

Le Chatelier's Principle

How do changes in concentration, pressure and temperature affect an equilibrium?

●●○ Intermediate

Equilibrium

Last reviewed 4 October 2026

Before you start: it helps to understand [Chemical Equilibrium and K_c](#).

What is it?

Le Chatelier's principle: if a system at equilibrium is disturbed by a change in concentration, pressure or temperature, the position of equilibrium shifts in the direction that **partly counteracts** the change.

“Shifts to the right” means more products form; “shifts to the left” means more reactants form. The system never fully undoes the change; it only reduces its effect.

🔑 KEY IDEA

Changes in concentration and pressure shift the equilibrium **without changing K**: they make $Q \neq K$, and the reaction runs until $Q = K$ again. A change in **temperature changes K itself**. A catalyst changes neither.

Why does it matter?

- **Industry.** Chemists choose temperatures and pressures that push equilibria toward the product, as in the Haber process for ammonia.
- **Everyday chemistry.** A fizzy drink goes flat when opened because lowering the pressure shifts $\text{CO}_2(\text{aq}) \rightleftharpoons \text{CO}_2(\text{g})$ toward the gas.
- **Living systems.** The body's buffers respond to added acid or base following the same principle.

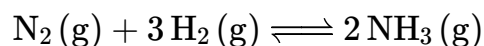
How does it work?

1. Changing a concentration

- **Add** a reactant (or remove a product): Q becomes smaller than K , so the equilibrium shifts **right** to use some of it up.
- **Add** a product (or remove a reactant): $Q > K$, so it shifts **left**.

2. Changing the pressure (gases)

Increasing the pressure by **reducing the volume** shifts the equilibrium toward the side with **fewer moles of gas**, which lowers the pressure:



The left side has 1 mol + 3 mol = 4 mol of gas; the right side has 2 mol.

Higher pressure shifts this equilibrium **right**, toward ammonia. If both sides have the same number of moles of gas (as in $\text{H}_2 + \text{I}_2 \rightleftharpoons 2 \text{HI}$), pressure has no effect on the position. Adding an inert gas at constant volume does not change any concentration, so it has no effect either.

3. Changing the temperature

Treat heat as a reactant or product:

- **Exothermic** forward reaction (ΔH negative): heat is a “product”. Raising the temperature shifts the equilibrium **left** and **decreases K**.
- **Endothermic** forward reaction (ΔH positive): heat is a “reactant”. Raising the temperature shifts it **right** and **increases K**.

For $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2 \text{NO}_2(\text{g})$, $\Delta H = +57 \text{ kJ}$: on heating, more brown NO_2 forms, so the colourless-to-brown mixture gets **darker**.

4. Adding a catalyst

A catalyst speeds up the forward and reverse reactions **equally**. Equilibrium is reached **sooner**, but its position and K are unchanged.

Change	Effect on position	Effect on K
Add reactant / remove product	shifts right	none
Add product / remove reactant	shifts left	none
Increase pressure (smaller volume)	toward fewer moles of gas	none
Increase temperature	in the endothermic direction	changes
Add a catalyst	none (equilibrium reached faster)	none

∞ THINK OF IT LIKE THIS

Think of a crowded room with two doors and people moving between it and the corridor at equal rates. Push more people into the room and, for a while, more people leave than enter, until the flows balance again. The room ends up a bit more crowded than before, but less crowded than right after the push: the change is partly counteracted.

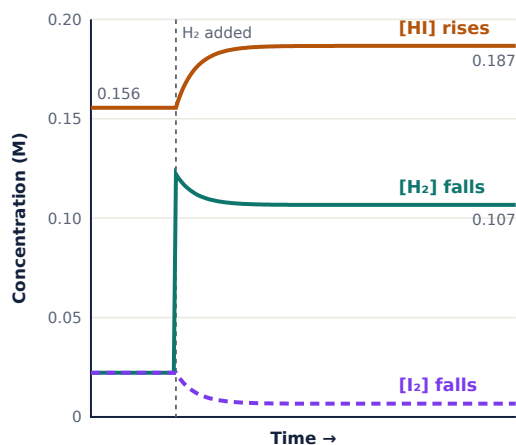
Q MORE PRECISELY

Le Chatelier's principle predicts the **direction** of a shift; comparing Q with K explains **why**, and an ICE table gives the **new concentrations**. How strongly K depends on temperature is described by the van 't Hoff equation, which links it to ΔH° .

Visualise it

Adding H_2 to $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$

0.100 M H_2 is added to the equilibrium mixture ($K_c = 49.0$).



The system uses up some of the added H_2 (and some I_2), making more HI , until $Q = K_c = 49.0$ again.

Adding H_2 shifts the equilibrium to the right: some of the added H_2 is used up and more HI forms. Calculated from a simple rate model.

Worked example

WORKED EXAMPLE: Predicting shifts in the Haber process

Question: For $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$, $\Delta H = -92\text{ kJ}$, predict the effect on the amount of ammonia of (a) adding N_2 (b) increasing the pressure (c) increasing the temperature (d) adding an iron catalyst.

- Step 1** Adding N_2 (a reactant): shifts **right**, so **more** NH_3 .
- Step 2** Higher pressure: 4 mol gas on the left, 2 mol on the right, so it shifts toward fewer moles: **right**, **more** NH_3 .
- Step 3** The forward reaction is exothermic, so heating shifts **left**: **less** NH_3 (and K decreases).
- Step 4** A catalyst does not shift the equilibrium: **no change** in the amount, but equilibrium is reached faster.

WORKED EXAMPLE: Checking a shift with Q

Question: An equilibrium mixture of $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$ ($K_c = 49.0$) contains 0.0222 M H_2 , 0.0222 M I_2 and 0.156 M HI . 0.100 M H_2 is added. Which way does the equilibrium shift?

- Step 1** Just after the addition, $[\text{H}_2] = 0.0222\text{ M} + 0.100\text{ M} = 0.122\text{ M}$
- Step 2**
$$Q = \frac{(0.156\text{ M})^2}{(0.122\text{ M})(0.0222\text{ M})} = 8.99 \text{ (the units of } \text{M}^2 \text{ cancel)}$$
- Step 3** $Q = 8.99 < K_c = 49.0$, so the reaction goes **forward**: H_2 and I_2 are used up and more HI forms, as Le Chatelier's principle predicts.

Common mistake

COMMON MISTAKE: Thinking concentration changes alter K

Adding a reactant shifts the **position**, but K stays the same. Only temperature changes K .

COMMON MISTAKE: Thinking a catalyst increases the yield

A catalyst makes equilibrium arrive faster but does not change how much product is present at equilibrium.

 COMMON MISTAKE: Counting all moles, not moles of gas

For pressure changes, count only **gases**. In $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$ there is 0 mol of gas on the left and 1 mol on the right.

 NOTATION NOTE

- “Shifts to the right / forward” = more products; “to the left / backward” = more reactants.
- ΔH refers to the forward reaction as written; the reverse reaction has the opposite sign.

Remember this

 REMEMBER THIS

- A disturbed equilibrium shifts to partly counteract the change.
- Add reactant or remove product: right. Higher pressure: fewer moles of gas.
- Heating favours the endothermic direction and is the only change that alters K .
- A catalyst gives equilibrium faster but does not move it.

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.

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