

4.3 Equilibrium amounts, K_c and K_p

Name: _____ Class: _____ Date: _____

Total: 19 marks

KEY WORDS partial pressure

Objective

Calculate an equilibrium constant from the actual equilibrium composition. Questions and marking guidance are original.

1 Key knowledge and worked examples

Write the balanced equation before the expression. Raise each concentration or partial pressure to its coefficient; omit pure solids and liquids. Constants refer to a specified temperature. A catalyst leaves K unchanged.

■ Worked method

For $A(g) + B(g) \rightleftharpoons C(g)$, initially 0.500 mol each A and B in 2.00 dm³ gives 0.200 mol C at equilibrium. The equilibrium amounts are A = 0.300, B = 0.300, C = 0.200 mol.

$$K_c = \frac{[C]}{[A][B]} = \frac{0.200/2.00}{(0.300/2.00)^2} = 4.44 \text{ dm}^3 \text{ mol}^{-1}$$

For gases, $p_i = x_i P$ where $x_i = n_i/n_{\text{total}}$. Derive K_p units from the powers and use the specified pressure unit. For an exothermic forward reaction, raising temperature lowers K. Concentration/pressure changes cause a shift while the same K is restored at fixed temperature.

2 Practice

2.1 Write K_c for $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, and derive its units for concentrations in mol dm⁻³. [3]

2.2 A gas mixture contains 2.0 mol X and 3.0 mol Y at total pressure 5.0 atm. Calculate each partial pressure. [2]

2.3 Write K_p for $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$. Explain the omitted terms. [2]

3 Exam-style questions

3.1 For $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$, equilibrium amounts are 0.60 mol N_2O_4 and 0.40 mol NO_2 at 2.50 atm total pressure. Calculate K_p with units. [4]

3.2 For $\text{A} + \text{B} \rightleftharpoons \text{C}$, 0.40 mol each A and B initially in 1.50 dm^3 produces 0.10 mol C. Calculate K_c . Explain why simply using the equilibrium mole amounts in the expression fails. [4]

3.3 An appropriately standardised dimensionless equilibrium constant is $K = 20.0$ at 298 K. Use $\Delta S(\text{total}) = R \ln K$ and $R = 8.31 \text{ J mol}^{-1} \text{ K}^{-1}$ to calculate total entropy. The forward reaction is endothermic. Explain the predicted change in K when temperature rises, assuming ΔH and system entropy are approximately constant. [4]

Solutions

Original marking guidance: M1 method, A1 accuracy, B1 explanation.

2.1 $K_c = [NH_3]^2 / ([N_2][H_2]^3)$. Units $\text{dm}^6 \text{mol}^{-2}$.

2.2 $p_X = (2/5)5 = 2.0 \text{ atm}$; $p_Y = (3/5)5 = 3.0 \text{ atm}$.

2.3 $K_p = p(\text{CO}_2)$; pure solids do not enter the expression, while both solid phases remain present.

3.1 Total amount 1.00 mol; partial pressures 1.50 and 1.00 atm. $K_p = 1.00^2 / 1.50 = 0.667 \text{ atm}$.

3.2 Equilibrium A and B are 0.30 mol each. Concentrations are 0.200, 0.200 and $0.0667 \text{ mol dm}^{-3}$; $K_c = 0.0667 / 0.200^2 = 1.67 \text{ dm}^3 \text{mol}^{-1}$. Unequal total powers mean volume factors do not cancel.

3.3 $\Delta S_{\text{total}} = 8.31 \ln(20.0) = +24.9 \text{ J mol}^{-1} \text{K}^{-1}$. For positive ΔH , the negative surroundings term $-\Delta H/T$ becomes smaller in magnitude as T rises, so total entropy and $\ln K$ increase, hence K increases.