Equilibrium Guidebook + WS Review

Darren Nathaniel Khosma

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1 Introduction

So yeah long story short I got cooked in the WS, this is a redemption arc and I am doing this out of my ego. Anyways yeah take this with a grain of salt because I am not the best in chemistry. Without a further ado, I'll start writing, you start reading.

2 Equilibrium Guidebook

2.1 Defenition

Chemical equilibrium is like a "balance point" in a reaction. It's when the reaction is happening, but you don't see any more changes because the forward and backward parts of the reaction are happening at the same speed. Imagine pouring water into a cup. At first, the cup fills up quickly.

- At first, the reaction goes forward really fast because there are lots of reactants (the starting materials). As the cup gets full, less water fits in, so pouring slows.

- Over time, the reactants are used up, so the forward reaction slows down. As the cup gets full, less water fits in, so pouring slows.

- At the same time, the products (what the reaction makes) start turning back into reactants (reverse reaction). This happens faster as there are more products.

- Equilibrium happens when the rates of pouring and spilling are equal. The cup looks full, but water is still moving in and out. The amount of water in the cup stays the same, but there's movement you don't see.

Key points are that the forward and reverse reactions don't stop, they just balance each other. And the amounts of reactants and products stay constant but might not be equal.

2.2 Types of Equilibrium

2.2.1 Chemical Equilibrium

Where the reactants and products balance each other as their forward and reverse reactions happen at the same rate. For example,

$$H_2 + I_2 \longleftrightarrow 2HI$$

This you can see, the coefficients of H_2 and I_2 when added together will be the same as the coefficient of HI

2.2.2 Physical Equilibrium

This involves physical changes, such as changes in state (solid, liquid, gas) or dissolution. No new substances are formed, only the state or distribution of substances changes. For example,

$$H_2O(l) \longleftrightarrow H_2O(g)$$

Water in a closed container can evaporate to form vapor and condense back into liquid. At equilibrium, water molecules evaporate at the same rate as vapor molecules condense, keeping the amount of liquid and vapor constant.

2.2.3 Homogeneous Equilibrium

All the reactants and products are in the same phase (state of matter). This phase is usually either all gases or all liquids. For example,

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

Where the reactants N_2 and H_2 are both gases, and product NH_3 is a gas.

2.2.4 Heterogeneous Equilibrium

Reactants and/or products are in different phases (solid, liquid, gas, or aqueous). The equilibrium only involves concentrations of substances in the gas or aqueous phase because the concentration of solids and liquids remains constant. For example,

$$CaCO_3(s) \longleftrightarrow CaO(s) + CO_2(g)$$

Both $CaCO_3$ and CaO are solids, so they are ignored, hence in calculations such as K_c and K_p, only the CO_2 is counted.

2.3 Law of Mass Action

2.3.1 K_c (Equilibrium Constant)

K_c is the equilibrium constant that uses the concentrations of reactants and products in a chemical reaction to describe the balance at equilibrium. **Do remember that only (aq) or (g) are counted in all these formulas**. In this chemical reaction, the formulas are

$$aA + bB \longleftrightarrow cC + dD$$

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

This is only used in equilibrium, so in questions, solve until equilibrium if it is asked for K_c, it can be simplified to, by the way use the Molarity (M), where

$$M = \frac{n}{V}$$

n is the mole, V is the volume in litres.

$$K_c = (\frac{[C][D]}{[A][B]})^{coefficient}$$

That, of course is assuming the coefficient for all products are the same, for unit, it depends on the coefficients of the equation, it is $M^{(c+d)-(b+a)}$, basically product coefficient minus the reactant's coefficient, such as M^{-2}, an example is this

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

HI = 0.8 mol/L
H_2 = 0.2 mol/L
I_2 = 0.2 mol/L
Putting this on the formula,

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

$$K_c = \frac{0.8^2}{0.2 * 0.2}$$

$$K_c = \frac{0.64}{0.04} = 64M^0$$

$$K_c = 16$$

2.3.2 Q (Quotient)

Q is the reaction quotient, the formula is like K_c but it's calculated using the current concentrations of reactants and products at any point during the reaction, not just at equilibrium. Again it uses M.

$$aA + bB \longleftrightarrow cC + dD$$

$$Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

There are some relations with equilibrium and K_c, if

$$Q < K_C$$

Reaction will move to the right, causing more products to be formed.

$$Q > K_c$$

Reaction will move to the left, causing more reactants to be formed

$$Q = K_C$$

The system, the reaction is in equilibrium. For example, let us assume that our K_c is the same one from the last example, 64 and the reaction is the same

$$H_2 + I_2 \longleftrightarrow 2HI$$

$[H_2] = 0.5$ mol/L
$[I_2] = 0.4$ mol/L
$[2HI] = 0.3$ mol/L
Putting this on the formula,

$$Q = \frac{[HI]^2}{[H_2][I_2]}$$

$$Q = \frac{0.5^2}{0.4 * 0.3}$$

$$Q = \frac{0.25}{0.12} \approx 2.08$$

since Q is smaller than K_c, it will shift the equilibrium to the right, making more products.

2.3.3 K_p (Pressure Equilibrium Constant)

K_p is the equilibrium constant expressed in terms of the partial pressures of gases in a reaction at equilibrium. In a chemical reaction where,

$$aA(g) + bB(g) \longleftrightarrow cC(g) + dD(g)$$

The formula is the same as K_c, except it is using partial pressure, but first the formula for partial pressure is. **Remember to use the moles you got from the equilibrium**,

$$P_a = \frac{n_a}{n_{total}} * P_{total}$$

Now the formula for K_p

$$K_p = \frac{[P_c]^c[P_d]^d}{[P_a]^a[P_b]^b}$$

Now for the example,

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

If the moles of the reactant and products are,
$N_2 = 5$

$O_2 = 3$
$NO = 2$
First find the partial pressure

$$P_{N_2} = \frac{5}{10} * 10 = 5atm$$
$$P_{O_2} = \frac{3}{10} * 10 = 3atm$$
$$P_{NO} = \frac{2}{10} * 10 = 2atm$$

Time to find the K_p

$$K_p = \frac{[P_{NO}]^2}{[P_{N_2}][P_{O_2}]}$$
$$K_p = \frac{2^2}{5 * 3}$$
$$K_p = \frac{4}{15} \approx 0.267$$

2.3.4 K_c in relation with K_p

The equilibrium constant are related to each other using a form of the $pV = nRT$ formula, where T is temperature expressed in Kelvin, to convert from Celsius to Kelvin, add 273 (273.15 actually, but just use 273), and the R is 0.082 , a constant.

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

Δn is the difference between the right sum of the coefficient minus the left sum of the coefficient, for example

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

The Δn is 2 - (1 + 3), where it is -2. Sorry for the bad numbers there, I use ChatGPT for these examples, anyways example question now

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

Where T = 27 C (27 + 273 = 300 K), $K_c = 0.1$, using the formula

$$K_P = K_c(RT)^{\Delta n}$$
$$K_p = 0.1(0.082 * 300)^{-2}$$
$$K_p = 0.1 * (24.6)^{-2}$$
$$K_p = \frac{0.1}{24.6^2}$$
$$K_p = \frac{0.1}{605.16} \approx 0.0001652455549$$

2.4 α (Dissociation Equilibrium)

Dissociation equilibrium refers to the state in which a compound breaks down into its individual components (atoms, molecules, or ions) and the rate of the dissociation process equals the rate of the recombination process, resulting in no net change in the concentrations of the species involved, it uses the M-R-Stb thing (I can't make one rn, my mouse is ass)

$$\alpha = \frac{R}{M}$$

2.5 Le Châtelier's Principle

Le Châtelier's Principle states that if an equilibrium is disturbed by changing the conditions (such as concentration, temperature, or pressure), the system will adjust to counteract the change and restore equilibrium. Basic laws include,

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

$Product \downarrow$ equilibrium shift left, counter acting the product going fewer hence reducing the reactants, $Reactant \downarrow$ equilibrium, the product will be reduced to counter act the reactant.
$P \downarrow$, $V \uparrow$ = Equilibrium goes to the bigger coefficient side
$P \uparrow$, $V \downarrow$ = Equilibrium goes to the lesser coefficient side

$$2NO + O_2 \longleftrightarrow 2NO_2 \Delta H = -27kJ$$

First you have to determine if it's exothermic or endothermic, since its - it releases heat, hence its exothermic.

$T \uparrow$ = Equilibrium shifting endothermic
$T \downarrow$ = Equilibrium shifting exothermic
In this case since it's a exothermic reaction, lowering the heat will cause more of the $2NO_2$ to be formed, and vice versa.

3 WS Review

3.1 PG

1. Chemical equilibrium is a state in a chemical reaction when the concentration of products and reactants does not change over time. The following are the characteristics of an equilibrium reaction, except … .
   a. reversible reaction
   b. the rate of the reaction to the left is equal to the rate of the reaction to the right
   c. no macroscopic changes occurred
   d. Irreversible reaction
   e. happens in an enclosed space

The answer is **D**, because all equilibrium reactions can be reversed.

2. Equilibrium reaction
   $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$ $\Delta H = -27$ kJ
   If the volume is increased, the equilibrium will shift to ....
   a. right, the amount of NO2 decreases
   b. right, the amount of NO2 increases
   c. right, the amount of NO increases
   d. left, the amount of NO increases
   e. left, the amount of NO decreases

Reactant side have a total of 3 as their coefficient, product have 2 as their total coefficient. According to the Le Châtelier's Principle, if the volume is increased, the equilibrium will shift to the side with the larger total coefficient, in which this case it's the reactants, since the reactants are increased, hence NO is increased, hence the answer is **D**.

3. A gas equilibrium system has the following reaction equation:
   $2AB2(g) + B2(g) \rightleftharpoons 2AB3(g)$ $\Delta H = -X$ kJ/mol
   If the temperature of the system is increased, the equilibrium system will shift towards....
   a. right, because it will shift towards the smaller number of moles
   b. right, because it shifts towards the exothermic direction
   c. left, because it shifts towards the exothermic direction
   d. left, because it shifts towards the larger number of moles
   e. left, because it shifts towards the endothermic direction

This reaction is a exothermic reaction, it can be seen by the - sign. Because this is a exothermic reaction, increasing the temperature will shift it to the reactants, because to make it a product, it must be a exothermic reaction, according to the Le Châtelier's Principle, increasing the temperature will make it a more endothermic as a reaction, hence the answer is **E**.

4. In a 1-liter space, there is an equilibrium reaction:
   $2\ HI(g) \rightleftarrows H2(g) + I2(g)$
   If initially there are 0.4 moles of HI, and 0.1 moles of hydrogen gas are obtained at equilibrium, then the degree of dissociation of HI is … .
   a. 0,25
   b. 0,50
   c. 0,60
   d. 0,75
   e. 0.80

Using the M-R-Stb table, you will find that the R of the HI is 0,2 moles, to find the degree of dissociation of HI it is.

$$\alpha = \frac{R}{M}$$

$$\alpha = \frac{0.2}{0.4} = 0.5$$

Hence the answer is **B**.

5. Observe the equilibrium that occurs at 300 K as follows:
   $SCl_2(g) + 2C_2H_4(g) \rightleftharpoons S(CH_2CH_2Cl)_2(g)$
   In a 1 L container, there are 0.6 mol of SCl_2 and 0.3 mol of C_2H_4. If at equilibrium there are 0.1 mol of $S(CH_2CH_2Cl)_2$, the equilibrium constant, Kc, for the reaction is ....
   a. 5
   b. 10
   c. 15
   d. 20
   e. 25

Since it's all gas, we must count it all.

$$SCL_2 + 2C_2H_4 \longleftrightarrow S(CH_2CH_2Cl)_2$$

First we make the M-R-Stb table to make it all in equilibrium, and if you did it correctly it should be,
$SCL_2 = 0.5$
$C_2H_4 = 0.1$
$S(CH_2CH_2Cl)_2 = 0.1$

Put this on the formula (since it's 1 L its the same)

$$K_c = \frac{[S(CH_2CH_2Cl)_2]}{[SCl_2][C_2H_4]^2}$$

$$K_c = \frac{0.1}{0.5 * 0.1^2}$$

$$K_c = \frac{0.1}{0.005} = 20M^{-2}$$

Hence the answer is **D**.

6. A sample of 1.00 mol of NOCl gas is placed in a 2 L container at a temperature of 227 °C until equilibrium is reached. At that state, it is known that there are 0.056 mol of Cl_2. The Kc value of the NOCl decomposition equilibrium at that temperature according to the reaction $2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$ is ....
   a. 2.2×10^{-4}
   b. 9.0×10^{-4}
   c. 2.25×10^{-3}
   d. 18.0×10^{-3}
   e. 4.5×10^{-4}

$$2NOCl \longleftrightarrow 2NO + Cl_2$$

Using the M-R-Stb, in equilibrium, you will get these values for the mole
$NOCl = 0.888$ $NO = 0,112$ $Cl_2 = 0,056$
Making this into M..

$$M = \frac{n}{V}$$

$M_{NOCl} = \frac{0.056}{2}$
$M_{NO} = \frac{0.112}{2}$
$M_{Cl_2} = \frac{0.888}{2}$
Put this on the K_c formula...

$$Kc = \frac{[NO]^2[Cl_2]}{[NOCl]^2}$$

$$Kc = \frac{\left[\frac{0.112}{2}\right]^2\left[\frac{0.056}{2}\right]}{\left[\frac{0.888}{2}\right]^2}$$

$$Kc \approx 4.5 * 10^{-4}$$

Hence the answer is **E**.

7. In the equilibrium reaction:
   $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -kJ$
   Which of the following will cause more ammonia to be formed?
   a. Decrease in pressure
   b. Decrease in temperature
   c. Addition of ammonia gas
   d. Addition of a catalyst
   e. Increase in volume

According to Le Châtelier's Principle, a decrease in temperature will make the system more exothermic, it will release more gas, since this is a exothermic reaction (you can see by the - sign), it releases heat to make the ammonia, hence decreasing the temperature will make more of the ammonia, a catalyst will only speed up the reaction but won't do anything else, and everything else is also already mentioned in Le Châtelier's Principle. Hence the answer is **B**.

8. At temperature T K, the same values of Kc and Kp are shown in the equilibrium reaction … .
   a. $2SO2 + O2 \rightleftharpoons 2SO3$
   b. $H2 + CO2 \rightleftharpoons H2O + CO$
   c. $CO + 3H2 \rightleftharpoons CH4 + H2O$
   d. $N2O4 \rightleftharpoons 2NO2$
   e. $N2 + 3H2 \rightleftharpoons 2NH3$

Equilibrium means that both sides must have the same amount of coefficient on both sides, hence the answer is **B**.

9. The gas NOBr decomposes according to the following equilibrium:
   2NOBr(g) ⇌ 2NO(g) + Br2(g)
   A closed container of fixed volume is filled with NOBr gas until its pressure reaches 80 atm. If, after equilibrium is reached, 50% of the NOBr gas decomposes. The equilibrium constant, Kp, for the reaction above is ....
   a. 10
   b. 20
   c. 30
   d. 40
   e. 60

In equilibrium, half of the initial 80 atm of NOBr gas decomposed, making it 40 atm. Assuming using the ratio, since it's 2NOBr, it would be the same for 2NO, and half for Br_2. Using the formula assuming it's like that...

$$K_p = \frac{[NO]^2[Br_2]}{[NOBr]^2}$$

$$K_p = \frac{40^2 * 20}{40^2} = 20$$

Hence the answer is **B**.

10. For the hypothetical equilibrium reaction
    A(s) + B(l) ⇌ C(g) + D(g) ΔH° = 84.3 kJ
    If both reactants are placed in a tightly sealed container and allowed to reach equilibrium, the number of moles of C can be decreased by ....
    a. Adding more A to the system
    b. Decreasing the volume of the reaction vessel
    c. Reducing the amount of D in the system
    d. Increasing the reaction temperature
    e. Increasing the amount of catalyst

Since A is a solid, and B is a liquid, we will ignore that according to the Le Châtelier's Principle. According to the Le Châtelier's Principle, decreasing the volume will make it go to the smaller coefficient, 0, so the number of moles of C will decrease, but to be honest it's better to use the process of elimination on this one, as all the other options are just plain wrong. Hence the answer, is **B**.

3.2 Essay

1. In a 5 litre space, and a pressure of 0.4 atm, there is an equilibrium reaction.

$$2NO + O_2 \longleftrightarrow N_2O_4$$

If 0.2 moles of NO gas are mixed with 0.2 moles of O_2 gas, and there is 0.05 moles of N_2O_4 in equilibrium, find α of NO, the K_c, and K_p.

a. Using the M-R-Stb table, we can find the degree of dissociation of NO, which is

$$\alpha = \frac{R}{M}$$
$$\alpha = \frac{0.1}{0.2} = 0.5$$

b. Next off, find the K_c, we first have to find the individual M of each compound.
$M_{NO} = \frac{0.1}{5}$
$M_{O_2} = \frac{0.15}{5}$
$M_{N_2O_4} = \frac{0.05}{5}$

Put this in the formula..

$$K_c = \frac{[N_2O_4]}{[NO]^2[O_2]}$$
$$K_c = \frac{\frac{0.05}{5}}{[\frac{0.1}{5}]^2[\frac{0.15}{5}]}$$
$$K_c = 833.333M^{-2}$$

c. Find K_p, first we have to find the individual partial pressure.

$P_{NO} = \frac{0.1}{0.3} * 0.4 = 0.1333$
$P_{O_2} = \frac{0.15}{5} * 0.4 = 0.2$
$P_{N_2O_4} = \frac{0.05}{0.3} * 0.4 = 0.0666$
Putting this on the formula...

$$K_p = \frac{[P_{N_2O_4}]}{[P_{NO}]^2[P_{O_2}]}$$
$$K_p = \frac{0.066666}{0.133^2 * 0.2}$$
$$K_p \approx 18.8252$$

Number 2 is asking if in a 2 litre space, CCl_4 gas will decompose according to this reaction

$$CCl_4(g) \longleftrightarrow C(s) + 2Cl_2(g)$$

In equilibrium, 4 moles of C is obtained.

a. Find the α of CCl_4, in this case using the M-R-Stb table, if you did it right we will get 4 as the R.

$$\alpha = \frac{R}{M}$$
$$\alpha = \frac{4}{5} = 0.8$$

b. Find K_c. Since C is a solid, we will ignore that one.

$M_{CCl_4} = \frac{1}{2} = 0.5M$

$M_{Cl_2} = \frac{8}{2} = 4M$
Using the formula, it is

$$K_c = \frac{[Cl]^2}{[CCl_4]}$$
$$K_c = \frac{16}{0.5}$$
$$K_c = 32$$

c. Find the K_p if the T = 500 K, reminding to find Δn you have to substract the total of right coefficient minus the left sum coefficient.

$$K_p = K_c(RT)^{\Delta n}$$
$$K_p = 32(0.082 * 500)$$
$$K_p = 1312$$

3. If the $K_c = 0.04$ for,

$$2X + 2Y \longleftrightarrow 4Z$$

a. Find the K_c of

$$2Z \longleftrightarrow X + Y$$

$$\frac{1}{0.04} = \frac{[Y]^2[X]^2}{[Z]^4}$$

Using indices, root both sides

$$(\frac{1}{0.04})^{\frac{1}{2}} = (\frac{[Y]^2[X]^2}{[Z]^4})^{\frac{1}{2}}$$
$$\frac{1}{0.2} = \frac{[X][Y]}{[Z]^2}$$
$$K_c = \frac{1}{0.2} = 5$$

b. Find the K_c for

$$X + Y \longleftrightarrow 2Z$$

$$\sqrt{0.04} = \frac{Z^2}{Y * X}$$
$$K_c = 0.2$$

With this I am done with the Equilibrium guidebook and WS Review PG + Essay, it is 23:20, I've been doing this since 17:30, rip my sleep schedule ig. GLHF for tmr's test! GBU!