

# Edexcel Chemistry A-level

## Topic 11: Equilibrium II

### Detailed Notes

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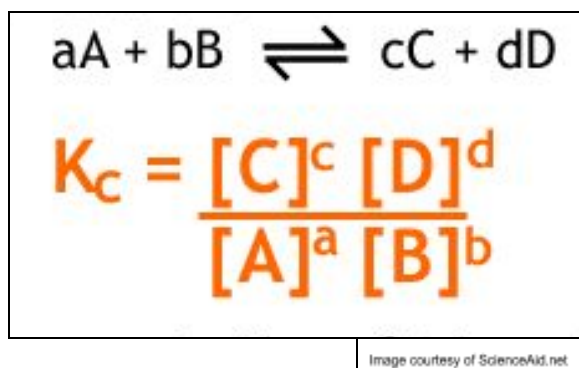




## Calculating K<sub>c</sub>

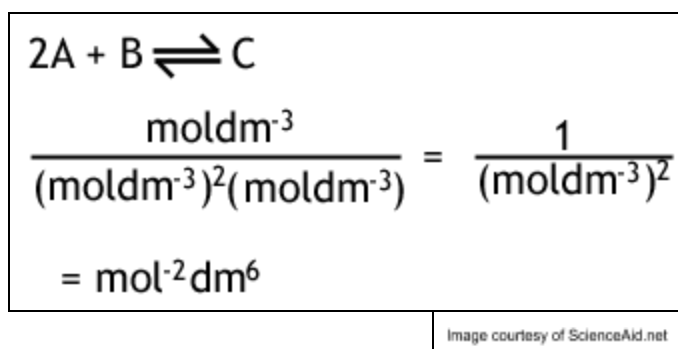
K<sub>c</sub> is equal to the **concentration of the products divided by the concentration of the reactants**. The concentration terms are **raised to a power** of the same value as the number of moles of that substance.

*Example:*



The equilibrium constant has **varying units**, depending on the chemical reaction. The units can be calculated by **substituting the concentration units into the K<sub>c</sub> expression**. Some of these units will then cancel, giving the overall units of K<sub>c</sub> for that reaction.

*Example:*



## Homogeneous and Heterogeneous

The equilibrium constant, K<sub>c</sub>, can be found for both **homogeneous** and **heterogeneous** reactions. Homogeneous reactions are reactions in which the reactants and products are in the same **phase**, whereas heterogeneous reactions are reactions in which some of the reactants and/or products are in different phases to each other.

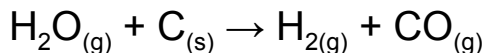
For **homogeneous** reactions, K<sub>c</sub> is calculated as shown **above**.

The **difference** when calculating K<sub>c</sub> for **heterogeneous** reactions is that any terms representing a **solid** are **not included** in the calculation.





Example:



The solid carbon is not included in the equation for  $K_c$ :

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

## Gaseous Equilibrium Constant ( $K_p$ )

$K_p$  is the equilibrium constant used for **gaseous equilibria**.  $K_p$  is calculated from gaseous reactants and products. If all reactants and products are in the **gaseous state**, the system is said to be **homogeneous**.

### Partial Pressure

Within a gaseous system, each gas has a **partial pressure**. The partial pressures add up to give the total system pressure. This partial pressure of a substance is found using the **molar fraction** of that substance and the **total pressure of the system**.

$$\text{Partial Pressure of A} = \frac{\text{Moles of A}}{\text{Total Moles}} \times \text{Total Pressure}$$

Partial pressure of A would be shown as (**PA**). Partial pressures are commonly measured in **Pascals** but are occasionally measured in atmospheres.

Example:

Image courtesy of SlidePlayer

A mixture of gases contains 0.51 mol  $\text{N}_2$ , 0.28 mol  $\text{H}_2$ , and 0.52 mol  $\text{NH}_3$ . If the total pressure of the mixture is 2.35 atm, what is the partial pressure of  $\text{H}_2$ ?

$$\begin{aligned} \text{Total moles in the system} &= 0.51 + 0.28 + 0.52 \\ &= 1.31 \text{ moles} \end{aligned}$$

$$\begin{aligned} \text{Molar fraction of } \text{H}_2 &= 0.28 / 1.31 \\ &= 0.21 \end{aligned}$$

$$(P_{\text{H}_2}) = 0.21 \times 2.35 = 0.50 \text{ atm}$$





## Calculating Kp

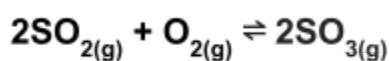
Partial pressures allow the value of Kp for a gaseous equilibrium to be found. Kp is equal to the product of the **partial pressures of products** over the product of the **partial pressure of reactants**. It is similar to Kc in that any variation in moles **raises the partial pressure to a power** of equal quantity to the number of moles.



$$K_p = \frac{(P_Y)(P_Z)^2}{(P_A)^2(P_B)^3}$$

The gaseous equilibrium constant has **varying units**, depending on the chemical reaction. The units can be calculated by **substituting the partial pressure units into the Kp expression**. Some of these units will then cancel, giving the overall units of Kp for that reaction.

Example:



$$K_p = \frac{(kPa)^2}{(kPa)^2 \times kPa} \\ = kPa^{-1}$$

## Homogeneous and Heterogeneous

Similarly to Kc, the gaseous equilibrium constant, Kp, can be found for both **homogeneous** and **heterogeneous** reactions.

For **homogeneous** reactions, Kp is calculated as shown **above**.

The **difference** when calculating Kp for **heterogeneous** reactions is that any terms representing a **solid** are **not included** in the calculation.



The solid CaCO<sub>3</sub> and CaO are not included in the equation for Kp:  $K_p = (P_{CO_2})$





## Factors Affecting $K_c$ and $K_p$

The values of  $K_c$  and  $K_p$  are **not affected by concentration or pressure change** or by the use of a **catalyst**. However, they are affected by changing the reaction **temperature**.

**Catalyst** affect the **rate** of the reaction (the kinetics) but not the **position** of the equilibrium. They only affect how fast the system reaches equilibrium, hence they have **no impact** on the equilibrium constant.

**Temperature**, on the other hand, does affect the position of the equilibrium, as this changes the **position** of the equilibrium, resulting in different concentrations of reactants and products. Therefore, the equilibrium constant changes with temperature.

If the forward reaction is **exothermic**, an increase in temperature will decrease the rate of the forward reaction because the equilibrium shifts to the **left** to oppose the change. This will decrease the concentrations of products and increase the concentrations of reactants, therefore the equilibrium constant ( $K_c$  or  $K_p$ ) **decreases**.

If the forward reaction is **endothermic**, an increase in temperature will increase the rate of the forward reaction because the equilibrium shifts to the **right** to oppose the change. This will increase the concentrations of products and decrease the concentrations of reactants, therefore the equilibrium constant ( $K_c$  or  $K_p$ ) **increases**.

Similar arguments can be made for the effect of **decreasing** the temperature.

