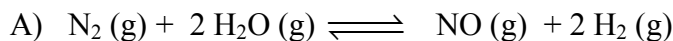


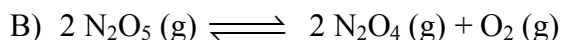
REVIEW QUESTIONS

Chapter 16

1. Write equilibrium constant expressions for each reaction below:

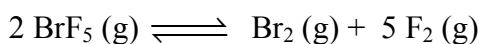


$$K_{\text{eq}} = \frac{[\text{NO}][\text{H}_2]^2}{[\text{N}_2][\text{H}_2\text{O}]^2}$$



$$K_{\text{eq}} = \frac{[\text{N}_2\text{O}_4]^2[\text{O}_2]}{[\text{N}_2\text{O}_5]^2}$$

2. At 1500 K the following equilibrium is established:



At equilibrium,

$$[\text{BrF}_5] = 0.0064 \text{ M}$$

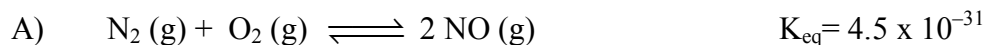
$$[\text{Br}_2] = 0.0018 \text{ M}$$

$$[\text{F}_2] = 0.0090 \text{ M.}$$

Calculate the equilibrium constant for this reaction.

$$K_{\text{eq}} = \frac{[\text{Br}_2][\text{F}_2]^5}{[\text{BrF}_5]^2} = \frac{(0.0018)(0.0090)^5}{(0.0064)^2} = 2.6 \times 10^{-9}$$

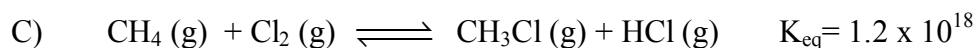
3. For each reaction below, determine whether reactants or products are favored at equilibrium:



Equilibrium constant is small, therefore reactants are favored at equilibrium

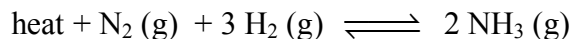


Equilibrium constant is large, therefore products are favored at equilibrium

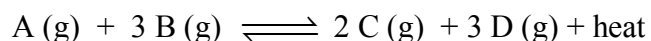


Equilibrium constant is large, therefore products are favored at equilibrium

4. In the equilibrium shown below, identify four changes that would increase the concentration of ammonia equilibrium:



1. **Add N₂ (equilibrium would shift forward and produce more NH₃)**
 2. **Add H₂ (equilibrium would shift forward and produce more NH₃)**
 3. **Increase temperature (equilibrium would shift forward and produce more NH₃)**
 4. **Decrease volume (pressure would increase and equilibrium would shift forward and produce more NH₃)**
5. For the equilibrium shown below, describe the changes that occur after each action listed:



- a) Temperature is increased

Equilibrium would shift reverse; A and B would increase; C and D would decrease.

- b) More A is added

Equilibrium would shift forward; B would decrease; C and D would increase.

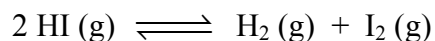
- c) Some D is removed

Equilibrium would shift forward; A and B would decrease; C would increase

- d) Volume is decreased

Equilibrium would shift reverse; A and B would increase; C and D would decrease

6. 0.100 moles of HI was placed in a 1.0-L flask and allowed to decompose as shown below:



At equilibrium, the 0.011 moles of H_2 was present. Calculate the equilibrium constant for this reaction. (Hint: set up a reaction table)

	2 HI (g)	\rightleftharpoons	H₂ (g)	+	I₂ (g)
Initial	0.100		0		0
Change (Δ)	-0.022		+0.011		+0.011
Equilibrium	0.078		0.011		0.011

$$K_{\text{eq}} = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{(0.011 \text{ M})(0.011 \text{ M})}{(0.078 \text{ M})^2} = 2.0 \times 10^{-2} \text{ M}$$

7. A 0.15 M solution of butanoic acid ($\text{HC}_4\text{H}_7\text{O}_2$) has $[\text{H}_3\text{O}^+] = 1.51 \times 10^{-3} \text{ M}$. Calculate K_a for this acid.



$$[\text{H}^+] = [\text{C}_4\text{H}_7\text{O}_2^-] = 1.51 \times 10^{-3} \text{ M}$$

$$K_a = \frac{[\text{H}^+][\text{C}_4\text{H}_7\text{O}_2^-]}{[\text{HC}_4\text{H}_7\text{O}_2]} = \frac{(1.51 \times 10^{-3})^2}{0.15} = 1.5 \times 10^{-5}$$

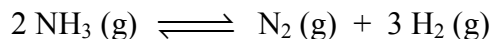
8. A 0.035 M solution of a weak acid (HA) has a pH of 4.88. Calculate the K_a for this acid.



$$[\text{H}^+] = [\text{A}^-] = 10^{-4.88} = 1.32 \times 10^{-5} \text{ M}$$

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{(1.32 \times 10^{-5})^2}{0.035} = 5.0 \times 10^{-9}$$

9. In a reaction, 0.15 moles of NH_3 is placed in a 1.5-L flask and allowed to decompose as shown below:



At equilibrium, 0.035 moles of N_2 were present. Determine the percent yield of this reaction.

$$[\text{NH}_3] = \frac{0.15 \text{ mol}}{1.5 \text{ L}} = 0.10 \text{ M}$$

$$[\text{N}_2] = \frac{0.035 \text{ mol}}{1.5 \text{ L}}$$

	$2 \text{NH}_3 (\text{g})$	\rightleftharpoons	$\text{N}_2 (\text{g})$	+	$3 \text{H}_2 (\text{g})$
Initial	0.10		0		0
Change (Δ)	-----		+0.023		-----
Equilibrium	-----		0.023		-----
No Equil.	0		0.050		-----

$$\% \text{ Yield} = \frac{0.023}{0.050} \times 100 = 46\%$$

10. Determine the pH and percent ionization of 1.0 M acetic acid ($K_a = 1.8 \times 10^{-5}$).



$$[\text{H}^+] = [\text{C}_2\text{H}_3\text{O}_2^-] = x$$

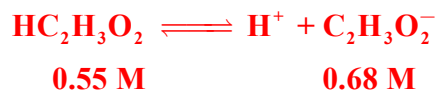
$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = \frac{x^2}{1.0} = 1.8 \times 10^{-5}$$

$$x = \sqrt{1.8 \times 10^{-5}} = 4.24 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log(4.24 \times 10^{-3}) = 2.37$$

$$\% \text{ Ionization} = \frac{[\text{H}^+]}{[\text{HA}]} \times 100 = \frac{4.24 \times 10^{-3}}{1.0} \times 100 = 0.42\%$$

11. What is the pH of a buffer solution consisting of 0.55 M acetic acid and 0.68 M sodium acetate? (K_a for acetic acid = 1.8×10^{-5})



$$[\text{H}^+] = K_a \times \frac{[\text{HA}]}{[\text{A}^-]} = 1.8 \times 10^{-5} \times \frac{0.55}{0.68} = 1.46 \times 10^{-5}$$

$$\text{pH} = -\log(1.46 \times 10^{-5}) = 4.84$$

12. The K_{sp} value for AgI is 8.5×10^{-17} . What is the molar solubility of AgI?



$$K_{sp} = [\text{Ag}^+][\text{I}^-] = 8.5 \times 10^{-17}$$

$$[\text{Ag}^+] = \sqrt{8.5 \times 10^{-17}} = 9.2 \times 10^{-9} \text{ M}$$

13. The solubility of CaF_2 is 2.14×10^{-4} M. Calculate the K_{sp} for CaF_2 .



$$[\text{Ca}^{2+}] = 2.14 \times 10^{-4}$$

$$[\text{F}^-] = 2 \times [\text{Ca}^{2+}] = 4.28 \times 10^{-4} \text{ M}$$

$$K_{sp} = [\text{Ca}^{2+}][\text{F}^-]^2 = (2.14 \times 10^{-4})(4.28 \times 10^{-4})^2 = 3.92 \times 10^{-11}$$