Chemical Equilibrium Video 1 Review Sheet

Description of Acid/Base Video7: "Chemical Equilibrium Video 1":
This essential review material will focus on how to describe and define an equilibrium state, followed by an introduction to the law of mass action, predicting shifts in equilibrium, and expressing the change by examining the reaction quotient, Q.

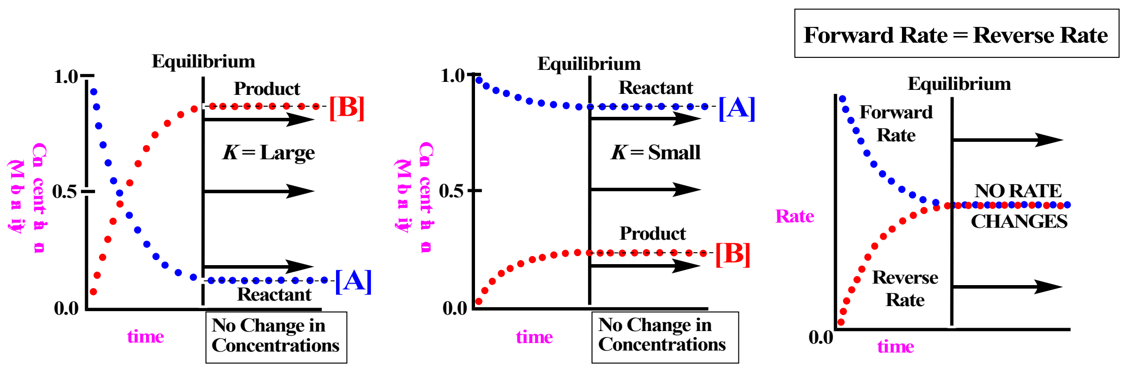
1) There are two common ways to state that a chemical reaction is at equilibrium, what are they?

First, the chemical reaction is indicated by an equilibrium set of arrows (as opposed to a single arrow, as shown below), which means the reaction is at equilibrium when: a) there is no change in concentrations of reactants and products and b) the forward rate is equal to the reverse rate of reaction as shown below:

$$\underline{a}A + \underline{b}B \longrightarrow \underline{c}C + \underline{d}D$$

$$\underline{a}A + \underline{b}B \rightleftharpoons \underline{c}C + \underline{d}D$$

(equilibrium arrows)



2) The Law of Mass Action indicates the extent of an equilibrium and often one is able to predict if the equilibrium lies toward reactants or products if given the equilibrium constant. How is this possible?

The law of mass action is a ratio of concentrations of products over reactants, or more specifically, the product of the product concentrations raised to their stoichiometric coefficients divided by the product of the concentrations of the reactants raised to their stoichiometric coefficients as shown below:

$$\underline{a}A + \underline{b}B \rightleftharpoons \underline{c}C + \underline{d}D \qquad K = \frac{[\text{Products}]}{[\text{Reactants}]} \qquad K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

A large value for K indicates that the concentrations of products is large (the numerator) and the concentrations of reactants is small, which is in the denominator (equilibrium lies to the right, toward products). A small value for K indicates that the concentrations of products is small (the numerator) and the concentrations of reactants is large, which is in the denominator (equilibrium lies to the left, toward reactants) as shown below. Remember, that K has <u>NO</u> units and that <u>NO</u> pure liquids or pure solids are included in the equilibrium expression.

$$\text{LARGE } K = \frac{\uparrow [\text{Products}]}{\downarrow [\text{Reactants}]} \qquad \text{SMALL } K = \frac{\downarrow [\text{Products}]}{\uparrow [\text{Reactants}]}$$

$$\underline{a}A + \underline{b}B \rightleftharpoons \underline{c}C + \underline{d}D \qquad \underline{a}A + \underline{b}B \rightleftharpoons \underline{c}C + \underline{d}D$$

3) The Haber reaction is an equilibrium when hydrogen gas reacts with nitrogen gas to afford ammonia gas. At a certain temperature the equilibrium concentrations were measured and found to be $[H_2] = 0.763\ M$, $[N_2] = 0.921\ M$, and $[NH_3] = 0.157\ M$. What is the equilibrium constant at this temperature?

Write the balanced equation for the Haber reaction (we will need the stoichiometric coefficients within the law of mass action), plug the equilibrium concentrations into the law of mass action, and calculate the equilibrium constant (note: NO units required for equilibrium constants).

$$N_{2\,(g)} + \underline{3}H_{2\,(g)} \rightleftharpoons \underline{2}NH_{3\,(g)} \qquad K = \frac{[NH_3]^2}{[N_2][H_2]^3} \qquad K = \frac{[0.157]^2}{[0.921][0.763]^3} = \boxed{\underline{0.0602}}$$

4) The Haber reaction is an equilibrium when hydrogen gas reacts with nitrogen gas to afford ammonia gas. At 300^0C the equilibrium constant was found to be K = 9.60. Please convert this to a K_p value at 300^0C.

Write the balanced equation for the Haber reaction (we will need the stoichiometric coefficients within the law of mass action), calculate change in moles (Δn), convert ^{0}C to K, and plug in the given K value as shown:

$$N_{2\,(g)} + \underline{3}H_{2\,(g)} \overset{K=9.60}{\rightleftharpoons} \underline{2}NH_{3\,(g)}$$

Δn = Sum mol of products - Sum mol of reactants
= 2 mol NH_3 - (1 mol N_2 + 3 mol H_2)
= $\underline{(-)\ 2}$

T = 300^0C + 273 = $\underline{573\text{K}}$

$K_p = K_c(RT)^{\Delta n}$
= $(9.60)(0.0821 \times 573)^{-2}$
= $\underline{4.3\times10^{-3}}$

5) The Haber reaction is an equilibrium when hydrogen gas reacts with nitrogen gas to afford ammonia gas. At 472^0C the equilibrium partial pressures were measured and found to be $[H_2]$ = 7.38 atm, $[N_2]$ = 2.46 atm, and $[NH_3]$ = 0.166 atm. What is the K_p and K at this temperature?

$$N_{2\,(g)} + \underline{3}H_{2\,(g)} \rightleftharpoons \underline{2}NH_{3\,(g)} \qquad K_p = \frac{[NH_3]^2}{[N_2][H_2]^3} \qquad K_p = \frac{(0.166\text{ atm})^2}{(2.46\text{ atm})(7.38\text{ atm})^3} = \boxed{\underline{2.79\times10^{-5}}}$$

K_p to K Calculation:

$K_p = K_c(RT)^{\Delta n}$

Δn = Sum mol of products - Sum mol of reactants
= 2 mol NH_3 - (1 mol N_2 + 3 mol H_2)
= $\underline{(-)\ 2}$

T = 472^0C + 273 = $\underline{745\text{K}}$

$2.79\times10^{-5} = K_c(0.0821 \times 745)^{-2}$

$K_c = \boxed{\underline{0.104}}$

6) Please explain what the reaction quotient, Q, is equal to, and explain shift in equilibrium if Q = K, Q > K, and if Q < K.

The reaction quotient, Q, is calculated using the initial concentrations within the law of mass action. If Q = K, then there will be no change in concentrations and the system is at equilibrium. However, if Q > K, then there will be a shift in equilibrium to the left (toward reactants), which will cause the concentration of products to decrease and the concentration of reactants to increase. It may help to abstractly think that "Q will want to become K" and for this to happen the numerator must become smaller (concentration of products) and the denominator will want to become larger (concentration of reactants) as shown below:

$$\underline{a}A + \underline{b}B \rightleftharpoons \underline{c}C + \underline{d}D \qquad \boxed{IF\ Q > K} \qquad \text{"Q wants to become K"} \qquad Q = \frac{\downarrow [\text{Products}]}{\uparrow [\text{Reactants}]}$$

$\uparrow$[A] $\uparrow$[B] $\downarrow$[C] $\downarrow$[D]

Shift to left $\Longleftarrow$

If Q < K, then there will be a shift in equilibrium to the right (toward products), which will cause the concentration of products to increase and the concentration of reactants to decrease as shown below:

$$\underline{a}A + \underline{b}B \rightleftharpoons \underline{c}C + \underline{d}D \qquad \boxed{IF\ Q < K} \qquad \text{"Q wants to become K"} \qquad Q = \frac{\downarrow [\text{Products}]}{\uparrow [\text{Reactants}]}$$

$\downarrow$[A] $\downarrow$[B] $\uparrow$[C] $\uparrow$[D]

Shift to right $\Longrightarrow$

7) Gaseous hydrogen and gaseous iodine are in equilibrium with gaseous hydrogen iodide (also known as hydroiodic acid). The equilibrium constant was measured and found to be 50.5 at 448^{0}C. Random amounts of reactants and product ([H_2] = 1.00x10^{-2} mol, [I_2] = 3.00x10^{-2} mol, and [HI] = 2.0x10^{-2} mol) were placed into a 2.00L reaction vessel and allowed to obtain an equilibrium state at 448^{0}C. Please predict if the given initial quantities will increase or decrease at equilibrium.

Write the balanced equation, place the given equilibrium constant over the equilibrium arrows, place the given quantities under each reactant and product, convert to concentrations by dividing by the volume of the reaction vessel, and calculate the Q value as shown below. Q < K; thus, a shift to right, which will affect concentrations as shown.

$$H_{2\,(g)} + I_{2\,(g)} \overset{K = 50.5}{\rightleftharpoons} \underline{2}HI_{\,(g)} \quad \text{at } 448\,^0\text{C}$$

$$\frac{1.00\text{x}10^{-2}\text{ mol}}{2.00\text{L}} = 5.00\text{x}10^{-3}\,M \qquad \frac{3.00\text{x}10^{-2}\text{ mol}}{2.00\text{L}} = 1.50\text{x}10^{-2}\,M \qquad \frac{2.0\text{x}10^{-2}\text{ mol}}{2.00\text{L}} = 1.00\text{x}10^{-2}\,M$$

$\downarrow$[H_2] $\downarrow$[I_2] $\uparrow$[HI]

$$Q = \frac{(1.00\text{x}10^{-2})^2}{(5.00\text{x}10^{-3})(1.50\text{x}10^{-2})} = \underline{1.30}$$

$$\boxed{Q < K}$$

1.30 < 50.5

Shift to right $\Longrightarrow$

8) Please indicate if shift will be to left or right and express the change for the Haber reaction at 773 ^{0}C (K=0.0602) given the following concentrations.

a) [H_2] = 0.38 *M*, [N_2] = 2.4 *M*, and [NH_3] = 0.0 *M*

b) [H_2] = 0.0 *M*, [N_2] = 3.8 *M*, and [NH_3] = 9.3 *M*

c) [H_2] = 1.20 *M*, [N_2] = 0.40 *M*, and [NH_3] = 0.20 *M*

For both a and b the shift will be toward the zero quantity.

a) The change means nitrogen and hydrogen will decrease (loss) and the ammonia concentration will increase (gain). From the stoichiometry of the balanced equation we know that for every two moles of ammonia that form (+2x), one mole of nitrogen and three moles of hydrogen will be lost, which we can write as our change as shown below.

b) The change means nitrogen and hydrogen concentrations will increase (gain) and the ammonia concentration will decrease (loss). From the stoichiometry of the balanced equation we know that for every three moles of hydrogen that form (+3x), one mole of nitrogen will form (+x), and two moles of ammonia will be lost (-2x), which we can write as our change as shown below.

a) $K = 0.0602$

a)	$N_{2\,(g)}$ +	$\underline{3}H_{2\,(g)}$ $\rightleftharpoons$	$\underline{2}NH_{3\,(g)}$
Initial:	**2.4 *M***	**0.38 *M***	**0.0 M**
	Shift to right ⟹		
Equilibrium:	**2.4-x**	**0.38-3x**	**+ 2x**

b) $K = 0.0602$

b)	$N_{2\,(g)}$ +	$\underline{3}H_{2\,(g)}$ $\rightleftharpoons$	$\underline{2}NH_{3\,(g)}$
Initial:	**3.8 *M***	**0.0 *M***	**9.3 M**
	⟸ **Shift to left**		
Equilibrium:	**3.8+x**	**+3x**	**9.3-2x**

c) A Q calculation with these random concentrations is first required to predict shift and Q < K (0.058 < 0.0602); thus, will shift to right because our Q value is less than K. That means the concentrations of nitrogen and hydrogen will decrease and the ammonia concentration will increase. From the stoichiometry of the balanced equation we know that for every two moles of ammonia that form (+2x), one mole of nitrogen (-x) and three moles of hydrogen (-3x) will be lost, which we can write as our change.

c) $K = 0.0602$

c)	$N_{2\,(g)}$ +	$\underline{3}H_{2\,(g)}$ $\rightleftharpoons$	$\underline{2}NH_{3\,(g)}$
Initial:	**0.40 *M***	**1.20 *M***	**0.20 M**
	Shift to right ⟹		
Equilibrium:	**0.40-x**	**1.20-3x**	**0.20+ 2x**

$$Q = \frac{[0.20]^2}{[0.40][1.20]^3} = \underline{0.058}$$

$$Q < K$$

$$0.058 < 0.0602$$