Notes for Chapter 13 Equilibrium Part I

Equilibrium is a state that is reached in a reaction in which the rate of the forward reaction is equal to the rate of the reverse reaction.

As you have seen in the video, many reactions do not go forward only but reverse back to the reactants.

This situation is referred to as **DYNAMIC** meaning it is constantly moving back and forth but no movement in the reaction is visible. It is called **dynamic equilibrium** because no apparent change in the amount of reactants or products are apparent but the forward and reverse reactions are still occurring.

Example: $2\ NO_2\ (g) \leftarrow\rightarrow N_2O_4\ (g)$ (Double arrows represent a reversible reaction.)

The above reaction is a Homogeneous reaction because all the reactants and products are in the same state. We will be also working with Heterogeneous reactions (not seen in the video).

Note: In the video, gas states were treated as [molar concentrations] which is very common. However, they can also be treated as (Partial Pressures) usually in atm. units. Both molar concentrations and partial pressures can be used to find the equilibrium constant (K). When concentrations are used to find the equilibrium constant, it is $\mathbf{K_C}$. When partial pressures are used to find the equilibrium constant, it is $\mathbf{K_P}$. So, the K_c can be found using the number of moles of each gas in a certain volume (molar concentration) or using the molarity of an aqueous solution of different reactants or products in a reaction. **(I will give examples of this later in the notes.)**

The equilibrium constant (K) for a reaction at a certain temperature is equal to the product of the "products" divided by the product of the "reactants" each raised to the power of their coefficients. The same is true for the Reaction Quotient (Q) for a reaction that may or may not have reached equilibrium. When Q = K, the reaction is at equilibrium. When Q < K, the reaction is moving toward the products (right) to make more products and reach equilibrium. When Q > K, the reaction is moving toward the reactants (left) to make more reactants and reach equilibrium.

For heterogeneous reactions, solids(s) and liquids (l) have a numerical value of "1", aqueous substances (aq) have a numerical value equal to their molar concentration [M], and gaseous substances (g) have a numerical value equal to their molar concentrations [M] or their partial pressures (P).

Examples:
Write the equilibrium constant K_c expression for the following reactants:
If the reaction is not at equilibrium, write the Q expression for these reactions:

$3\ O_2\ (g) \leftarrow\rightarrow 2\ O_3\ (g) \quad K_c = \dfrac{[O_3]^2}{[O_2]^3} \quad Q = \dfrac{[O_3]^2}{[O_2]^3}$

$N_2\ (g) + 3\ H_2\ (g) \leftarrow\rightarrow 2\ NH_3(g) \quad K_c = \dfrac{[NH_3]^2}{[N_2]\ [H_2]^3} \quad Q = \dfrac{[NH_3]^2}{[N_2]\ [H_2]^3}$

$S_8\ (s) + 8\ O_2\ (g) \leftarrow\rightarrow 8\ SO_2\ (g) \quad K_c = \dfrac{[SO_2]^8}{(1)\ [O_2]^8} \quad Q = \dfrac{[SO_2]^8}{(1)\ [O_2]^8}$

Likewise, the K_P for the above reaction would be

$$K_P = \frac{(P_{O3})^2}{(P_{O2})^3} \qquad K_P = \frac{(P_{NH3})^2}{(P_{N2})\,(P_{H2})^3} \qquad K_P = \frac{(P_{SO2})^8}{(1)\,(P_{O2})^8}$$

Remember the reaction quotient (Q) is found the same way the equilibrium constant (K) is found but Q isn't always the same. It moves toward equilibrium. The equilibrium constant (K) is always the same for a reaction at a particular temperature.

The value of Q changes as a reaction approaches equilibrium.

1) Pure reactants mixed: Q = 0 because there are no products formed at the moment of mixing so products are 0 and then Q has to be 0.
2) Reaction proceeds: the value of Q increases as the concentration of the products increase and reactants decrease.
3) Reaction reaches equilibrium and Q no longer changes because the concentrations do not change.
4) The value of Q for the reaction is now called the equilibrium constant, K. for the reaction.

Equilibrium Constant. K

The magnitude of K is a measure of the yield of a reversible reaction.

1) If K > 1, its value indicates that at equilibrium, more reactants are converted to products favoring the products.
2) If K < 1, its value indicates that at equilibrium, less of the reactants are converted to products favoring the reactants.

Reactions reach equilibrium only if reactants and products cannot escape from the reaction mixture. As long as a product is pulled out of a reaction (for example a precipitate forms and falls out of solution or a gas forms in an open system and escapes) then Q will stay less than K and will not reach equilibrium.

Calculating an equilibrium constant. K_C

For the reaction

$$I_2\,(aq) + I^-\,(aq) \leftrightarrow I_3^-\,(aq)$$

Equilibrium concentrations of the reactants and products are 6.64×10^{-4} M for I_2 and I^{-1} and 3.36×10^{-4} M for I_3^-. What is the equilibrium constant?

$$K_C = \frac{[I_3^-]}{[I_2]\,[I^-]} = \frac{[3.36 \times 10^{-4}]}{[6.64 \times 10^{-4}]^2} = 776\ M^{-1}$$

Calculating the equilibrium concentration of a product or reactant when the others are known.

N_2 (g) + O_2 (g) $\leftrightarrow$ 2 NO (g)

K_C for this reaction is 4.1×10^{-4} (no units because all "M" cancel)
The concentration of N_2 is 0.036 M and of O_2 is 0.0089 M at equilibrium.
What is the concentration of NO at equilibrium?

$$K_C = \frac{[NO]^2}{[N_2]\,[O_2]} \qquad 4.1 \times 10^{-4} = \frac{(X)^2}{(.036)(.0089)}$$

Cross multiply and take the square root to solve for X.
$X = 3.60 \times 10^{-4}$ M

Calculating equilibrium concentrations when only initial concentrations of reactants and K_C is known.

We are now going to learn how to construct a RICE (or RICEE) BOX.

As see in the video RICE stands for Reaction, Initial con., Change in con., and Equilibrium concentrations. Let's take the above reaction of nitrogen and oxygen producing nitrogen monoxide with the initial concentrations of the reactants both 1.00×10^{-3} M and the K_C being 4.1×10^{-4}.

R	**N_2**	**+**	**O_2**	**$\leftrightarrow$**	**2 NO**
I	**($1.00x10^{-3}$)**		**($1.00x10^{-3}$)**		**0 .**
C	**-X**		**-X**		**+2X .**
E	**($1.0x10^{-3}$- X)**		**($1.0x10^{-3}$-X)**		**2X .**

(2X is added because 2 times as much product is made.)

E <u>This is where I plug the value for X back in to get the actual values.</u>

Now I do the math and solve for X. $4.1x10^{-4} = \dfrac{(2X)^2}{(1.0x10^{-3}-X)(1.0x10^{-3}-X)}$

Since top and bottom of the equation is squared, we could take the square root. However, many times this is not the case and the reactants have different concentrations. We would have to foil the bottom then use the quadratic equation.
But! If the $K_C \leq 10^{-4}$ so little product is formed that the –X 's in the C of the RICE Box can be dropped because they are insignificant. (a majority of K_C values fall Into this range)

So now we have an easier way to solve for X. $4.1x10^{-4} = \dfrac{(2X)^2}{(1.0x10^{-3})(1.0x10^{-3})} = \dfrac{4X^2}{(1.0x10^{-3})^2}$ **$X = 1.01x10^{-5}$M**
(Cross multiply and square root.)
Plug X back in and we have **[NO] = 2($1.0x10^{-5}$)**
$= 2.0X10^{-5}$ M

Equilibrium Constant changes as equation changes.

If I have the K_C (or K_P) for a reaction and I change the reaction, the K_C changes also.

For the following reaction

$2\ SO_3 \leftrightarrow 2\ SO_2 + O_2$ $K_C = .230$

I change the reaction to

$SO_3 \leftrightarrow SO_2 + \frac{1}{2}\ O_2$ K_C = square root of .230 = 0.480 (halved the moles)

I change reaction to

$2\ SO_2 + O_2 \leftrightarrow 2\ SO_3$ $K_C = \frac{1}{.230} = 4.35$ (reversed the rxn)

I change the reaction to **$4\ SO_3 \leftrightarrow 4\ SO_2 + 2\ O_2$ $K_C = (.230)^2$ (doubled the moles)**

Relation between K_C and K_P

$K_P = K_C\,(RT)^{\Delta n}$ **Where Δn is the number of gas moles in the products minus the number of gas moles in the reactants.**
R is the gas constant for atm (.0821) and T is the kelvin temperature.
If you know K_C, you can find K_P and if you know K_P, you can find K_C.

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