Unit 2: Chemical Equilibrium

What this unit is about?

When two reacting species are reacted with each other, *something* will happen, but what happens when the reaction stops? Is it because the reaction ran out of reactants or has the reaction reached its equilibrium? What controls the amount of products?

Dynamic Equilibrium

Many chemical reactions are reversible, meaning the reactants can react to make products or the products can react to make reactants.

Reversible reactions are denoted by $\rightleftharpoons$ such as

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$$

This reaction is stating that when Hydrogen gas and Iodine gas is reacted with each other, it will produce hydrogen iodide **BUT** hydrogen iodide gas can be decomposed back into the reactants.

This is called Dynamic Equilibrium.

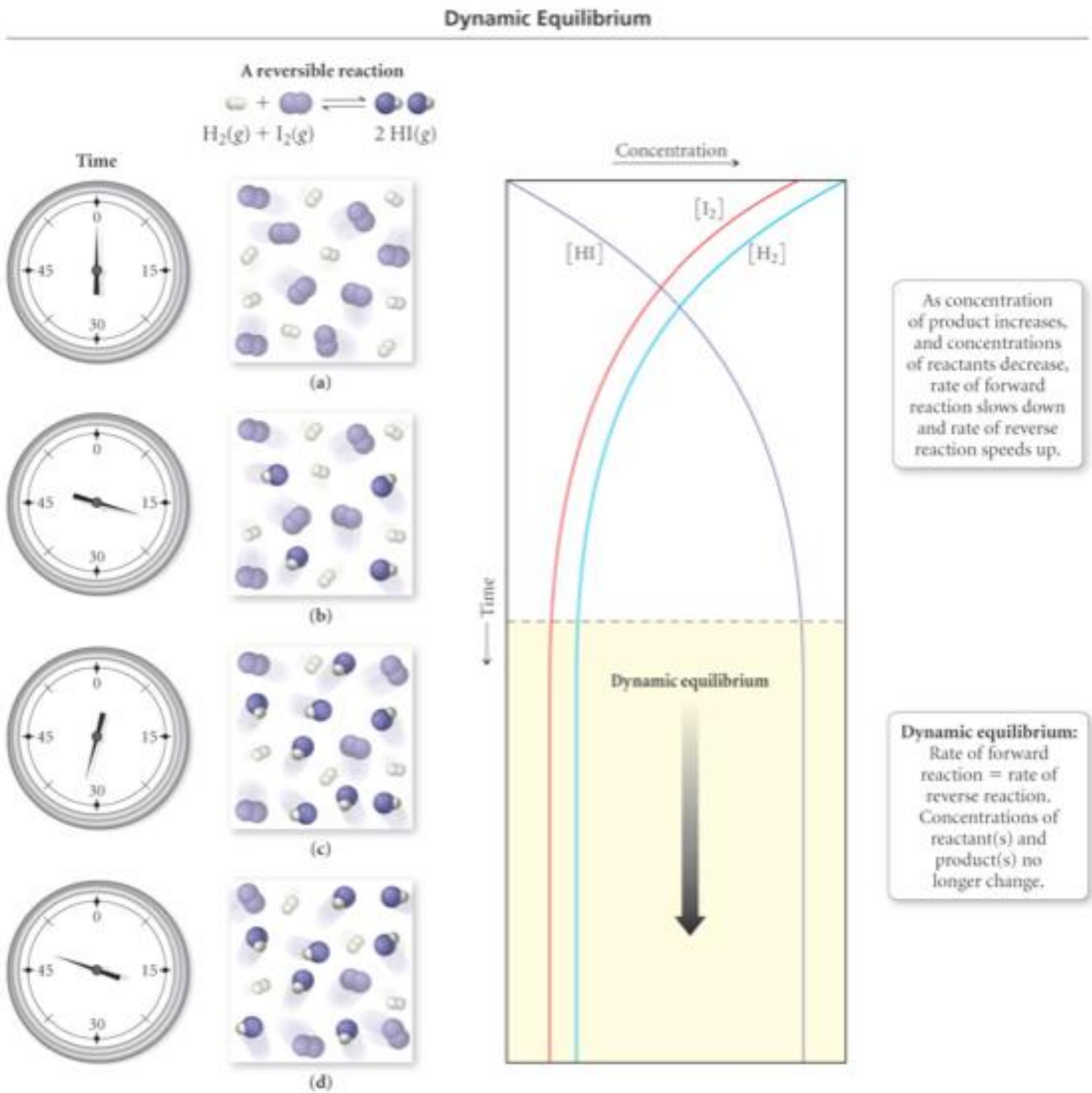
What is dynamic equilibrium?

The condition where the *rate* of forward and backward reactions are identical.

Notice that it is the *RATE* and *NOT* the *CONCENTRATION*! At equilibrium, concentrations of reactants and products do not have to be identical!

It doesn't matter if we start with all products or all reactants, eventually the same *rate* of each would result after reaction at equilibrium.

Dynamic equilibrium can only be achieved in a <u>closed system.</u>



In the diagram above, at equilibrium, the rate of products and reactants are the same even though if the concentrations are different.

The equilibrium Constant, K

The equilibrium constant, K, is used to to see if the reaction favors the amount of products produced or the reactants.

K, is calculated by comparing the relative amounts of products to reactants.

General reaction equation:

$$aA + bB \rightleftharpoons cC + dD$$

The coefficients, a,b,c,d are used as exponents in the K (Law of Mass Action Expression) while A, B, C and D are concentrations of the reacting compounds in mol/L.

$$K = \frac{\overbrace{[C]^c \cdot [D]^d}^{\text{Products}}}{\underbrace{[A]^a \cdot [B]^b}_{\text{Reactants}}}$$

Key Mathematical Concepts that can be learned from this expression:

1. As you can see, K, is merely a ratio between the amount of products to the amount of reactants.
2. When K is a large number (ex. 1.5×10^{19}), that means there are A LOT more products produced than reactants at equilibrium. Products are favored

3. When K is a small number (ex. 2.5 x 10^{-4}), that means there are MORE reactants than products at equilibrium. Reactants are favored.
4. K cannot be negative

“Favoured” means that side of the equation has a higher concentration(moles) than the other.

<u>Writing Law of Mass Action (K constant expression)</u>

- Liquids and solids are not included in the Law of Mass Action expression, only aqueous (aq) and gas (g) are included!
- Coefficients become exponents

Example:

$$CO_{(g)} + 3H_{2(g)} \rightleftharpoons CH_{4(g)} + H_2O_{(g)}$$

$$K = \frac{[CH_4][H_2O]}{[CO][H_2]^3}$$

(Notice the 3 became an exponent)

Example 2:

$$2\ TiCl_{3(s)} + 2\ HCl_{(g)} \rightleftharpoons 2\ TiCl_{4(s)} + H_{2(g)}$$

Example 3:

$$P_{4(s)} + 6\ Cl_{2(g)} \rightleftharpoons 4\ PCl_{3(s)}$$

Example 4:

$$SO_{2(g)} + NO_{2(g)} \rightleftharpoons SO_{3(g)} + NO_{(g)}$$

The value of K is AT equilibrium and can determine if the reaction is AT equilibrium, before, or past equilibrium.

Example 5:

At a given temperature, the reaction $CO_{(g)} + H_2O_{(g)} \rightleftharpoons H_{2(g)} + CO_{2(g)}$ produces the following concentrations at equilibrium: . Calculate the equilibrium constant, K, if $[H_2] = 0.320$, $[CO_2] = 0.420$, $[CO] = 0.200$ and $[H_2O] = 0.500$.

Example 6:

Comment on the favorability of product formation in each of the reactions.

a. $H_2(g) + F_2(g) \leftrightarrow 2\ HF(g)$ $K_C = 1.0 \times 10^{13}$

b. $SO_2(g) + NO_2(g) \leftrightarrow NO(g) + SO_3(g)$ $K_C = 1.0 \times 10^{2}$

c. $2\ H_2O(g) \leftrightarrow 2\ H_2(g) + O_2(g)$ $K_C = 6.0 \times 10^{-28}$

Example 7: Chemists have determined the equilibrium constants for several reactions. In which of these reactions are the products favoured over the reactants?

a. $K_C = 1.0\times10^{2}$
b. $K_C = 3.5$
c. $K_C = 0.003$
d. $K_C = 6.0\times10^{-4}$

Calculating K from equilibrium concentrations

Recall that:

1. Equilibrium constant is calculated from concentrations of reactants and products AT equilibrium not initial concentrations (before reaction).
2. In order to calculate the equilibrium constant with initial concentrations, stoichiometry must be used.
3. The "I-C-E" table is used to calculate K when initial concentrations are provided. I – Initial, C – Change, and E – equilibrium

Example:

$$2CH_{4(g)} \rightleftharpoons C_2H_{2(g)} + 3H_{2(g)}$$

If we start with only $CH_{4(g)}$ with an initial concentration of 0.300M and at equilibrium the [$C_2H_{2(g)}$ is 0.045M, calculate the equilibrium constant.
(Note: the question has INITIAL and EQUILIBRIUM concentrations! Recognize this question type)

1. Setup ICE table & place the proper concentration at the appropriate box.

2. Make sure you have the correct Signs (+/-). Use + signs for products since you are gaining, and – signs for reactants since you are using them up if you are provided with the equilibrium concentration.

3. Use stoichiometry to calculate the amount change for other reacting species.

4. Calculate all concentrations at equilibrium and solve for K.

5. Reason if your K indicates product or reactant favouring.

Example 2: (Question gives you the K, and initial values).

$$N_2(g) + O_2(g) \quad 2\,NO(g)$$

The equilibrium constant is 6.76.
If 6.0 moles of nitrogen and oxygen gases are placed in a 1.0 L container, what are the concentrations of all reactants and products at equilibrium?

1. Setup ICE table.
2. In this question, since you are not provided with ANY concentrations AT equilibrium, you will need to use Algebra to substitute into the "Change" row of your ICE table.
3. Make sure your "x" is consistent with the correct +/- sign and also the stoichiometry ratio.
4. Solve for x once you have established your equilibrium concentrations and substitute your value of x back into your ICE table.

Calculating the Reaction Quotient, Q

Up to this point, we looked at if a reaction at equilibrium is favoring the products or reactants AT equilibrium. But can we predict the direction of a reaction at ***any*** time?

At any one time of a reaction, is the reaction producing more products (moving forward) or more reactants (moving backwards)?

This is where the reaction quotient, Q will come into play.

The reaction quotient, Q, is the *instantaneous ratio of products to reactants.* We then compare the K and the Q to see the direction of the reaction.

Setting up Reaction Quotient, Q

The setup is exactly identical to K.

$$aA + bB \rightleftharpoons cC + dD$$

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

But what is important is knowing what Q means if it is less than or greater than K.

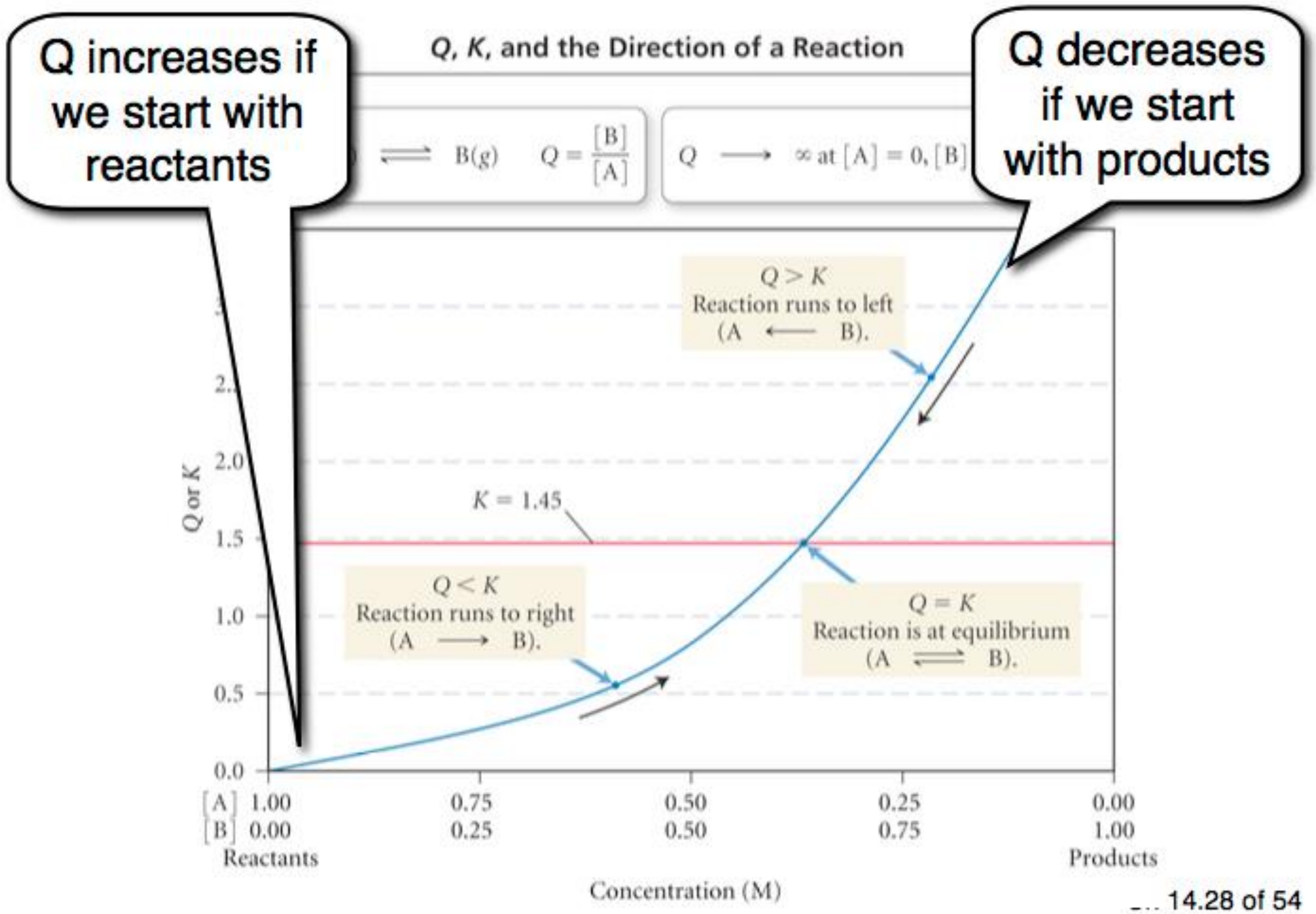
3 scenarios to Q to K.

If Q = K. If Q is equal to K, that means the rates of reactants and products generated are identical between Q and K. This means the reaction is at equilibrium.

If Q < K. If Q is less than K, that means that at that particular moment of time, the reaction would have more reactants than products since you would have more of

[A] and [B] than of [C] and [D] when compared to K. In other words, in order for Q to be smaller than K, the denominator of Q must be greater, meaning more reactants. The reaction is not at equilibrium and is moving forward in converting the reactants into products.

If Q > K. If Q is greater than K, that means that at that particular moment of time, the reaction would have more products than reactants. The reaction is not at equilibrium and is moving backwards in converting the reactants into reactants.



Notice from the graph above, the horizontal line represents K, and the curve is the progression of the reaction through time.

At the beginning of the reaction, the concentration of reactants is large (Q is smaller than K – remember larger denominator) and as the reaction progresses, Q increases, causing the reactants to decrease. When Q becomes greater than K, it means there are more products produced and must be reverted by to the reactants favoring a backwards reaction.

<u>Reaction Quotient Questions</u>

1. For the reaction below, K = 25.0:

   $A(s) + 2\,B(g) \leftrightarrow 2\,C(g)$

   If [A] = 12.0 mol/L [B] = 2.0 mol/L [C] = 30.0 mol/L

   a) Is the system is at equilibrium.
   b) Which reaction is faster (favoured), forward or reverse?
   c) Which concentrations are increasing or decreasing?

2. There exists an equilibrium if 5.0 moles of CO_2, 5.0 moles of CO and 0.20 moles of O_2 are in a 2.0 L container at 562°C. Find K_C for the reaction

   $2\,CO(g) + O_2(g) \leftrightarrow 2\,CO_2(g)$

   Would the system be at equilibrium if $[CO_2]$ = 15.8 mol/L, [CO] = 10.0 mol/L and $[O_2]$ = 0.25 mol/L? If not, which reaction is favoured?

3. For the reaction below, K = 16.0:

$$2\ SO_2(g) + O_2(g) \leftrightarrow 2\ SO_3(g)$$

Initially, $[SO_2]$ = 5.0 mol/L, $[O_2]$ = 10.0 mol/L and $[SO_3]$ = 0. After two hours $[O_2]$ = 7.9 mol/L. Is the system at equilibrium? If not, which substances are increasing and which are decreasing?

4. The reaction

$$4\ NH_3(g) + 5\ O_2(g) \leftrightarrow 4\ NO_2(g) + 6\ H_2O(g)$$

is at equilibrium when $[H_2O]$ = 0.100 mol/L, $[O_2]$ = 2.00 mol/L, [NO] = 0.200 mol/L and $[NH_3]$ = 0.500 mol/L. If 0.75 moles of H_2O, 12.0 moles of NO, 30.0 moles of O_2 and 0.30 moles of NH_3 are in a 3.0 L container at the same temperature, is equilibrium achieved? If not, which reaction is favoured?

Review Questions

1. Why are solids and liquids not included in the equilibrium constant expression?

2. What does the value of K mean in terms of the amount of reactants and products?

3. What is the correct equilibrium constant expression for the following reaction: $2\ \mathrm{SO}_{2(g)} + \mathrm{O}_{2(g)} \rightleftharpoons 2\ \mathrm{SO}_{3(g)}$?
    a. $K = \frac{[\mathrm{SO_2}]^2[\mathrm{O_2}]}{[\mathrm{SO_3}]^2}$
    b. $K = \frac{[\mathrm{SO_3}]^2}{[\mathrm{SO_2}]^2[\mathrm{O_2}]}$
    c. $K = \frac{[2\ \mathrm{SO_2}][\mathrm{O_2}]}{[2\ \mathrm{SO_3}]}$
    d. $K = \frac{[2\ \mathrm{SO_3}]}{[2\ \mathrm{SO_2}][\mathrm{O_2}]}$
4. What is the correct equilibrium constant expression for the following reaction: $\mathrm{Cu(OH)}_{2(s)} \rightleftharpoons \mathrm{Cu}^{2+}_{(aq)} + 2\ \mathrm{OH}^{-}_{(aq)}$?
    a. $K = \frac{[\mathrm{Cu}_{2+}][\mathrm{OH}_{-}]^2}{[\mathrm{Cu(OH)}^2]}$
    b. $K = \frac{[\mathrm{Cu(OH)}^2]}{[\mathrm{Cu}^{2+}][\mathrm{OH}^-]^2}$
    c. $K = \frac{1}{[\mathrm{Cu}^{2+}][\mathrm{OH}^-]^2}$
    d. $K = [\mathrm{Cu}^2+][\mathrm{OH}^-]^2$

5. Consider the following equilibrium system: $2\ NO_{(g)} + Cl_{2(g)} \rightleftharpoons 2\ NOCl_{(g)}$. At a certain temperature, the equilibrium concentrations are as follows: $[NO] = 0.184\ mol/L,\ [Cl_2] = 0.165\ mol/L$, and $[NOCl] = 0.060\ mol/L$. What is the equilibrium constant for this reaction?
   a. 0.506
   b. 0.648
   c. 1.55
   d. 1.97

6. For the reaction $2\ MgCl_{2(s)} +\ O_{2(g)} \rightleftharpoons 2\ MgO_{(s)} + 2\ Cl_{2(g)}$, the equilibrium constant was found to be 3.86 at a certain temperature. If $0.560\ mol\ O_{2(g)}$ is placed in a $1.00\ L$ container, what is the concentration of $Cl_{2(g)}$ at equilibrium?
   a. $1.47\ mol/L$
   b. $2.16\ mol/L$
   c. $2.88\ mol/L$
   d. not enough information is available

7. Write the equilibrium constant expressions for each of the following equations:
   a.
   b. $2\ CaSO_{4(s)} \rightleftharpoons 2\ CaO_{(s)} + 2\ SO_{2(g)} + O_{2(g)}$
   c. $2\ Fe^{3+}_{(aq)} + 3\ S^{2-}_{(aq)} \rightleftharpoons Fe_2S_{3(s)}$
   d. $Hg_{(l)} + H_2S_{(g)} \rightleftharpoons HgS_{(s)} + H_{2(g)}$

8. A mixture at equilibrium at 827°C contains 0.552 moles of CO_2, 0.552 moles H_2, 0.448 moles CO, and 0.448 moles of H_2O in a 1.00 L container. What is the value of the equilibrium constant, Keq?

$$CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g)$$

9. The equilibrium constant for the reaction

$$4\ H_2(g) + CS_2(g) \rightleftharpoons CH_4(g) + 2\ H_2S(g)$$

at 755°C is 0.256. What is the equilibrium concentration of H_2S if at equilibrium $[CH_4]$ = 0.00108 mol/L, $[H_2]$ = 0.316 mol/L, $[CS_2]$ = 0.0898 mol/L?

10. Find the value of Keq for the equilibrium system

$$ZnO(s) + CO(g) \rightleftharpoons Zn(s) + CO_2(g)$$

if at equilibrium there are 3.0 moles of CO, 4.0 moles of Zn and 4.0 moles of CO_2 in a 500.0 mL container.

USE ICE TABLES for the following

11. The decomposition of hydrogen iodide to hydrogen and iodine occurs by the reaction

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

Hydrogen iodide is placed in a container at 450°C an equilibrium mixture contains 0.50 moles of hydrogen iodide. The equilibrium constant is 0.020 for the reaction. How many moles of iodine and hydrogen iodide are present in the equilibrium mixture?

12. $H_2(g) + Cl_2(g) \rightleftharpoons 2\ HCl(g)$

A student places 2.00 mol H_2 and 2.00 mol Cl_2 into a 0.500 L container and the reaction is allowed to go to equilibrium at 516°C. If K_C is 76.0, what are the equilibrium concentrations of H_2, Cl_2 and HCl?

13. If Keq = 78.0 for the reaction $A(s) + 2\ B(g) \rightleftharpoons 2\ C(g)$ and initially there are 5.00 moles of A and 4.84 moles of B in a 2.00 L container, how many moles of B are left at equilibrium?

Le Châtelier's Principle

In the late 1800s, a chemist by the name of Henry-Louis Le Châtelier was studying stresses that were applied to chemical equilibria. He formulated a principle, **Le Châtelier's Principle**, which states that when a stress is applied to a system at equilibrium, the equilibrium will **<u>shift in a direction to partially counteract</u>** the stress and once again reach equilibrium.

For instance, if a stress is applied by increasing the concentration of a reactant, the equilibrium position *will shift toward the right* and remove that stress by using up some of the reactants. The reverse is also true. If a stress is applied by lowering a reactant concentration, the equilibrium position will shift toward the left, this time producing more reactants to partially counteracting that stress. The same reasoning can be applied when some of the products is increased or decreased.

Le Châtelier's principle does not provide an explanation of what happens on the molecular level to cause the equilibrium shift. Instead, it is simply a quick way to determine which way the reaction will run in response to a stress applied to the system at equilibrium.

There are 4 conditions:

b) Concentration Change
c) The addition of a Catalyst
d) Pressure change
e) Temperature change

1. Effect of Concentration Changes

Let's use Le Châtelier's principle to explain the effect of concentration changes on an equilibrium system.

$$N_2O_{4(g)} \rightleftharpoons 2NO_{2(g)}$$

If there is a sudden increase of NO_2, then the is an added stress to the products side. Le Chatelier's principle states that reactions will alleviate the stress by shifting the equilibrium, which is to change the products to reactants.

Result: An increase of $[N_2O_{4(g)}]$

The reaction can be viewed as a seesaw

$$N_2O_{4(g)} \rightleftharpoons 2NO_{2(g)}$$

If there is an **increase** of the **reactants**, the **stress** is on the **reactants**, so the reaction will then **shift** to the right by producing more products. Tipped on the left side so the product side is to increase.

Likewise, if there is an **increase** of **products**, the **stress** is on the **products**, so the reaction will the **shift** to the left by producing **more** reactants.

If there is an DECREASE of reactants, then the **stress** is on the **product** side, so the equilibrium shifts back toward the reactants to alleviate the **stress** on the **product** side. Elevated on the left side so the seesaw shifts to reactant side and increases it.

If there is an DECREASE of products, then the **stress** is on the reactant side, so the equilibrium shifts towards the products.

Example:

For the reaction $\mathrm{SiCl}_{4(g)} + \mathrm{O}_{2(g)} \rightleftharpoons \mathrm{SiO}_{2(s)} + 2\ \mathrm{Cl}_{2(g)}$, what would be the effect on the equilibrium system if:

1. $[\mathrm{SiCl}_4]$ increases
2. $[\mathrm{O}_2]$ increases
3. $[\mathrm{Cl}_2]$ increases

Solution:

1. The equilibrium would shift to the right. $[\mathrm{Cl}_2]$ would increase, more SiO_2 would be produced (but that does not increase its concentration since its a solid), and $[\mathrm{O}_2]$ would decrease.
2. The equilibrium would shift to the right. $[\mathrm{SiCl}_4]$ would decrease, more SiO_2 would be produced (but again no change in concentration), and $[\mathrm{Cl}_2]$ would increase.
3. The equilibrium would shift left. $[\mathrm{SiCl}_4]$ and $[\mathrm{O}_2]$ would increase, and SiO_2 would be used up but not change its concentration.

Example:

For the reaction $\mathrm{PCl}_{3(g)} + \mathrm{Cl}_{2(g)} \rightleftharpoons \mathrm{PCl}_{5(g)}$, which way will the equilibrium shift if:

1. $[\mathrm{PCl}_3]$ decreases
2. $[\mathrm{Cl}_2]$ decreases
3. $[\mathrm{PCl}_5]$ decreases

Solution:

1. left
2. left
3. right

Example:

Here's a reaction at equilibrium. Note the phases of each reactant and product.

$$A_{(aq)} + B_{(s)} \rightleftharpoons C_{(aq)} + D_{(aq)}$$

1. Which way will the equilibrium shift if you add some A to the system without changing anything else?
2. After A has been added and a new equilibrium is reached, how will the new concentration of D compare to the original concentration of D?
3. After A has been added and a new equilibrium has been established, how will the new concentration of A compare to the original concentration of A?
4. After A has been added and a new equilibrium has been established, how will the new concentration of B compare to the original concentration of B?
5. Which way will the equilibrium shift if you add some C to the system without changing anything else?
6. After C has been added and a new equilibrium has been established, how will the new concentration of D compare its original concentration?
7. After C has been added and a new equilibrium has been established, how will the new concentration of A compare its original concentration?
8. Which way will the equilibrium shift if you add some B to the system without changing anything else?

Solution:

1. The equilibrium will shift toward the products.
2. The new concentration of D will be higher than the original.
3. The new concentration of A will be higher than the original, but lower than the concentration right after A was added.
4. Since B is a solid, its concentration will be the same as the original. There will be less of it since some was used in the equilibrium shift, but the concentration will be the same.
5. The equilibrium will shift toward the reactants.
6. The new concentration of D will be lower than the original.
7. The new concentration of A will be higher than the original.
8. Since B is a solid, adding B will not change its concentration and therefore has no effect on the equilibrium. It is possible that adding some B will increase the surface area of B and therefore increase the forward reaction rate, but it will also increase the reverse reaction rate by approximately the same amount, hence no shift in equilibrium.

2. Addition of a catalyst

Adding a catalyst to a system at equilibrium will NOT affect the equilibrium position. However, if a catalyst is added to a system which is not at equilibrium, the system will reach equilibrium much quicker since forward and reverse reaction rates are increased.

3. Pressure Change / Volume change

The pressure and volume change in a system have an inversely proportional effect since the increase in volume will decrease the pressure while increase in pressure will decrease in volume. If you can understand 1, then the other is the opposite.

In Haber's Ammonia producing process

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

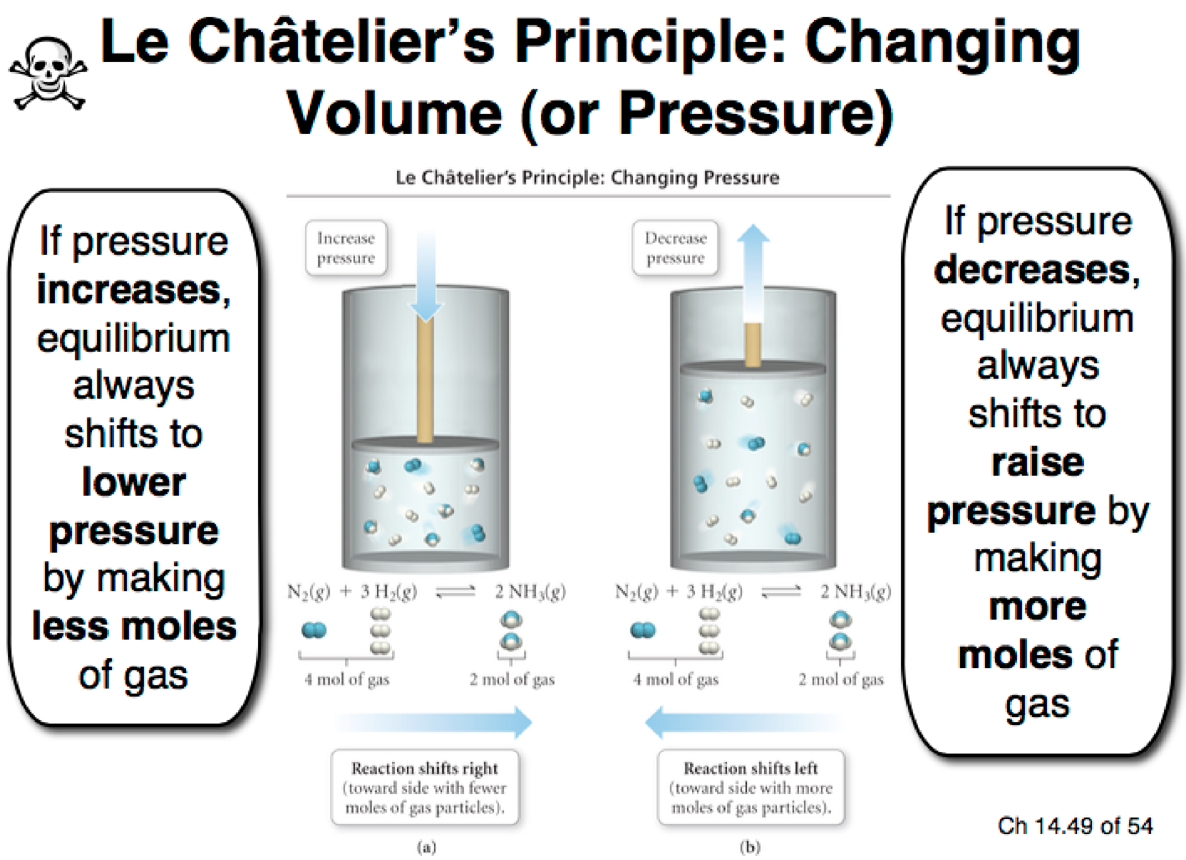
Pressure change effects:

If the pressure holding the nitrogen and hydrogen is increased, the **number** of collisions between the reactants will be more **frequent** since there are more of them. The reaction will have to **decrease** the **stress** on the reactant side due to the **increased** in **pressure** and **produce more products** to minimize the collisions.

If the pressure is decreased however, we need to shift the reaction to the direction where there will be ***more*** collisions to compensate the decreased in pressure.

Summary:

To see which direction a reaction will shift, the number of moles (coefficients) must be taken into account. Pressure = collisions, the more pressure, the more collisions, the more stress, thus the shift must be towards the opposite to alleviate the stress.



3a. The effects of Volume change

In order to see the effect of a volume change, understanding the effects of pressure change will make it easier.

Going back to Haber's Ammonia producing process

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

If the volume is decreased, where is the stress and which direction will the reaction shift?

If the volume is decreased, then the pressure is increased. That means the stress is where there are more collisions (reactant side with more moles). In order to alleviate the stress, the reaction shifts to the right to make less moles.

However, if the volume is increased, then the pressure is decreased. The equilibrium will now have to shift to the side where there are MORE moles, in this case, reactants.

If both sides contain the same number of reactant and product moles, increasing or decreasing volume will have no effect.

4. **The effects Temperature change**

Some reactions will be exothermic (produces heat)

$$A + B \rightleftharpoons C + D + \text{Heat}$$

While some reactions are endothermic (takes in heat)

$$A + B + \text{heat} \rightleftharpoons C + D$$

A simple way of looking at it is that the effects of temperature change is similar to concentration change that an increase in temperature will add stress to the temperature side. The result will be shifting the reaction to the opposite side.

Example:

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)} + \text{heat}$$

If reaction is heated up, the reaction will shift to make more *reactants*, to *absorb* the *excess* heat.

If the reaction is cooled down, then the reaction will shift to make *products*, to *release more* heat.

Example:

$$N_2O_{4(g)} \text{ (Colorless)} \rightleftharpoons 2NO_{2(g)} \text{ (Brown)}$$

Is the reaction an endothermic or exothermic reaction if the reaction turns brown when heated?

It must be an endothermic reaction

$$N_2O_{4(g)} \text{ (Colorless)} + \text{heat} \rightleftharpoons 2NO_{2(g)} \text{ (Brown)}$$

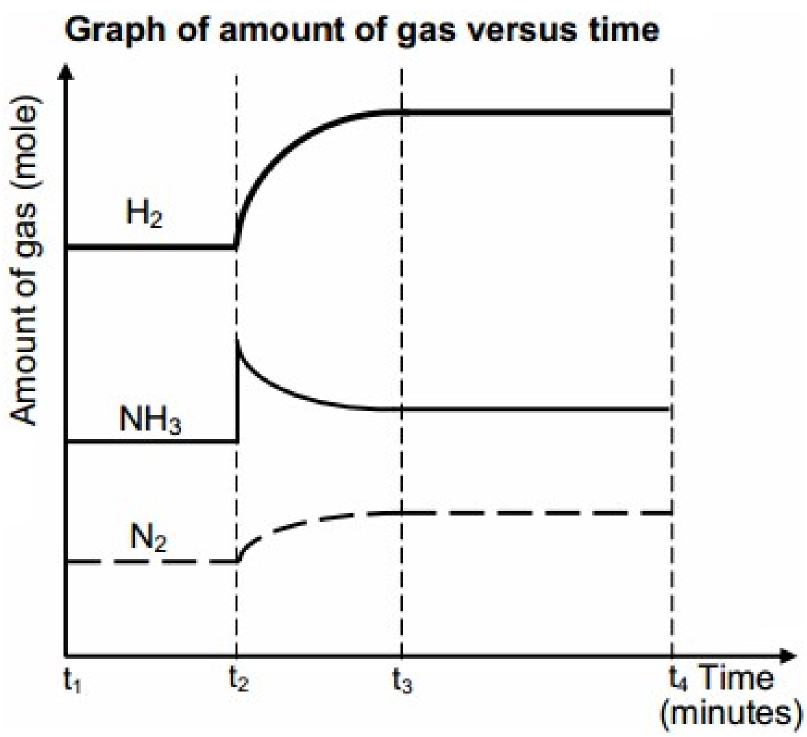
Since if we were to heat it up, the stress will be on the reactant side, and to alleviate the stress, it has to shift to the right to absorb the heat.

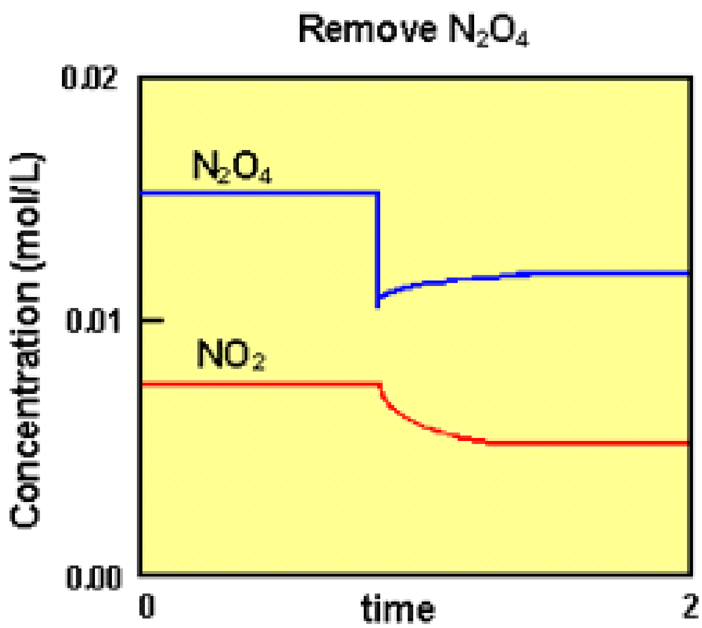
Interpreting Graphs during reaction stress

- These graphs will show what is happening when a stress is applied
- The concentration vs. time graphs show what is happening to each reacting species when the stress is applied
- On any graph an equilibrium is shown by a flat horizontal line
- After the stress has been applied, the line curves up or down depending on the situation, and then it levels off at a new equilibrium level
- The stress is indicated by a sharp peak (up for increases and down for decreases)
- The rest of the lines gently curve up or down depending on Le Chatelier
- Temperature changes cause all reacting species to have gentle curves in the appropriate directions

Example: $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$ + heat

At t2, there is a sharp peak at $NH_{3(g)}$, that means there was a sudden increase of ammonia gas. As a result, the stress is on the product side, and the reaction shifts to the left to achieve equilibrium, which is seen at t3.





$2\ NO_2(g) \leftrightarrow N_2O_4(g)$

Sudden drop of N_2O_4 concentration shifts the reaction to the right by creating MORE of the N_2O_4 compound until equilibrium.

1. For the reaction

   $PCl_3(g) + Cl_2(g) \leftrightarrow PCl_5(g)$ $\Delta H = -92.5$ kJ

   predict the effect on the position of the equilibrium that results from

   a) increasing the total pressure by decreasing volume.
   b) injecting more Cl_2 gas without changing the volume.
   c) increasing the temperature.
   d) increasing the volume of the container.
   e) adding a catalyst.

2. For the reaction

   $CH_4(g) + H_2O(g) + 49.3\text{ kJ} \leftrightarrow CO(g) + 3\,H_2(g)$

   predict the effect on the position of the equilibrium that results from

   a) increasing temperature.
   b) decreasing temperature.
   c) decreasing the pressure.
   d) decreasing the volume of the container.
   e) adding a solid drying agent such as $CaCl_2$ which reacts with $H_2O(g)$.

3. For the reaction

   $9.4\text{ kJ} + 2\,HI(g) \leftrightarrow H_2(g) + I_2(g)$

   a) What is the effect on [HI] if a small amount of H_2 is added?
   b) What is the effect on [HI] if the pressure of the system is increased?
   c) What is the effect on [HI] if the temperature is increased?
   d) What is the effect on [HI] if a catalyst is added?

4. For the reaction
   $CO(g) + 2\,H_2(g) \leftrightarrow CH_3OH(g)$ + energy
   predict the effect of the following changes on the equilibrium concentration of $CH_3OH(g)$

   a) a decrease in temperature.
   b) an increase in pressure.
   c) addition of $H_2(g)$.
   d) addition of a catalyst.

5. In the equilibrium reaction

   $2\,NO(g) + O_2(g) \leftrightarrow 2\,NO_2(g) + 114.6\text{ kJ}$

   What will be the change in the equilibrium $[NO_2]$ under each of the following conditions?

   a) O_2 is added.
   b) NO is removed.
   c) energy is added.

6. For the following reaction ΔH = +58.9 kJ

   $N_2O_4(g) \leftrightarrow 2\ NO_2(g)$

   how will the equilibrium $[NO_2]$ be affected by the following?

   a) an increase in pressure.
   b) an increase in temperature.
   c) the addition of a catalyst.

Additional Le Chatelier's Questions

1. What is the effect on the equilibrium position if the pressure is increased?
2. What is the effect on the equilibrium position if the pressure is decreased?
3. Use Le Châtelier's Principle to predict what will happen to the following reaction at equilibrium if the pressure is increased: $2\ NH_{3(g)} \rightleftharpoons N_{2(g)} + 3\ H_{2(g)}$. Mark all that apply.
   a. equilibrium position shifts right
   b. equilibrium position shifts left
   c. $[N_2]$ will decrease
   d. $[NH_3]$ will increase
4. Use Le Châtelier's principle to predict what will happen to the following reaction at equilibrium if the pressure is decreased: $2\ NO_{(g)} + 2\ H_{2(g)} \rightleftharpoons N_{2(g)} + 2\ H_2O_{(g)}$. Mark all that apply.
   a. equilibrium position will not shift
   b. equilibrium position shifts left
   c. $[N_2]$ will increase
   d. $[NO]$ will increase
5. Use Le Châtelier's principle to predict what will happen to the following reaction at equilibrium if the volume is decreased: $2\ NCl_{3(g)} \rightleftharpoons N_{2(g)} + 3\ Cl_{2(g)}$. Mark all that apply.
   a. equilibrium position shifts right
   b. equilibrium position shifts left
   c. $[N_2]$ will increase
   d. $[NCl_3]$ will decrease
6. For the reaction $2\ N_2O_{(g)} + O_{2(g)} \rightleftharpoons 4\ NO_{(g)}$, what would be the effect on the equilibrium system if the pressure increases (or the volume decreases)?
7. For the reaction $2\ IBr_{(g)} \rightleftharpoons I_{2(g)} + Br_{2(g)}$, what would be the effect on the equilibrium system if the pressure decreases (or the volume increases)?
8. For the reaction $H_{2(g)} + CO_{2(g)} \rightleftharpoons H_2O_{(g)} + CO_{(g)}$, what would be the effect on the equilibrium system if
   a. the pressure increases?
   b. the volume decreases?
9. For the reaction $3\ NO_{(g)} \rightleftharpoons N_2O_{(g)} + NO_{2(g)}$, what would be the effect on the equilibrium system if
   a. the pressure increases?
   b. the volume decreases