# Edexcel Chemistry A-level

## Topic 11: Equilibrium II

Detailed Notes

## Calculating Kc

Kc 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:*

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

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

Image courtesy of ScienceAid.net

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

*Example:*

$$2A + B \rightleftharpoons C$$

$$\frac{moldm^{-3}}{(moldm^{-3})^2(moldm^{-3})} = \frac{1}{(moldm^{-3})^2}$$

$$= mol^{-2}dm^6$$

Image courtesy of ScienceAid.net

### Homogeneous and Heterogeneous

The equilibrium constant, Kc, 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, Kc is calculated as shown **above**.

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

*Example:* $H_2O_{(g)} + C_{(s)} \rightarrow H_{2(g)} + CO_{(g)}$

The solid carbon is not included in the equation for Kc:

$$Kc = \frac{[H_2]\,[CO]}{[H_2O]}$$

## Gaseous Equilibrium Constant (Kp)

Kp is the equilibrium constant used for **gaseous equilibria**. Kp 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 **($P_A$).** Partial pressures are commonly measured in **Pascals** but are occasionally measured in atmospheres.

*Example:*

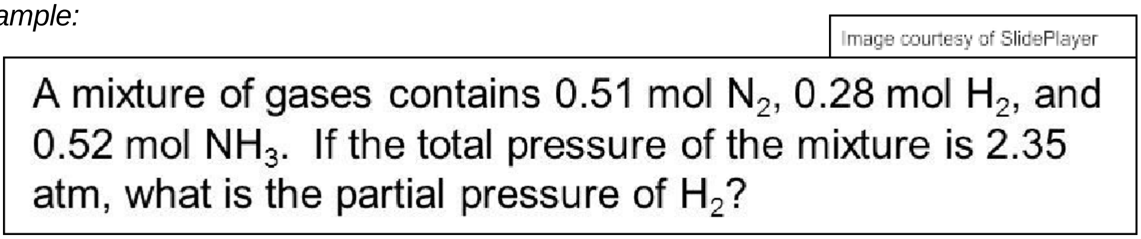

Image courtesy of SlidePlayer

A mixture of gases contains 0.51 mol $N_2$, 0.28 mol $H_2$, and 0.52 mol $NH_3$. If the total pressure of the mixture is 2.35 atm, what is the partial pressure of $H_2$?

***Total moles in the system = 0.51 + 0.28 + 0.52***
***= 1.31 moles***
***Molar fraction of $H_2$ = 0.28 / 1.31***
***= 0.21***
***($P_{H_2}$) = 0.21 x 2.35 = 0.50 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.

$$2A_{(g)} + 3B_{(g)} \longrightarrow Y_{(g)} + 2Z_{(g)}$$

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

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:*

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

$$Kp = \frac{(kPa)^2}{(kPa)^2 \times kPa}$$

$$= kPa^{-1}$$

## 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.

Example: $CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$

The solid $CaCO_3$ and CaO are not included in the equation for Kp: $Kp = (P_{CO_2})$

## Factors Affecting Kc and Kp

The values of Kc and Kp 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 (Kc or Kp) **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 (Kc or Kp) **increases**.

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