

Chemical Equilibrium and K_c

What is chemical equilibrium and how do you calculate K_c?

●●○ Intermediate

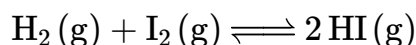
Equilibrium

Last reviewed 4 October 2026

Before you start: it helps to understand [Molarity and Solution Stoichiometry](#) and [Balancing Chemical Equations](#).

What is it?

Many reactions do not go to completion. They are **reversible**: the products can react to reform the reactants. In a closed container, a reversible reaction reaches **chemical equilibrium**:



At equilibrium the **forward and reverse reactions continue at equal rates**, so the concentrations of reactants and products stay **constant**. This is a **dynamic** equilibrium: molecules keep reacting in both directions, but there is no overall change.

💡 KEY IDEA

At equilibrium, the concentrations are constant, **not equal**. The ratio between them is fixed by the **equilibrium constant**, K_c .

Why does it matter?

- **Yield.** Equilibrium sets the maximum amount of product a reversible reaction can give, which matters in industry (ammonia, sulfuric acid, methanol).
- **Acids, bases and solubility.** K_a for weak acids, buffers and solubility are all equilibrium constants.
- **The body.** Oxygen binding to haemoglobin and the carbonic acid buffer in blood are equilibria.

How does it work?

1. The equilibrium constant, K_c

For a general reaction $a\text{A} + b\text{B} \rightleftharpoons c\text{C} + d\text{D}$, the **equilibrium constant expression** is:

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

- Products go on top, reactants on the bottom.
- Each concentration (in mol/L) is raised to the power of its **coefficient** in the balanced equation.
- **Pure solids and pure liquids are left out** (their “concentration” does not change). For $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$, $K_c = [\text{CO}_2]$.
- K_c is constant at a given **temperature**; it changes only when the temperature changes.

Units of K_c . Strictly, each concentration is divided by the standard concentration $c^\circ = 1 \text{ M}$, so the units cancel and K_c has **no unit**. The worked examples show this step.

2. What the size of K tells you

Value of K_c	At equilibrium the mixture contains...
very large (e.g. 10^{10})	mostly products: the reaction goes almost to completion
about 1	significant amounts of both
very small (e.g. 10^{-10})	mostly reactants: very little reaction

3. Related equilibria

- **Reverse** the equation: $K_{\text{reverse}} = \frac{1}{K}$
- **Multiply** the equation by n : $K_{\text{new}} = K^n$ (for example, halving the coefficients gives $\sqrt{K}$)

4. The reaction quotient, Q

Q has the same form as K , but uses the concentrations **at any moment**, not only at equilibrium. Comparing Q with K predicts the direction of change:

Comparison	What happens
$Q < K$	too few products: the reaction goes forward ($\rightarrow$)
$Q = K$	at equilibrium: no net change
$Q > K$	too many products: the reaction goes in reverse ($\leftarrow$)

5. ICE tables

To find equilibrium concentrations, use an **ICE table**: **I**nitial, **C**hange, **E**quilibrium. The changes follow the mole ratio of the equation, written in terms of an unknown, x .

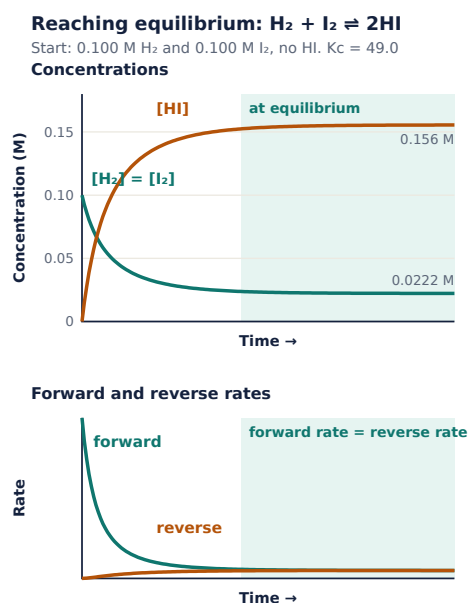
∞ THINK OF IT LIKE THIS

Picture two shops connected by a door, with customers walking both ways. After a while, as many people walk from shop A to shop B each minute as walk back. The number of people in each shop stays the same, even though people never stop moving. That's dynamic equilibrium, and the shops need not hold equal numbers.

🔍 MORE PRECISELY

The thermodynamic equilibrium constant is defined with **activities**, which for dilute solutions are approximately $[X]/c^\circ$ and for gases p_X/p° . That is why equilibrium constants are dimensionless. For gases, K_p uses partial pressures instead of concentrations; K_p and K_c are equal only when the number of moles of gas is the same on both sides.

Visualise it



At equilibrium the concentrations stop changing because the forward and reverse rates are equal. Calculated from a simple rate model.

Worked example

WORKED EXAMPLE: Calculating K_c

Question: At a certain temperature, an equilibrium mixture of $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$ contains $[\text{N}_2\text{O}_4] = 0.0450 \text{ M}$ and $[\text{NO}_2] = 0.0300 \text{ M}$. Calculate K_c .

Step 1 Write the expression: $K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$

Step 2 Substitute, dividing each concentration by $c^\circ = 1 \text{ M}$:

$$\begin{aligned} K_c &= \frac{(0.0300 \text{ M}/1 \text{ M})^2}{0.0450 \text{ M}/1 \text{ M}} \\ &= \frac{(0.0300)^2}{0.0450} = 0.0200 \end{aligned}$$

Step 3 The units cancel: $K_c = \mathbf{0.0200}$.

WORKED EXAMPLE: Predicting the direction with Q

Question: For $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$, $K_c = 49.0$. A mixture contains 0.050 M H_2 , 0.050 M I_2 and 0.20 M HI . Is it at equilibrium? If not, which way will it shift?

Step 1 $Q = \frac{(0.20 \text{ M})^2}{(0.050 \text{ M})(0.050 \text{ M})} = 16$ (the units of M^2 cancel)

Step 2 $Q = 16 < K_c = 49.0$, so there is too little HI : the reaction goes **forward**, making more HI .

 WORKED EXAMPLE: An ICE table

Question: 0.100 M H_2 and 0.100 M I_2 are mixed with no HI ($K_c = 49.0$). Find the equilibrium concentrations.

Step 1 ICE table, where x is the decrease in $[\text{H}_2]$, in M:

	H_2	I_2	HI
Initial	0.100 M	0.100 M	0 M
Change	$-x$	$-x$	$+2x$
Equilibrium	$0.100 \text{ M} - x$	$0.100 \text{ M} - x$	$2x$

Step 2 Substitute: $49.0 = \frac{(2x)^2}{(0.100 \text{ M} - x)^2}$. Both sides are perfect squares, so take the square root: $7.00 = \frac{2x}{0.100 \text{ M} - x}$

Step 3 $0.700 \text{ M} - 7.00x = 2x$, so $x = \frac{0.700 \text{ M}}{9.00} = 0.0778 \text{ M}$

Step 4 $[\text{HI}] = 2x = 0.156 \text{ M}$; $[\text{H}_2] = [\text{I}_2] = 0.100 \text{ M} - 0.0778 \text{ M} = 0.0222 \text{ M}$

Step 5 Check: $\frac{(0.156 \text{ M})^2}{(0.0222 \text{ M})^2} = 49.4 \approx 49.0$ (the small difference comes from rounding).

Common mistake

 **COMMON MISTAKE:** Thinking equilibrium means equal amounts

At equilibrium the concentrations are **constant**, not equal. In the example above there is seven times as much HI as H_2 .

 **COMMON MISTAKE:** Forgetting the coefficients as powers

For $\text{H}_2 + \text{I}_2 \rightleftharpoons 2 \text{HI}$, $K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$, not $\frac{2[\text{HI}]}{[\text{H}_2][\text{I}_2]}$. Coefficients become **exponents**.

 **COMMON MISTAKE:** Including solids, liquids or water

Pure solids and liquids (including water as the solvent) do not appear in K_c . That's why K_a has no $[\text{H}_2\text{O}]$ term.

NOTATION NOTE

- The double half-arrow $\rightleftharpoons$ shows a reversible reaction at equilibrium.
- Square brackets mean concentration in mol/L: $[\text{HI}]$ = concentration of HI.
- “Position of equilibrium” describes whether reactants or products are favoured; it can change, but K changes only with temperature.

Remember this

REMEMBER THIS

- Dynamic equilibrium: forward rate = reverse rate; concentrations constant, not equal.
- K_c = products over reactants, each to the power of its coefficient; leave out solids and liquids.
- $Q < K$: forward. $Q > K$: reverse. $Q = K$: at equilibrium.
- Use an ICE table; K_c depends only on temperature.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

Practise this topic with flashcards and a quiz at chemistryclarity.com/chemistry/chemical-equilibrium/