La Medica Academy

Study Smarter Not Harder

General Chemistry Review

Chapter 13 :

Chemical Equilibrium

Mustafa Kharma

Chapter 13 – Chemical Equilibrium

⇒ The Equilibrium Constant:

- **K** is the equilibrium expression in terms of concentrations.
- **Kp** is the equilibrium expression in terms of the partial pressures of the gaseous constituents of the reaction

⇒ Consider The Following Equilibria:

- $H_2O_{(g)} + CO_{(g)} \rightleftharpoons H_{2(g)} + CO_{2(g)}$; K
- $2H_2O_{(g)} + 2CO_{(g)} \rightleftharpoons 2H_{2(g)} + 2CO_{2(g)}$; K'
- $H_{2(g)} + CO_{2(g)} \rightleftharpoons H_2O_{(g)} + CO_{(g)}$ K''
- $K' = K^2$ ⇌ (concentrations doubled)
- $K'' = (1/K)$ ⇌ (reversing the reaction)

⇒ The Relationship Between K and Kp

$$Kp = K(RT)^{\Delta ng}$$

$$\Delta ng = ng\ (P) - ng\ (R)$$

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

Using the value of Kp (3.9×10^4) from the previous example, **calculate the value of K at 35°C.**

$$Kp = K(RT)^{\Delta ng}$$

$$3.9 \times 10^4 = K(0.0821\ L.atm/mol.K \times 308K)^{2-4}$$

$$K = 2.5 \times 10^7$$

Note: Kp < K (Δng is negative)

- $Kp = K(RT)^{\Delta ng}$; $\Delta ng = ng(P) - ng(R)$
- **Kp = K** when Δng **= 0**

$$H_2(g) + I_2(g) \longrightarrow 2HI(g)$$

- **Kp > K** when Δng **> 0**

$$N_2O_2(g) \longrightarrow 2NO_2(g)$$

- If the equilibrium **lies to the right** (toward the products), the value for **K** is **Larger than 1**.
- If the equilibrium **lies to the left** (towards the reactants), the value for **K** is **Smaller than 1.**

⇒ Calculations:

At equilibrium: [C] = 3.0M

	$A_{(g)}$	+ $2B_{(g)}$	$C_{(g)}$
Initially:	8	10	0
Change:	-3	-6	+3
At equilibrium:	5	4	3

$$K = \frac{[C]}{[B]2[A]} = \frac{3}{4^2.5} = 0.0375$$

Calculate K_p at 25 °C?

If the Initial concentrations are: 10.0 M Fe^{3+} and 8.00 M SCN-, **What is the equilibrium concentration of $FeSCN^{2+}$ (K=0.33333) The equilibrium reaction is shown below:**

	$Fe^{3+}{}_{(aq)}$	+ $SCN^{-}{}_{(aq)}$	$FeSCN^{2+}{}_{(aq)}$
Initially:	10.0 M	8.0 M	0
At Equilibrium.:	(10.0 – x)	(8.0 – x)	x

$$K = \frac{x}{(8.0-x)(10.0-x)} = 0.3333$$

5.00 *M* $FeSCN^{2+}$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

It is found that x=5M so, At eq.: $[Fe^{3+}]$ = 10 – 5 = 5, [SCN] = 8 – 5 = 3 , [FeSCN] = 5

⇒ Reaction Quotient, Q:

- **Q = K**; The system is at equilibrium. No shift will occur.
- **Q > K**; The system shifts to the **left**. **After equilibrium** ←
- Consuming products and forming reactants, until equilibrium is achieved.
- **Q < K**; The system shifts to the **right**. **Before equilibrium** →
- Consuming reactants and forming products, to attain equilibrium.

- **Consider the reaction represented by the equation:**

$$Fe^{+3} + SCN^{-} \longrightarrow FeSCN^{+2}$$

- **Consider the following initial concentrations**

	Fe^{3+}	SCN^{-}	$FeSCN^{2+}$	Q
Q1:	9.00 M	5.00 M	1.00 M	**0.0222**
Q2:	3.00 M	2.00 M	5.00 M	**0.8333**
Q3:	2.00 M	9.00 M	6.00 M	**0.3333**

K = 0.3333

Find the equilibrium concentrations for all species.

⇒ Le Châtelier's Principle:

1. **Concentration:** The system will shift away upon addition of any component that is a part of the system. If a component is removed, the opposite effect occurs.
2. **Temperature:** K will change depending upon the temperature. (endothermic – energy is a reactant; exothermic – energy is a product).
3. **Pressure effect at constant temperature**, pressure that is due to a volume change (volume pressure): Favor the reaction that decreases the number of gas molecules.
4. **Addition of inert gas** does not affect the equilibrium position.