1.56 mol of hydrogen iodide gas is present. Calculate K_c for the reaction.

	$H_{2(g)}$	+ $I_{2(g)}$	$\rightleftharpoons$ $2\,HI_{(g)}$
I	1.00	1.00	0
C	-0.78	-0.78	+1.56
E	0.22	0.22	1.56

$$k_c = \frac{[HI]^2}{[H_2][I_2]} = \boxed{50.3}$$

2. In an experiment, 2.00 mol of $H_2(g)$ and 2.00 mol of $F_2(g)$ are introduced into a 1.00 L flask at 500°C. After equilibrium was reached, the concentration of HF(g) was 0.500 mol/L. Calculate the K_c for this reaction at 500°C.

	$H_{2(g)}$	+ $F_{2(g)}$	$\rightleftharpoons$ $2HF_{(g)}$
I	2.00 mol/L	2.00 mol/L	0
C	-0.250	-0.250	+0.500
E	1.75 mol/L	1.75 mol/L	0.500 mol/L

$$k_c = \frac{[HF]^2}{[H_2][F_2]} = \boxed{0.0816}$$

3. Phosphorus pentachloride gas decomposes into phosphorus trichloride gas and chlorine gas. If the $[PCl_5(g)]_i = 8.1 \times 10^{-3}$ mol/L and the $[PCl_3(g)]_i = 0.298$ mol/L, calculate the K_c. The $[Cl_2(g)]_{eq} = 2.00 \times 10^{-3}$ mol/L.

See notes

Graphical Analysis

The Haber-Bosch process of the industrial production of ammonia involves the equilibrium

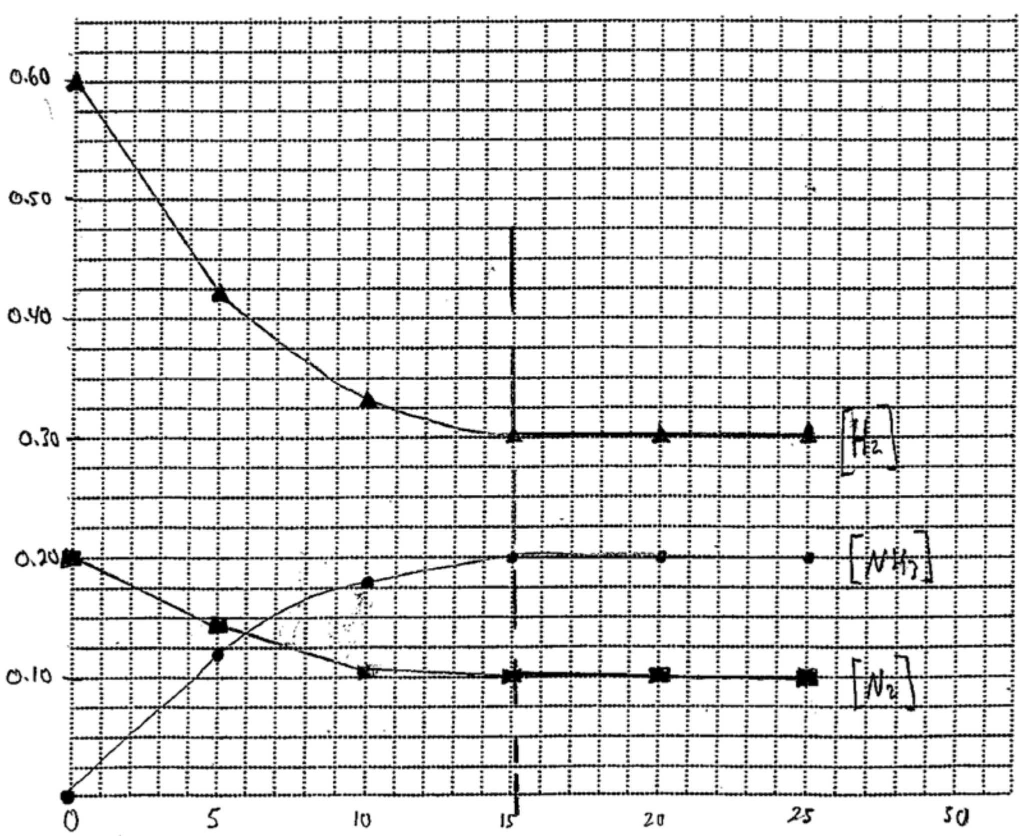
$$N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g) + 92.2\text{ kJ}.$$

In a laboratory experiment designed to study this equilibrium, a chemical engineer injects 0.20 mol of $N_2(g)$ and 0.60 mol of $H_2(g)$ into a 1.0 L flask at 500°C. She records her analysis of the flask at 5 s intervals in the table shown.

Time (s)	Concentration (mol/L)		
	$N_2(g)$	$H_2(g)$	$NH_3(g)$
0	0.20	0.60	0.00
5	0.14	0.42	0.12
10	0.11	0.33	0.18
15	0.10	0.30	0.20
20	0.10	0.30	0.20
25	0.10	0.30	0.20

Analyze the data by:

1. Draw a graph of the concentrations of $N_2(g)$, $H_2(g)$ and $NH_3(g)$ versus time on the graph paper below. Include a legend with your graph.
2. State the time required for equilibrium to be established
3. Calculate the equilibrium constant for this reaction…showing all work.



$$k_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

$= 14.81\ldots$

$= \boxed{15}$

equilibrium established @ 15 s