Key

Chemistry 12

Worksheet for K_{eq} calculations

1. Write the expression for the equilibrium constant for each of the following reactions.

(a) $H_{2(g)} + Cl_{2(g)} \Leftrightarrow 2\,HCl_{(g)}$

$$K_{eq} = \frac{[HCl]^2}{[H_2][Cl_2]}$$

(b) $2\,SO_{2(g)} + O_{2(g)} \Leftrightarrow 2\,SO_{3(g)}$

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

(c) $N_{2(g)} + 3\,H_{2(g)} \Leftrightarrow 2\,NH_{3(g)}$

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

(d) $2\,CO_{(g)} + O_{2(g)} \Leftrightarrow 2\,CO_{2(g)}$

$$K_{eq} = \frac{[CO_2]^2}{[CO]^2[O_2]}$$

(e) $CaCO_{3(s)} \Leftrightarrow CaO_{(s)} + CO_{2(g)}$

$$K_{eq} = [CO_2]$$

(f) $HCN_{(g)} \Leftrightarrow H^+_{(aq)} + CN^-_{(aq)}$

$$K_{eq} = \frac{[CN^-][H^+]}{[HCN]}$$

(g) $H_2SO_{4(l)} \Leftrightarrow H^+_{(aq)} + HSO_4^-{}_{(aq)}$

$$K_{eq} = [H^+][HSO_4^-]$$

(h) $H_2CO_{3(aq)} \Leftrightarrow 2\,H^+_{(aq)} + CO_3^{2-}{}_{(aq)}$

$$K_{eq} = \frac{[CO_3^{2-}][H^+]^2}{[H_2CO_3]}$$

2. From selected equilibrium constant expressions above, calculate the value of the equilibrium constant (K_{eq}).

(a) from 1(a) above: equilibrium $[H_2] = [Cl_2] = [HCl] = 1.0 \times 10^{-2}$ M

$$K_{eq} = \frac{[HCl]^2}{[H_2][Cl_2]} = \frac{(1.0 \times 10^{-2})^2}{(1.0 \times 10^{-2})(1.0 \times 10^{-2})} = 1.0$$

(b) from 1(b) above: equilibrium
$[SO_2] = 1.0 \times 10^{-3}$ M
$[O_2] = 2.0 \times 10^{-3}$ M
$[SO_3] = 3.0 \times 10^{-3}$ M

$$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(3.0 \times 10^{-3})^2}{(1.0 \times 10^{-3})^2(2.0 \times 10^{-3})} = 4500$$

(c) from 1(c) above: equilibrium $[N_2] = 4.4 \times 10^{-2}$ M
$[H_2] = 1.2 \times 10^{-1}$ M
$[NH_3] = 3.4 \times 10^{-3}$ M

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(3.4 \times 10\ )^2}{(4.4 \times 10\ )(1.2 \times 10^{-1})^3} = 0.15$$

3. From 1(a) above, assume the equilibrium constant to be 55.0. During an experiment, the equilibrium $[H_2] = 4.8 \times 10^{-3}$ M and $[Cl_2] = 2.1 \times 10^{-3}$ M. What is the equilibrium [HCl]?

$$K_{eq} = \frac{[HCl]^2}{[H_2][Cl_2]} \quad \therefore [HCl] = \sqrt{K_{eq}[H_2][Cl_2]}$$

$$= \sqrt{(55.0)(4.8 \times 10^{-3})(2.1 \times 10^{-3})}$$

$$= \sqrt{5.544 \times 10\ }$$

$$= 2.4 \times 10^{-2}$$

4. Under a given set of experimental conditions, the reaction;

$$N_{2(g)} + O_{2(g)} \Leftrightarrow 2\,NO_{(g)}$$

has a $K_{eq} = 6.2 \times 10^{-4}$. If the equilibrium $[N_2] = [O_2] = 5.2$ M, then what is the equilibrium [NO]?

$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]} \quad \therefore [NO] = \sqrt{K_{eq}[N_2][O_2]}$$

$$= \sqrt{(6.2 \times 10^{-4})(5.2)(5.2)}$$

$$= \sqrt{(1.6765 \times 10^{-2})}$$

$$= 0.13$$

5. According to 1(b) above, if 0.600 mole of SO_2 and 0.600 mole of O_2 are placed into a 1.00 L container and allowed to establish equilibrium, the equilibrium $[SO_3] = 0.500$ M. Calculate the value of K_{eq}.

Must use an ICE box to calculate the equilibrium values of SO_2 and O_2

	$2\,SO_2$ +	O_2 $\Leftrightarrow$	$2\,SO_3$
I	0.600	0.600	0
C	$-2x$ (= -0.500)	$-x$ (=-0.300)	$+2x$ ($2x = 0.500$)
E	0.100	0.300	0.500

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

$$= \frac{(0.500)^2}{(0.100)^2(0.300)}$$

$$= 83.3$$

6. Under a given set of experimental conditions, the reaction;

$$2\,HI_{(g)} \Leftrightarrow H_{2(g)} + I_{2(g)}$$

has a $K_{eq} = 2.4$. If the initial [HI] = 0.200 M, then what are the equilibrium concentrations of all of the chemicals in the reaction?

	$2\,HI$	$\Leftrightarrow$ H_2	$+$ I_2
I	0.200	0	0
C	$-2x$	$+x$	$+x$
E	$0.200 - 2x$	x	x

$$K_{eq} = \frac{[H_2][I_2]}{[HI]^2}$$

$$2.4 = \frac{(x)(x)}{(0.200 - 2x)^2}$$

$$2.4 = \frac{(x)^2}{(0.200 - 2x)^2}$$

$$\sqrt{2.4 = \frac{(x)^2}{(0.200 - 2x)^2}} \qquad 1.5492 = \frac{(x)}{(0.200 - 2x)}$$

$$(1.5492)(0.200 - 2x) = (x)$$

$$x = 0.3098 - 3.098x$$

$$4.098x = 0.3098$$

$$\frac{4.098x}{4.098} = \frac{\text{[illegible]}}{4.098}$$

$$x = 0.076 \text{ M}$$

$[H_2]$ and $[I_2]$ = 0.076 M. Since $[HI] = 0.200 - 2x$
$= 0.200 - 0.152$
$= 0.048$ M