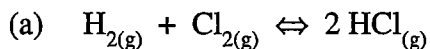


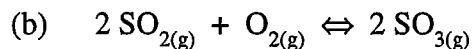
Chemistry 12

Worksheet for K_{eq} calculations

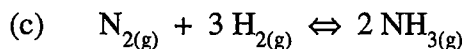
1. Write the expression for the equilibrium constant for each of the following reactions.



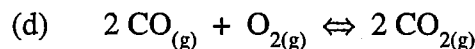
$$K_{eq} = \frac{[\text{HCl}]^2}{[\text{H}_2][\text{Cl}_2]}$$



$$K_{eq} = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]}$$



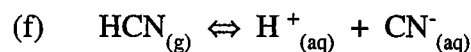
$$K_{eq} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$



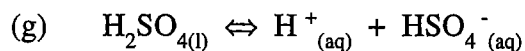
$$K_{eq} = \frac{[\text{CO}_2]^2}{[\text{CO}]^2[\text{O}_2]}$$



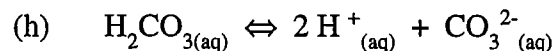
$$K_{eq} = [\text{CO}_2]$$



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



$$K_{eq} = [\text{H}^+][\text{HSO}_4^-]$$



$$K_{eq} = \frac{[\text{CO}_3^{2-}][\text{H}^+]^2}{[\text{H}_2\text{CO}_3]}$$

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

(a) from 1(a) above: equilibrium $[\text{H}_2] = [\text{Cl}_2] = [\text{HCl}] = 1.0 \times 10^{-2} \text{ M}$

$$K_{eq} = \frac{[\text{HCl}]^2}{[\text{H}_2][\text{Cl}_2]} = \frac{(1.0 \times 10^{-2})^2}{(1.0 \times 10^{-2})(1.0 \times 10^{-2})} = 1.0$$

(b) from 1(b) above: equilibrium $[\text{SO}_2] = 1.0 \times 10^{-3} \text{ M}$

$$[\text{O}_2] = 2.0 \times 10^{-3} \text{ M}$$

$$[\text{SO}_3] = 3.0 \times 10^{-3} \text{ M}$$

$$K_{eq} = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]} = \frac{(3.0 \times 10^{-3})^2}{(1.0 \times 10^{-3})^2(2.0 \times 10^{-3})} = 4500$$

(c) from 1(c) above: equilibrium

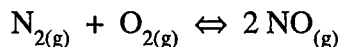
$$\begin{aligned}
 [\text{N}_2] &= 4.4 \times 10^{-2} \text{ M} \\
 [\text{H}_2] &= 1.2 \times 10^{-1} \text{ M} \\
 [\text{NH}_3] &= 3.4 \times 10^{-3} \text{ M}
 \end{aligned}$$

$$K_{eq} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(3.4 \times 10^{-3})^2}{(4.4 \times 10^{-2})(1.2 \times 10^{-1})^3} = 0.15$$

3. From 1(a) above, assume the equilibrium constant to be 55.0. During an experiment, the equilibrium $[\text{H}_2] = 4.8 \times 10^{-3} \text{ M}$ and $[\text{Cl}_2] = 2.1 \times 10^{-3} \text{ M}$. What is the equilibrium $[\text{HCl}]$?

$$\begin{aligned}
 K_{eq} &= \frac{[\text{HCl}]^2}{[\text{H}_2][\text{Cl}_2]} \quad \therefore [\text{HCl}] = \sqrt{K_{eq} [\text{H}_2][\text{Cl}_2]} \\
 &= \sqrt{(55.0)(4.8 \times 10^{-3})(2.1 \times 10^{-3})} \\
 &= \sqrt{5.544 \times 10^{-4}} \\
 &= 2.4 \times 10^{-2}
 \end{aligned}$$

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



has a $K_{eq} = 6.2 \times 10^{-4}$. If the equilibrium $[\text{N}_2] = [\text{O}_2] = 5.2 \text{ M}$, then what is the equilibrium $[\text{NO}]$?

$$\begin{aligned}
 K_{eq} &= \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} \quad \therefore [\text{NO}] = \sqrt{K_{eq} [\text{N}_2][\text{O}_2]} \\
 &= \sqrt{(6.2 \times 10^{-4})(5.2)(5.2)} \\
 &= \sqrt{1.6765 \times 10^{-2}} \\
 &= 0.13
 \end{aligned}$$

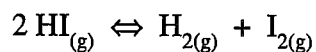
5. According to 1(b) above, if 0.600 mole of SO_2 and 0.600 mole of O_2 are placed into a 1.00 L container and allowed to establish equilibrium, the equilibrium $[\text{SO}_3] = 0.500 \text{ M}$. Calculate the value of K_{eq} .

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

	2SO_2	+	O_2	\rightleftharpoons	2SO_3
I	0.600		0.600		0
C	$-2 \times$ (= -0.500)		$- \times$ (= -0.300)		$+2 \times$ ($2 \times = 0.500$)
E	0.100		0.300		0.500

$$\begin{aligned}
 K_{eq} &= \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]} \\
 &= \frac{(0.500)^2}{(0.100)^2 (0.300)} \\
 &= 83.3
 \end{aligned}$$

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



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

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

$$K_{eq} = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2}$$

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

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

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

$$1.5492 = \frac{(x)}{(0.200 - 2x)}$$

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

$$x = 0.3098 - 3.098x$$

$$4.098x = 0.3098$$

$$\frac{4.098x}{4.098} = \frac{0.3098}{4.098}$$

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

$[\text{H}_2]$ and $[\text{I}_2] = 0.076 \text{ M}$. Since $[\text{HI}] = 0.200 - 2x$
 $= 0.200 - 0.152$
 $= 0.048 \text{ M}$

