

CHAPTER 10

CHEMICAL EQUILIBRIUM

10.1 EQUILIBRIUM

Equilibrium can be defined as a state when two opposing reactions are equal. Equilibrium is a state of a system where there is no observable change in the properties of the system with regard to time. There are two types of equilibrium namely:

10.1.1 STATIC EQUILIBRIUM

In this type of equilibrium, there is a total standstill of the two opposing reactions. An example is when two kids are balanced on a sea-saw. At the point of balance, the kids and the sea-saw are in a point of total stand still

10.1.2 DYNAMIC EQUILIBRIUM

An example of this type of equilibrium is a boy running up an escalator, and the escalator is descending at the same rate at which the boy is ascending. When these two opposing processes (i.e. the boy ascending and the escalator descending) occur at the same time and rate there will be no observable change in the properties of the system.

Generally, dynamic equilibrium involving physical change is called **physical equilibrium** e.g. when liquid water and solid water (ice) exist together. Also, dynamic equilibrium involving chemical change is called the **chemical equilibrium**.

10.2 CHARACTERISTICS OF DYNAMIC EQUILIBRIUM

10.2.1 Constant macroscopic properties: That means that there is no overall large scale change. As the two opposing processes are occurring at the same time and rate, no net change is observed. The system seems to have come to rest to an external observer.

10.2.2 Continuing microscopic process: This means at equilibrium, molecular scale (very small scale) processes continue, although at a balanced rate but cannot be measured by external observer.

10.2.3 Equilibrium is reached only in a closed system: Considering the production of carbon dioxide (CO₂) by the decomposition of calcium trioxocarbonate (IV) (CaCO₃). At equilibrium, CaCO₃ will be decomposing into CO₂ and calcium (II) oxide (CaO) while the CO₂ and CaO produced will be reacting together again at the same time to form CaCO₃.



When this reaction takes place in an open system. It is obvious that as the CaCO_3 decomposes to form CaO and CO_2 , the CO_2 will be lost into the surrounding and thus there will be no CO_2 to react with the CaO to give the original CaCO_3 . Therefore, equilibrium will not be reached. But in a closed system, either reaction can take place since there is no escape of any substance.

10.2.4 Equilibrium is attained from either direction: Again considering the above reaction, as CaCO_3 decomposes, CaO and CO_2 is formed. Also CaO then combines with CO_2 to produce CaCO_3 . As the CaCO_3 is decomposing, CaO and CO_2 are combining at the same time and rate to form CaCO_3 . Therefore, equilibrium is reached from the initial reaction and the final reaction.

10.2.5 None of the component in an equilibrium system is ever fully consumed: For the forward and the backward reaction to take place at the same time and rate, the reactants and products must be present in the system at all times.

10.3 EQUILIBRIUM IN REVERSIBLE REACTION

A reversible reaction is one that proceeds in both directions under suitable conditions. The two direction under which a reversible reaction can proceed are called the forward and the backward direction.

A reversible reaction is represented by two half arrows (\rightleftharpoons). The \sim arrow represents the forward reaction while --- represents the backward reaction.



10.4 EQUILIBRIUM CONSTANT

Equilibrium constant is the constant relationship that exists between concentrations of reactants and products in a reversible reaction. It is explained by the law of mass action which states that at constant temperatures the rate of a reaction is directly proportional to the active masses of each of the reactants. The concentration of the reactants and products can also be used in place of the active masses.

Considering the reaction;



Here a (moles) of A reacts with b (moles) of B to produce c (moles) of C and d (moles) of D . r_1 and r_2 represents the rate of forward and backward reaction respectively. At a constant temperature, the expression for the law of mass action is given by:

$r_1 \propto [A]^a [B]^b \rightarrow$ (Equation 10.1) for the forward reaction

$r_2 \propto [C]^c [D]^d \rightarrow$ (Equation 10.2) for the backward reaction

$$r_1 = K_1 [A]^a [B]^b \rightarrow \quad (10.3)$$

$$r_2 = K_2 [C]^c [D]^d \rightarrow \quad (10.4)$$

Where K_1 and K_2 = velocity constant for the forward and backward reactions respectively. At equilibrium, the rate of the forward reaction is equal to the rate of the backward reaction. Therefore equation (10.3) is equal to equation (10.4).

$$r_1 = r_2$$

$$\therefore K_1 [A]^a [B]^b = K_2 [C]^c [D]^d$$

Rearranging so as to bring the constants on the same side, we get: $\frac{K_1}{K_2} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$

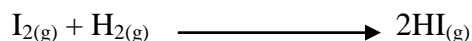
It should be noted that in the mass action expression, the concentration of the products must appear as the numerator while the concentration of the reactant as the denominator.

$$\frac{K_1}{K_2} = K = \text{equilibrium constant}$$

Therefore, at constant temperature, the expression of the equilibrium constant is given by:

$$K = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad \text{Equation 10.5}$$

The expression is also called the **Equilibrium law**. This law tells us that at constant temperature regardless of the concentration of the reactants and products, the equilibrium constant is the ratio of the rate of the forward reaction to the rate of the reversed reaction. This ratio is a constant for that reaction. An example, is the reversible reaction at equilibrium between hydrogen and iodine to produce hydrogen iodide at a temperature of 458°C



We have:

| S/N | Initial conc. in mol/dm ³ | | | Equilibrium conc. in mol/dm ³ | | | K _c |
|-----|--------------------------------------|-------------------------|-------------------------|--|-------------------------|-------------------------|----------------|
| | [H ₂] | [I ₂] | [H _I] | [H ₂] | [I ₂] | [H _I] | |
| 1. | 2.40 x 10 ⁻² | 1.68 x 10 ⁻² | 0 | 0.92 x 10 ⁻² | 0.21 x 10 ⁻² | 3.00 x 10 ⁻² | 46.58 |
| 2. | 2.44 x 10 ⁻² | 1.98 x 10 ⁻² | 0 | 0.77 x 10 ⁻² | 0.31 x 10 ⁻² | 3.34 x 10 ⁻² | 46.13 |
| 3. | 2.46 x 10 ⁻² | 1.76 x 10 ⁻² | 0 | 0.92 x 10 ⁻² | 0.22 x 10 ⁻² | 3.08 x 10 ⁻² | 46.87 |
| 4. | 0 | 0 | 3.04 x 10 ⁻² | 0.35 x 10 ⁻² | 0.35 x 10 ⁻² | 2.35 x 10 ⁻² | 46.40 |

From this results, it is observed that within the limits of experimental errors, and regardless of what the initial concentrations of hydrogen, iodine or hydrogen iodine were, at equilibrium when their concentration are substituted into the mass action expression, the result will always be constant.

The equilibrium constant K_c gives very useful information about the reaction. It tells how far a reaction proceed before equilibrium is reached.

1. If K_c is large, it indicates that the forward reaction goes almost to completion before equilibrium is reached. And at equilibrium, the product will be greater than the reactant.
2. If $K_c = 1$, it indicates that the concentration of the products and reactants are the same when equilibrium is reached.
3. If K_c is lesser than 1, it indicates that the forward reaction does not go far before equilibrium, is reached. And at equilibrium, the reactant will be greater than the product.

The equilibrium constant is not affected by changes in concentration, pressure and catalyst. But a change in temperature affects the equilibrium constant. At higher temperature, K values increases and the product of that particular reaction. Changes in concentration and pressure will not affect the value of the equilibrium constant because the reaction will automatically adjust itself to attain equilibrium again to give the same K_c value. Also, the presence of a catalyst will not affect the value of the equilibrium constant, it will only lower the activation energy so that the rate of the reaction becomes faster, but still the same K value is obtained.

Equilibrium constant does not have a definite unit. The unit depends on the unit on the numerator and the denominator of the equilibrium law expression. K_c will have no unit when the reaction involves equal number of particles „in the products (numerator) and the reactant (denominator) sides, e.g.

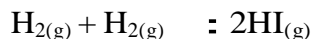
$$\begin{aligned} \text{I}_{2(g)} + \text{H}_{2(g)} &= 2\text{HI}_{(g)} \\ K_c &= \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} \text{ units} = \frac{(\text{mol dm}^{-3})^2}{(\text{mol dm}^{-3})(\text{mol dm}^{-3})} \\ &= \text{no units} \end{aligned}$$

For reactions where the mass-action expression gives unequal particles in the numerator and denominator, K will have a unit e.g.

$$\begin{aligned} K_c &= \frac{[\text{NH}]^2}{[\text{N}_2][\text{H}_2]^3} \text{ units} = \frac{(\text{mol dm}^{-3})^2}{(\text{mol dm}^{-3})(\text{mol dm}^{-3})^3} \\ &= \frac{1}{(\text{mol dm}^{-3})^2} = \text{mol}^{-2} \text{ dm}^6 \end{aligned}$$

10.5 EQUILIBRIUM CONSTANTS INVOLVING GASES

So far we have expressed equilibrium constant with K_c . For reaction involving gases, it is more convenient to express equilibrium constant in terms of the partial pressure of the gases. This is because the partial pressure of gases is proportional to its mole fraction. Therefore, the equilibrium expression of our earlier example;



Becomes:
$$K_p = \frac{[\text{P}_{\text{HI}}]^2}{[\text{P}_{\text{H}_2}] [\text{P}_{\text{I}_2}]}$$

8.6 THE RELATIONSHIP BETWEEN K_c AND K_p

Using the ideal gas equation:

$$Pv = nRT \dots\dots\dots \text{(Equation 10.6)}$$

Rearranging equation (10.6) we get:

$$P = \frac{n}{v} RT \dots\dots\dots \text{(Equation 10.7)}$$

Where:

P = Pressure of the gas in atmosphere

R = Gas Constant

T = Temperature in Kelvin

$\frac{n}{v}$ Number of moles of the gas per volume (is the effective concentration of the gas)

$$\therefore \frac{n}{v} = [\text{gas}]$$

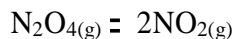
\therefore Equation (10.7) can be written thus:

$$P = [\text{gas}] RT \dots\dots\dots \text{(Equation 10.8)}$$

At constant temperature, the pressure of the gas will be directly proportional to the concentration of the gas. Thus,

$$P \propto [\text{gas}]$$

Applying this in an equilibrium expression



We get

$$K_c = \frac{[\text{NO}_{2(g)}]^2}{[\text{N}_2\text{O}_{4(g)}]} \dots\dots\dots (10.9)$$

$$K_p = \frac{[\text{P}_{\text{NO}_{2(g)}}]^2}{[\text{P}_{\text{N}_2\text{O}_{4(g)}}]} \dots\dots\dots (10.10)$$

From equation (8.8) $P = [\text{gas}] RT$. Substituting (10.8) in (10.10) we get:

$$K_p = \frac{[\text{NO}_2]^2 (RT)^2}{[\text{N}_2\text{O}_4] RT} = \frac{[\text{NO}_2]^2 (RT)}{[\text{N}_2\text{O}_4]} \dots\dots\dots (10.11)$$

Recall that from equation (10.9) that $K_c = K_c = \frac{[\text{NO}_{2(g)}]^2}{[\text{N}_2\text{O}_{4(g)}]}$

Therefore substituting (10.9) in (10.11) we get

$$K_P = K_c RT \dots\dots\dots \text{Equation 10.12}$$

Note: This is not always the case.

Let us consider another example:



$$PV = nRT \dots\dots\dots (10.13)$$

$$P = \frac{nRT}{V} \dots\dots\dots (10.14)$$

$$P = [\text{gas}]RT \dots\dots\dots (10.15)$$

$$K_c = \frac{[\text{NH}_{3(g)}]^2}{[\text{N}_{2(g)}][\text{H}_{2(g)}]^3} \dots\dots\dots (10.16)$$

$$K_c = K_p = \frac{(P_{\text{NH}_{3(g)}})^2}{(P_{\text{N}_{2(g)}})(P_{\text{H}_{2(g)}})^3} \dots\dots\dots (10.17)$$

$$K_p = \frac{(P_{\text{NH}_3})^2}{(P_{\text{N}_2})(P_{\text{H}_2})^3} \dots\dots\dots (10.18)$$

$$K_p = \frac{[\text{NH}_3]^2(RT)^2}{[\text{N}_{2(g)}]RT[\text{H}_{2(g)}]^3(RT)^3} \dots\dots\dots (10.19)$$

$$K_p = \frac{[\text{NH}_3]^2}{[\text{N}_{2(g)}][\text{H}_{2(g)}]^3(RT)^2} \dots\dots\dots (10.20)$$

Substituting (10.16) in (10.20) we get

$$K_P = K_C (RT)^2 \dots\dots\dots (10.21)$$

The relationship between K_p and K_c can be generalized thus

$$K_P = K_C(RT)^{A_n}$$

Where;

$$A_n = (\text{No. of moles on product side}) - (\text{No. of moles on reactant side})$$

Note: In this calculation, temperature must be expressed in Kelvin; pressure must be expressed in pascal.

Celsius temperature + 273 = Kelvin temperature 1 atm \approx 105 Pascal

10.7 HETEROGENEOUS EQUILIBRIA

So far, we have been talking about a single phase equilibria (homogeneous equilibria). An equilibria involving more than one phase is called **heterogeneous**

equilibria.

An example is $\text{CaCO}_{3(s)} \rightleftharpoons \text{CaO}_{(s)} + \text{CO}_{2(g)}$

Here, CaCO_3 and CaO are in a solid phase While CO_2 is in a gaseous phase.

The equilibrium expression for the reaction can be written thus:

$$K_C = \frac{[\text{CO}_{(s)}][\text{CO}_{2(g)}]}{[\text{CaCO}_{3(s)}]}$$

Since the ratio of moles to volume is constant for solid, then their concentration is constant. Therefore the equation can be simplified by writing the equilibrium expression „in such a way that the concentration of solid and liquid (note, not aqueous) is removed in the equilibrium expression. Therefore, our equation becomes:

$$K_C = [\text{CO}_{2(g)}] \text{ OR } K_P = P_{\text{CO}_{(g)}}$$

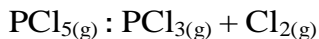
Also for $\text{CaO}_{(s)} + \text{SO}_{2(g)} \rightleftharpoons \text{CaSO}_{3(s)}$

We have:

$$K_C = \frac{1}{[\text{SO}_{2(g)}]} \text{ OR}$$

$$K_P = \frac{1}{P_{\text{SO}_{2(g)}}}$$

Examples 10.1:



If the equilibrium concentrations of the PCl_5 and PCl_3 ; are 0.015 mol/dm³ and 0.78 mol/dm³ respectively, what is the concentration of Cl_2 if K_{eq} of the system is 35 at a certain temperature?

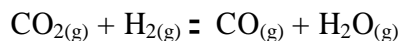
$$K_{eq} = \frac{[\text{Cl}_2][\text{PCl}_3]}{[\text{PCl}_5]}$$

$$35 = \frac{[\text{Cl}_2]0.78}{0.015}$$

$$\therefore [\text{Cl}_2] = \frac{35 \times 0.015}{0.78} = 0.67 \text{ mol/dm}^3$$

Example 10.2:

At equilibrium a 1.0L vessel contains 20mol of H₂, 18mol of CO₂, 12.0mol of water (H₂O) and 5.9mol of CO at 427°C. What is the value of K at this temperature?

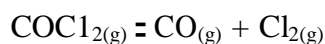


$$K_{\text{eq}} = \frac{[\text{CO}][\text{H}_2\text{O}]}{[\text{CO}_2][\text{H}_2]}$$

$$K_{\text{eq}} = \frac{5.9 \times 12.0}{18 \times 20} = 0.196 \quad = 0.2$$

Example 10.3:

At equilibrium, the quantities of the substances present in a 20dm³ container are: 1.70 mole of COCl₂, 0.76mol of CO and 1.50 mole of Cl₂ for the temperature at which equilibrium is reached. What is K_{eq} for the reaction?



$$[\text{COCl}_2] = \frac{1.70}{20} = 0.085$$

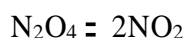
$$[\text{CO}] = \frac{0.76}{20} = 0.038$$

$$[\text{Cl}_2] = \frac{1.50}{20} = 0.075$$

$$K_{\text{eq}} = \frac{[\text{CO}][\text{Cl}_2]}{[\text{COCl}_2]} = \frac{0.038}{0.085} \times 0.075 = 0.0335 \text{ mol/dm}^2$$

Example 10.4:

When 1.0mol of N₂O₄ is placed into a 5.0dm³ container at 100°C, part of it decomposed to form NO₂. At equilibrium, when the temperature is 100°C. 1.0mol of NO₂ is present. Calculate K for the reaction at 100°C. The balanced equation for the reaction is:



| | | |
|----------|-------------------------------|-----------------|
| | N ₂ O ₄ | NO ₂ |
| At start | 1.0mol | 0 |

$$= \frac{1.0}{5} = 0.2 \text{ mol/dm}^3$$

| | | |
|----------------|-----|--------|
| At equilibrium | 0.1 | 1.0mol |
|----------------|-----|--------|

$$= \frac{1.0}{5} = 0.2 \text{ mol/dm}^3$$

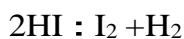
From the equation, 1 mol N_2O_2 decomposes to give 2mol of NO_2 . Therefore 0.1 was decomposed to give 0.2mol NO_2 and so 0.1 N_2O_4 will be left at equilibrium.

$$K_{\text{eq}} = \frac{(1.0)^2}{0.1} = \frac{0.04}{0.1} = 0.40$$

Example 10.5:

A 1.00mol sample of $\text{HI}_{(\text{g})}$ is heated to 510°C in a sealed flask with a volume of 1.00dm^3 . At equilibrium, 0.14mol of each of the product, $\text{H}_{2(\text{g})}$ and $\text{I}_{2(\text{g})}$ is present.

- Calculate the number of moles of HI that are present at equilibrium.
- Write the equilibrium expression for the reaction.
- Calculate K_{eq}



| | HI | I_2 | H_2 |
|------------------------|------|--------------|--------------|
| Moles before reaction | 1.0 | 0 | 0 |
| Moles at equilibrium x | 0.14 | 0.14 | |

From the equation, 2 moles of HI decomposes to give 1 mole of I_2 and H_2 . Therefore at equilibrium, 0.28 HI is used up and 0.72 HI will be left at equilibrium.

$$\text{HI left at equilibrium} = 1 - 0.28 = 0.72 \text{ moles}$$

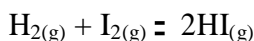
$$K_{\text{eq}} = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2}$$

$$K_{\text{eq}} = \frac{[0.14][0.14]}{[0.72]^2}$$

$$K_{\text{eq}} = \frac{0.0196}{0.5184} = 0.378 \text{ or } 3.73 \times 10^{-2}$$

Example 8.6:

Hydrogen and iodine react at 426°C according to:



If 1.00mol of H_2 and 1.00mol, I_2 are placed in a 1.00 litre vessel and allowed to react, what weight of HI will be present at equilibrium. $K_{\text{eq}} = 55.3$

| | H_2 | I_2 | 2HI |
|----------------|--------------|--------------|--------------|
| At start | 1 | 1 | 0 |
| At equilibrium | $1 - x$ | $1 - x$ | $2x$ |

$$K_{eq} = \frac{(2x)^2}{(1-x)^2}$$

$$55.3 = \frac{(2x)^2}{(1-x)^2}$$

$$\sqrt{55.3} = \frac{2x}{1-x}$$

$$7.4 (1 - x) = \frac{2x}{1-x}$$

$$7.4 (1 - x) = 2x$$

$$7.4 - 7.4x = 2x$$

$$7.4 = 2x + 7.4x$$

$$7.4 = 9.4x$$

$$x = \frac{7.4}{9.4} = 0.79 \text{ mol}$$

$$\text{Concentration of HI} = 2x$$

$$= 2 \times 0.79$$

$$= 1.6 \text{ mole}$$

$$\text{Reacting mass} = \text{Molar mass} \times \text{No. of mole}$$

$$= 126.9 \times 1.6$$

$$= 203.04 \text{ g}$$

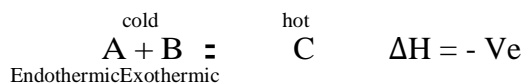
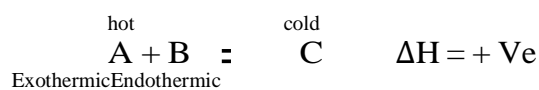
10.8 LECHATelier'S PRINCIPLE

Lechatelier's Principle states that if an external constraint such as change in temperature, pressure or concentration is imposed on a chemical system in equilibrium, the equilibrium position will shift so as to annul the constraint. This principle is used to determine the optimum operating condition for chemical processes in industries.

10.8.1 EFFECT OF TEMPERATURE ON EQUILIBRIUM

The effect of temperature on equilibrium of reversible reaction depends on the value of ΔH (enthalpy change) of the reaction. A positive value for ΔH implies that the reaction is endothermic in the forward reaction i.e. heat is absorbed in the forward reaction and the product tend to be cooler, while the backward is exothermic. A negative value of ΔH implies that the

reaction is exothermic in the forward reaction i.e. heat is evolved, or liberated in the forward reaction and the product will tend to be hotter, while the backward is endothermic.



According to Lechatechier's principle, a change in temperature will cause the equilibrium position to shift so as to annul the effect of the change in temperature. Therefore in an endothermic reaction increase in temperature will shift the equilibrium to the forward reaction. While decrease in temperature, will favour the backward reaction. In the same way, in an exothermic reaction, increase in temperature will shift the equilibrium position to the backward reaction and decrease in temperature will favour the forward reaction.

Examples 10.6



This is an endothermic reaction. An increase in the temperature will shift the equilibrium to the forward reaction (right) giving a higher value of K meaning a greater yield of product. While a decrease will favour the backward reaction (left), giving a lower value of K which means a greater yield of reactant.

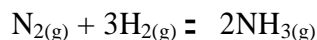


This is an exothermic reaction. An increase in temperature will shift the equilibrium to the left (backward reaction), while a decrease in temperature will shift the equilibrium position to the right.

10.8.2 EFFECT OF PRESSURE ON EQUILIBRIUM

For a change in pressure to affect a chemical equilibrium, the following must be in place;

1. At least one of the reactants or products in the reversible reaction must be gaseous.
2. The total moles of gaseous molecules in the forward and reversed reaction must not be the same e.g.



Backward reaction has 1 mole N_2 and 3 moles H_2 making a total of 4 moles, while forward reaction has 2 moles NH_3 .

According to Lechatechier's principle, if a high pressure is imposed on a system in equilibrium, the equilibrium will shift to the position of low pressure. Also if low pressure is imposed on an equilibrium system, the equilibrium position will shift to the position of higher pressure.

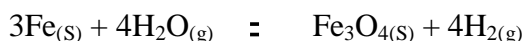
It should be noted however, that the position of high pressure is that with the greater volume or greater number of moles. The position of low pressure is that with less volume or lesser numbers of moles.

In the reaction above, increase in pressure will favour the forward reaction while decrease in pressure will favour the backward reaction. In both cases the equilibrium constant K remains unchanged.

10.8.3 EFFECT OF CONCENTRATION ON EQUILIBRIUM

In an equilibrium mixture, there is a balance between the concentrations of the reactants and the products i.e. these concentrations are in a definite ratio which depends on the condition of the reaction. If more reactants are introduced into the equilibrium system the equilibrium will shift to the right favouring the forward reaction. This results in a proportional increase in the concentration of the products and so the equilibrium constant remains constant. Another way of increasing the amount of products formed is by continually removing the products from the system. Again in order to remove the constraint, the equilibrium will shift to the right, favouring the forward reaction.

Let us consider the reaction below:



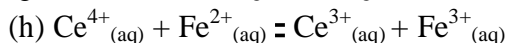
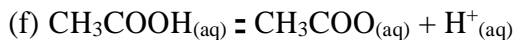
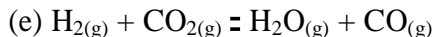
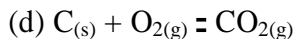
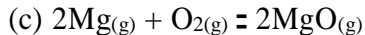
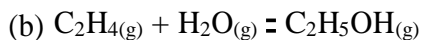
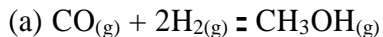
If the concentration of a reactant or product is changed, the equilibrium position shifts to annul the change. Increase in the concentration of Iron will shift the equilibrium to the right and a decrease in hydrogen concentration will also favour the forward reaction. But when Fe_3O_4 is increase or H_2O is reduced, the backward reaction is favoured.

10.8.4 EFFECT OF CATALYST ON EQUILIBRIUM

Catalyst does not affect the position of equilibrium. It cannot be used to increase product formation. However, catalyst is very important in chemical industry because they affect the time taken for a reaction to reach equilibrium by lowering the activation energy for a positive catalyst and increasing the activation energy for a negative catalyst.

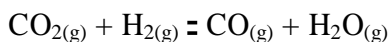
PRACTICE QUESTION

1. Write down the equilibrium constant for each of the following reaction?



Use K or K_c depending on the states of the chemical.

2. At equilibrium at 1.0L vessel contains 20.0mol of H₂, 18.0mol of CO₂ of 12.0mol of H₂O and 5.9mol of CO at 427°C. What is the value of K at this temperature? Equation for the reaction is:



3. A reaction between gaseous sulphur dioxide (SO₂) and Oxygen (O₂) gas to produce gaseous sulphur trioxide (SO₃) takes place at 600°C. At that temperature, the concentration of SO₂ is found to be 1.50mol/L, the concentration of O₂ is 1.25mol/L, and the concentration of SO₃, is 3.50mol/L. Using the balanced chemical equation, calculate the equilibrium constant for the system.
4. For the reaction:
- $$\text{H}_{2(\text{g})} + \text{I}_{2(\text{g})} \rightleftharpoons 2\text{HI}_{(\text{g})} \text{ at } 425^\circ\text{C}.$$

Calculate [HI], given [H₂] = [I₂] = 4.79 x 10⁻⁴ mol/L and K = 54.3

5. What is the effect of increase in (i) temperature (ii) pressure and (iii) concentration on the equilibrium reaction?

