

Gas Equilibrium and Pressure Constant (K_p)

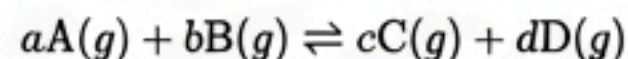
Mastering Partial Pressures and the $K_p \rightleftharpoons K_c$ Relationship for JEE

Partial pressure is the partial pressure of gaseous reactants and products, where is the equation by:

$$P_{\text{total}} = P_A + P_B + P_C + \dots$$

Note: Dalton's Law is a what colne chemain and courm? reaction of partial pressure, of the partial pressure.

Equilibrium Constant K_p with th the following reaction, this express to:



The equilibrium expression is:

$$K_p = \frac{P_C^c * P_D^d}{P_A^a * P_B^b}$$

where P_x represents the partial pressure of gas x at equilibrium.

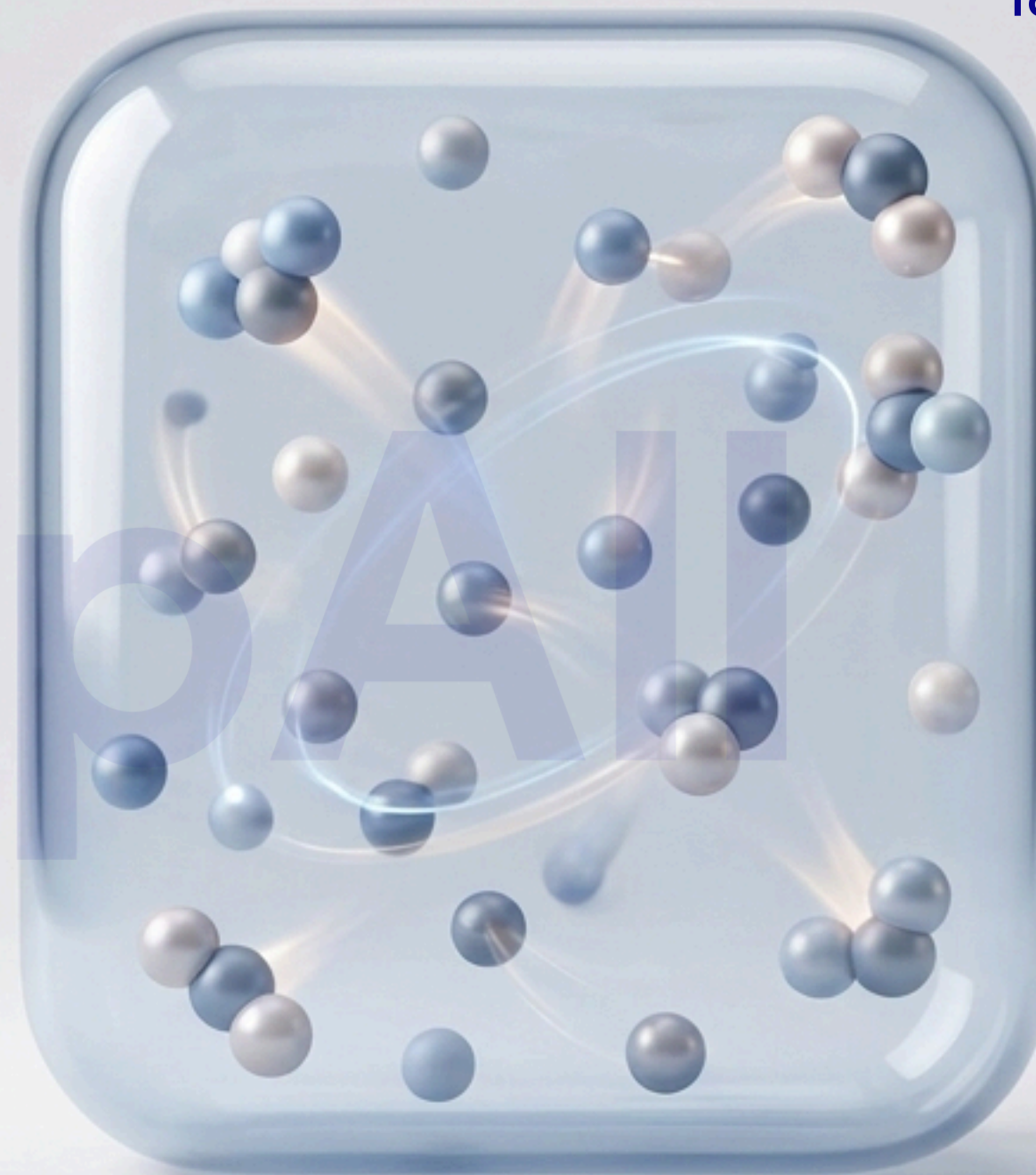
K_p is defined in terms of partial pressures of gaseous reactants and products. It is applicable only when all species in the equilibrium expression are in the gaseous phase.

Key Takeaway: The relationship between K_p and K_c is given by

$$K_p = K_c(RT)^{\Delta n_g},$$

where R is the ideal gas constant, T is the temperature in Kelvin, and $\Delta n_g = (c + d) - (a + b)$ is the change in the number of moles of gaseous substances.

Caution: Be mindful of the units of R. Use $R = 0.0821 \text{ L-atm/(mol-K)}$ when pressure is in atm, and $R = 8.314 \text{ J/(mol-K)}$ when pressure is in Pa and volume is in m^3 . Also, ensure T is always in Kelvin.



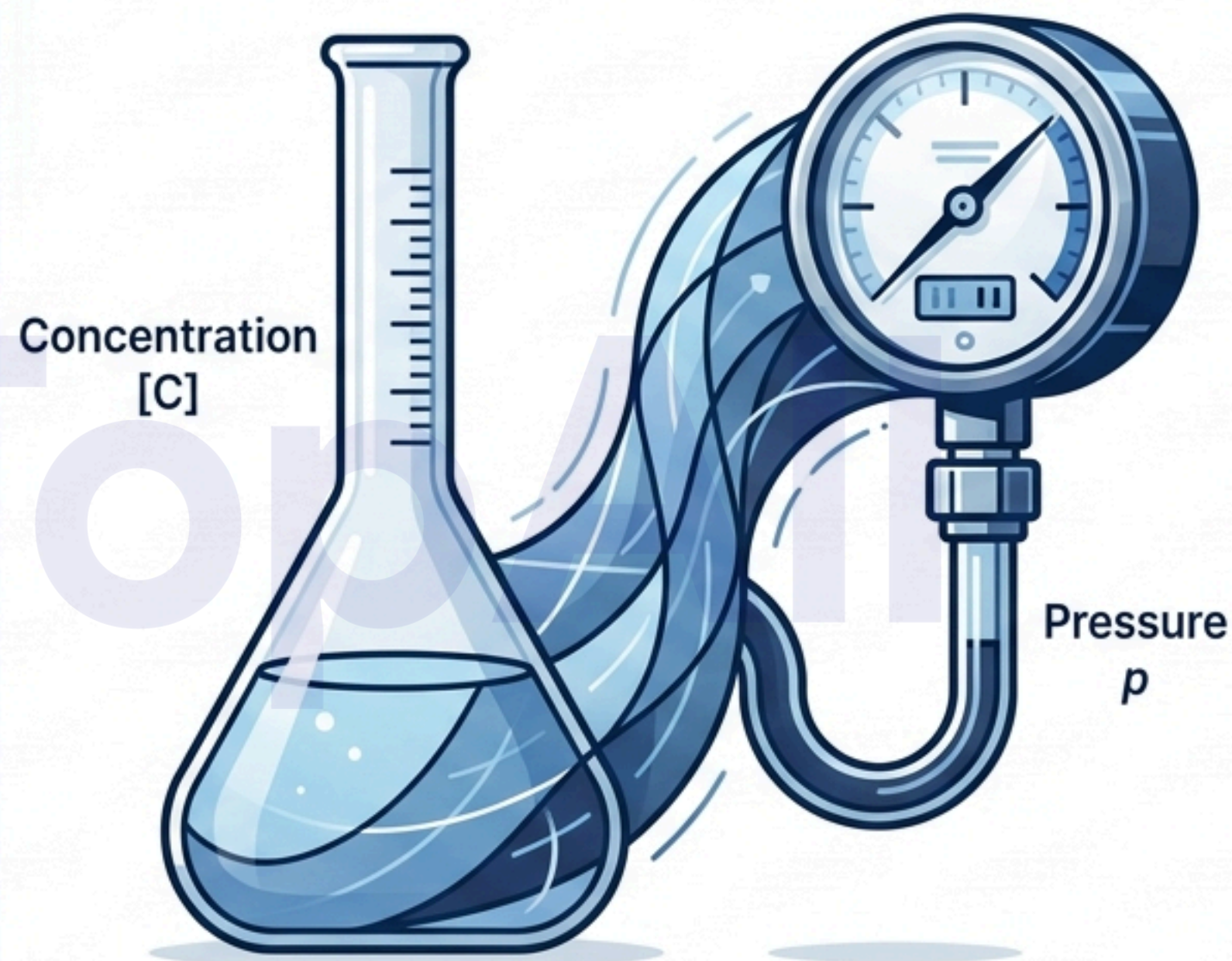
The Concept

Why Use K_p for Gases?

While K_c uses molar concentration (mol/L), measuring the exact concentration of gases in a mixture can be highly impractical. For gaseous systems, it is significantly more convenient to express the equilibrium constant in terms of partial pressure.

The Core Principle:

At a constant temperature, the pressure of a gas is directly proportional to its concentration. Therefore, partial pressures accurately represent the active mass of gaseous species in an equilibrium mixture.



First Principles

Linking Pressure to Concentration

How do we convert K_c to K_p ? We bridge the gap using the **Ideal Gas Equation**.



$$pV = nRT$$

↓ Rearranging for pressure p

$$p = (n/V)RT$$

↓ Since concentration $c = n/V$ in mol/L

$$p = cRT$$

↓ Expressed for any specific gas

$$p_{\text{gas}} = [\text{gas}]RT$$

Standard State Note: When calculating K_p , pressure should be expressed in **bar** (Standard state = 1 bar). Remember this constant for JEE: $R = 0.0831 \text{ L bar K}^{-1} \text{ mol}^{-1}$.



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Derivation

Formulating K_p for a General Reaction

Consider the general gas-phase reaction at equilibrium:



Instead of square brackets for concentration $[X]$, we use p_x to denote the equilibrium partial pressure of the gas. The stoichiometric coefficients still become the exponents.

The Equilibrium Law for Pressures:

$$K_p = \frac{(p_C)^c (p_D)^d}{(p_A)^a (p_B)^b}$$

Where:

p_A, p_B, p_C, p_D = Equilibrium partial pressures of the reactants and products.

Connecting K_p and K_c

By substituting $p = [\text{gas}]RT$ into the K_p expression:

$$K_p = \frac{([C]RT)^c ([D]RT)^d}{([A]RT)^a ([B]RT)^b} = \frac{[C]^c [D]^d}{[A]^a [B]^b} \times \frac{(RT)^{c+d}}{(RT)^{a+b}}$$

$$K_p = K_c (RT)^{\Delta n}$$

Where $\Delta n = (\text{moles of gaseous products}) - (\text{moles of gaseous reactants})$

If $\Delta n = 0 \rightarrow K_p = K_c$	(e.g., $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$)
If $\Delta n > 0 \rightarrow K_p > K_c$	(e.g., $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$)
If $\Delta n < 0 \rightarrow K_p < K_c$	(e.g., $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$)

Identify Gaseous Species ONLY!

For **Heterogeneous Equilibria**, pure solids (s) and pure liquids (l) are fundamentally excluded from the K_p (and K_c) expressions. Their concentrations and partial pressures are considered constant and are incorporated into the equilibrium constant.

Thermal decomposition of Limestone:



The Trap (Incorrect):

$$K_p = \frac{(p_{\text{CaO}})(p_{\text{CO}_2})}{p_{\text{CaCO}_3}}$$



The Truth (Correct):

$$K_p = p_{\text{CO}_2}$$

Chapter Master Summary

Formula Bank

$$p = [\text{gas}]RT$$

$$K_p = \frac{(p_C)^c (p_D)^d}{(p_A)^a (p_B)^b}$$

$$K_p = K_c (RT)^{\Delta n}$$

$$\Delta n = n_{p(\text{gas})} - n_{r(\text{gas})}$$

Key Takeaways

- 1. Convenience:** K_p is used exclusively for gas-phase reactions because measuring partial pressure is easier than measuring molarity.
- 2. Units matter:** Standard state pressure is 1 bar. When calculating K_p , ensure $R = 0.0831$.
- 3. State Symbols are critical:** Always ignore (s) and (l) phases in your K_p calculations. Only count the moles of (g) to find Δn .



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