=Unit 5=
Chemical Equilibrium
(IB Text Ch 7)

7.1 Equilibrium

Chemical Equilibrium:

1. The state in which forward and reverse reactions occur at equal rates.
2. The state at which the concentrations of all reactants and products remain constant.

Example:

$$NaCl(s) \longrightarrow Na^+(aq) + Cl^-(aq)$$

When the solution becomes saturated, the reaction will begin to reverse itself:

$$NaCl(s) \leftrightarrows Na^+(aq) + Cl^-(aq)$$

Both forward and reverse reactions occur simultaneously <u>and</u> continuously.

(Called a "dynamic" process.)

When
forward rate = reverse rate,
a <u>state of equilibrium</u>
is said to have been reached.

Consider the following reaction:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

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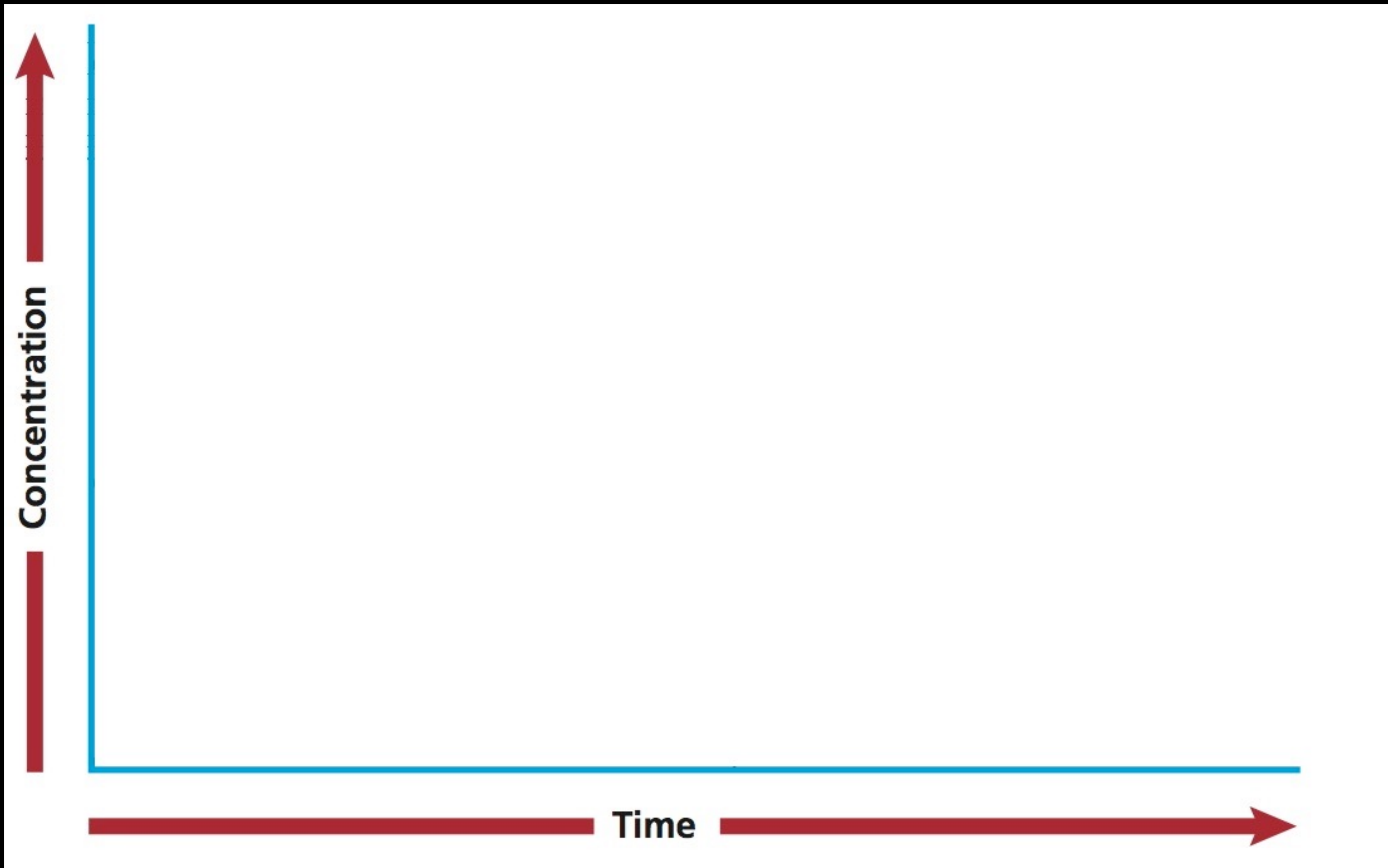

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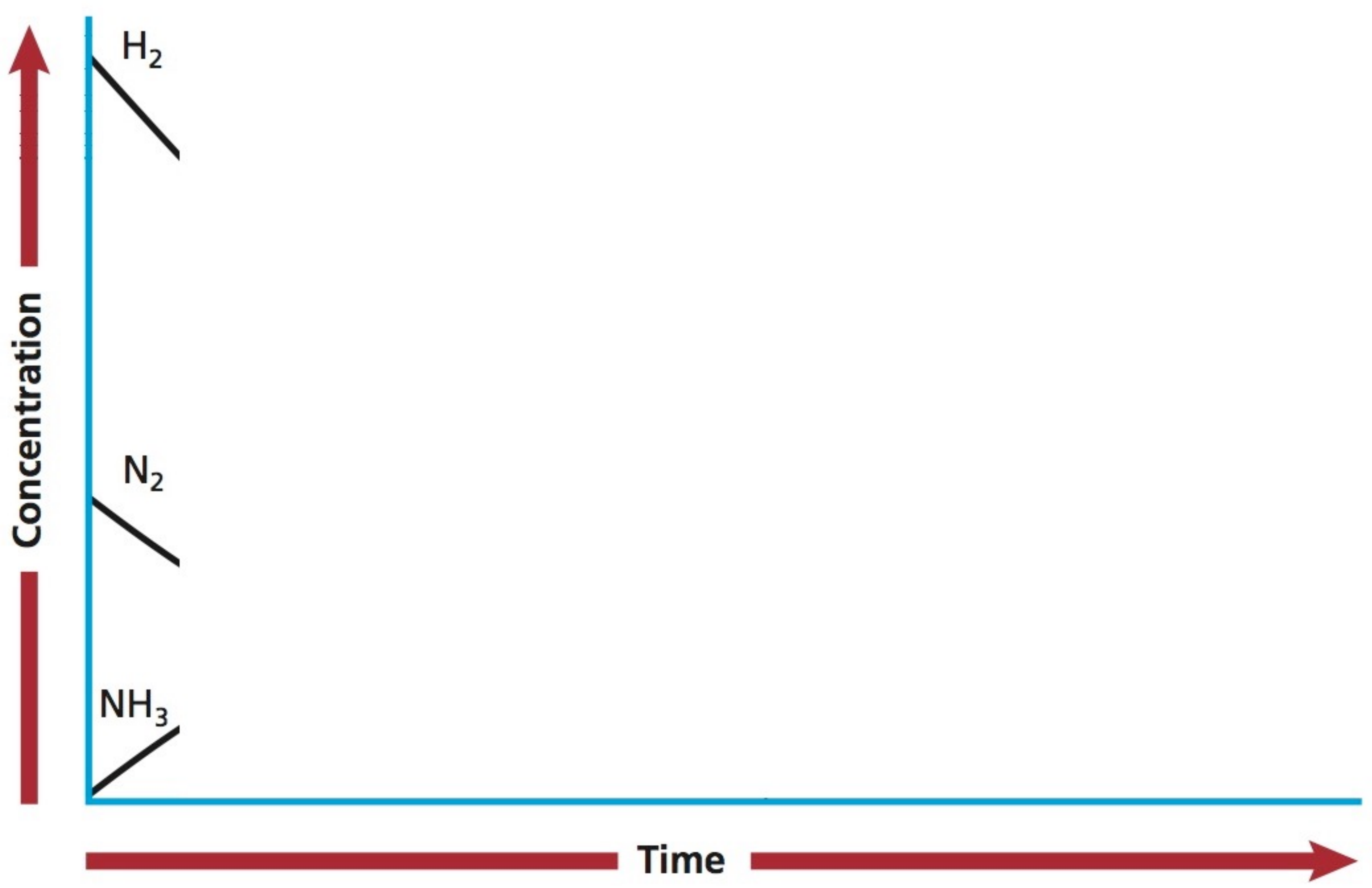

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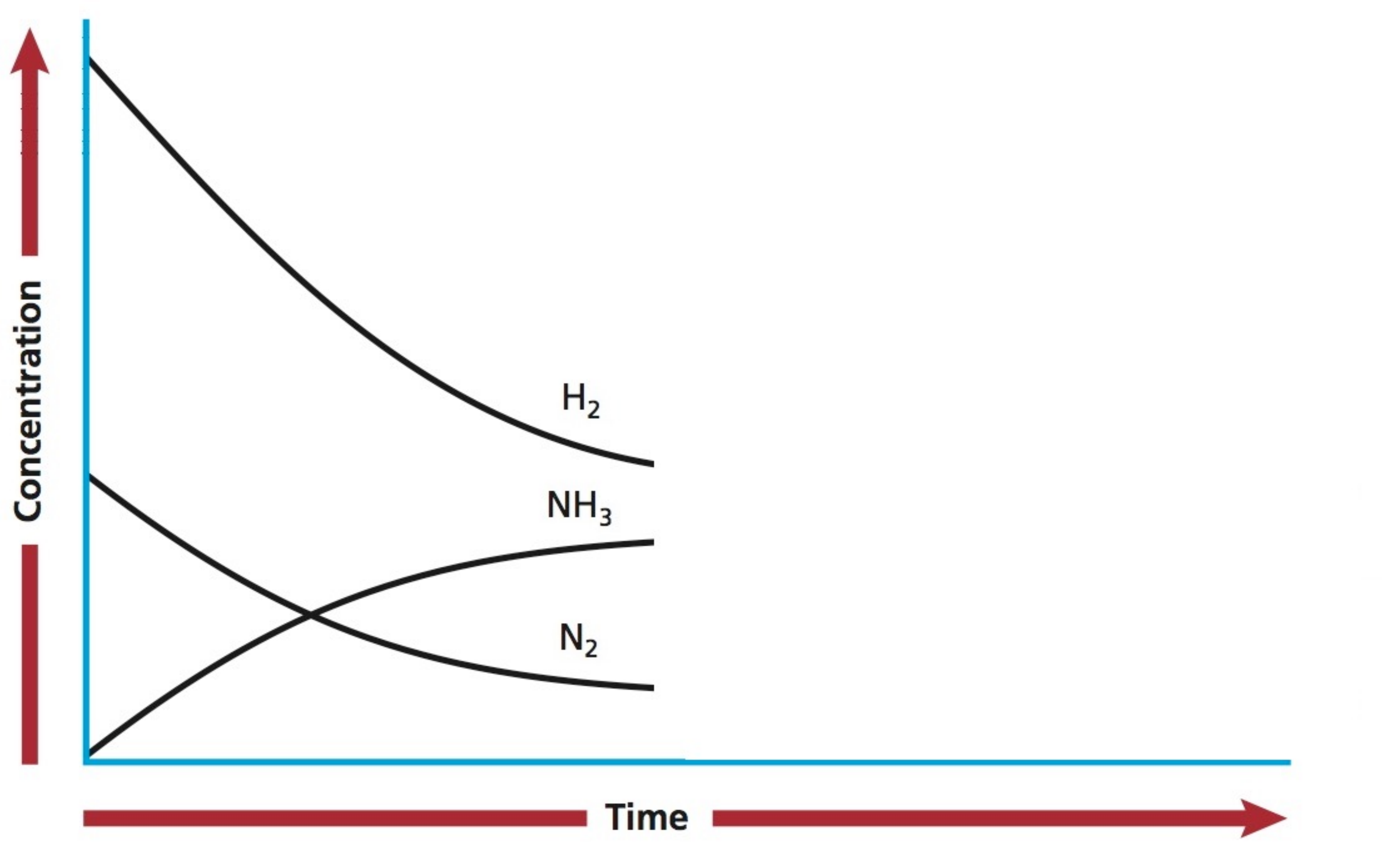

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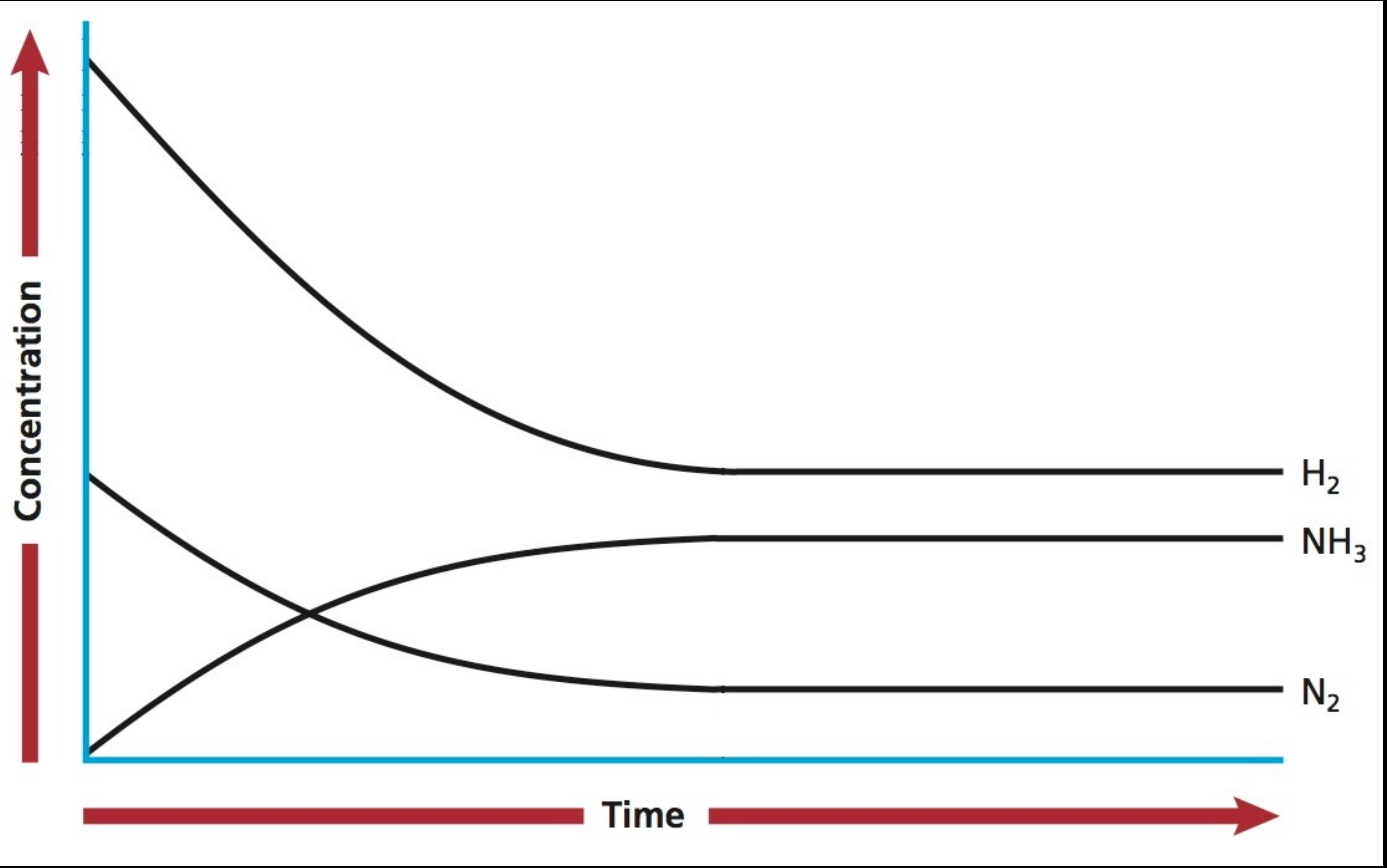

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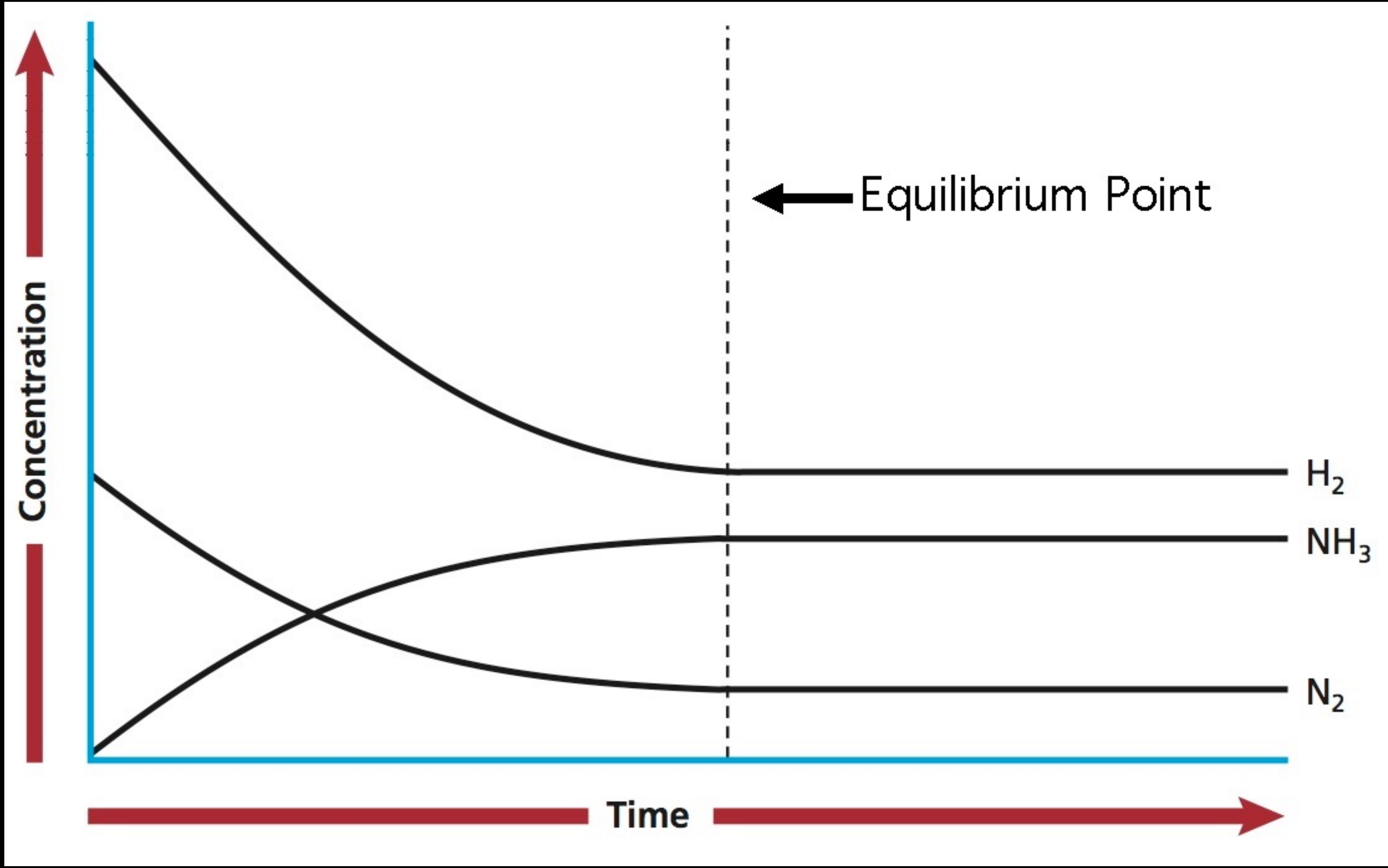


The Equilibrium Constant

The Law of Equilibrium

If at a given temperature,

$$aA + bB \leftrightharpoons cC + dD$$

Then,

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

The equilibrium constant

The value of K_c will be constant for any specified temperature.

The value of K_c :

If $K_c > 1$, rxn is product favored

If $K_c < 1$, rxn is reactant favored

If $K_c = 1$, neither reactant nor product favored

The magnitude of K_c :

If $K_c >> 1$, rxn goes to completion

(of the order $K_c > 1000$)

If $K_c << 1$, rxn does not take place

(of the order $K_c < 0.001$)

Phases and K_c

K_c is independent of solids and liquids, as their concentrations do not change. Therefore, solids and liquids are not included in the equilibrium expression.

Homogeneous equilibria: one phase present for all reactants and products.

Heterogeneous equilibria: two or more phases present.

The Position of Equilibrium

Each set of
equilibrium concentrations
is called an...

EQUILIBRIUM POSITION

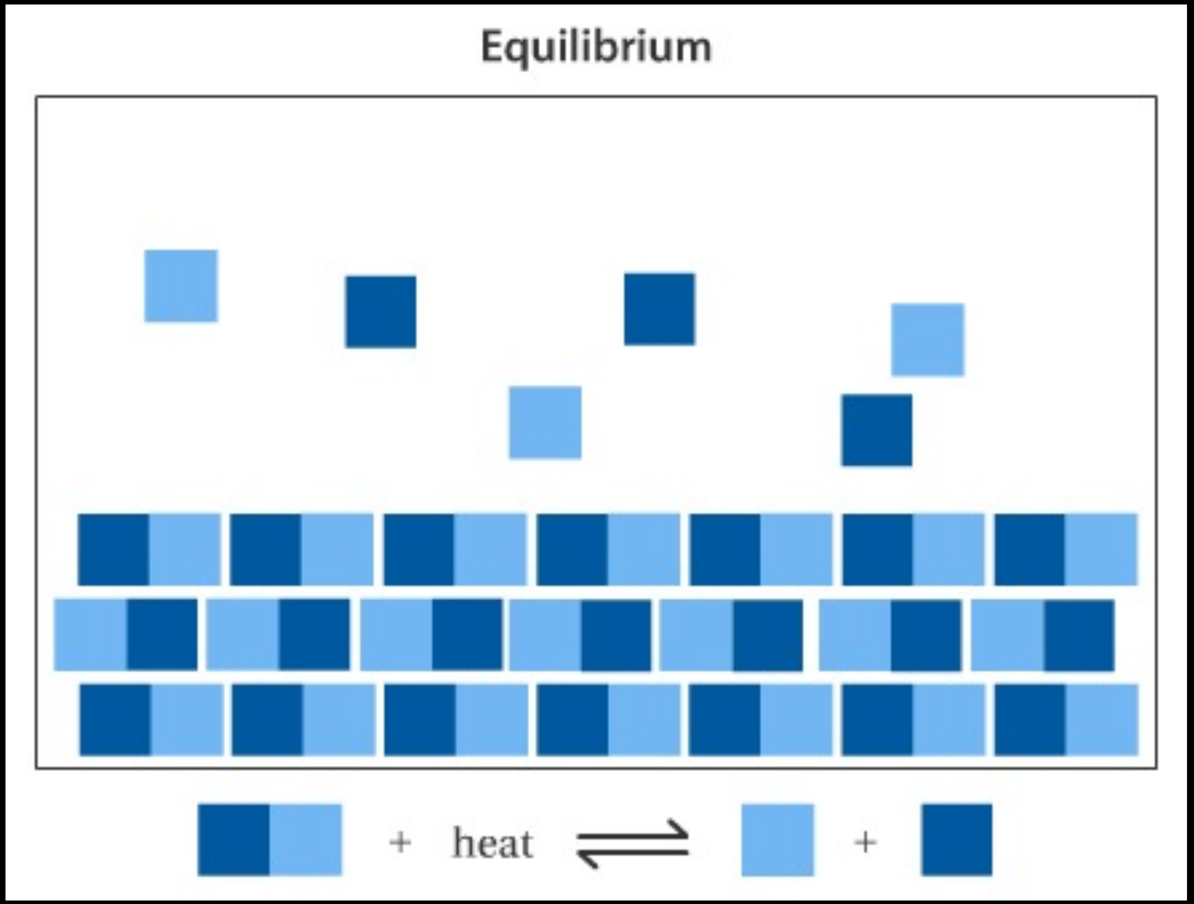
For any reversible reaction at a given temperature, there is only one value for $K_{c,}$

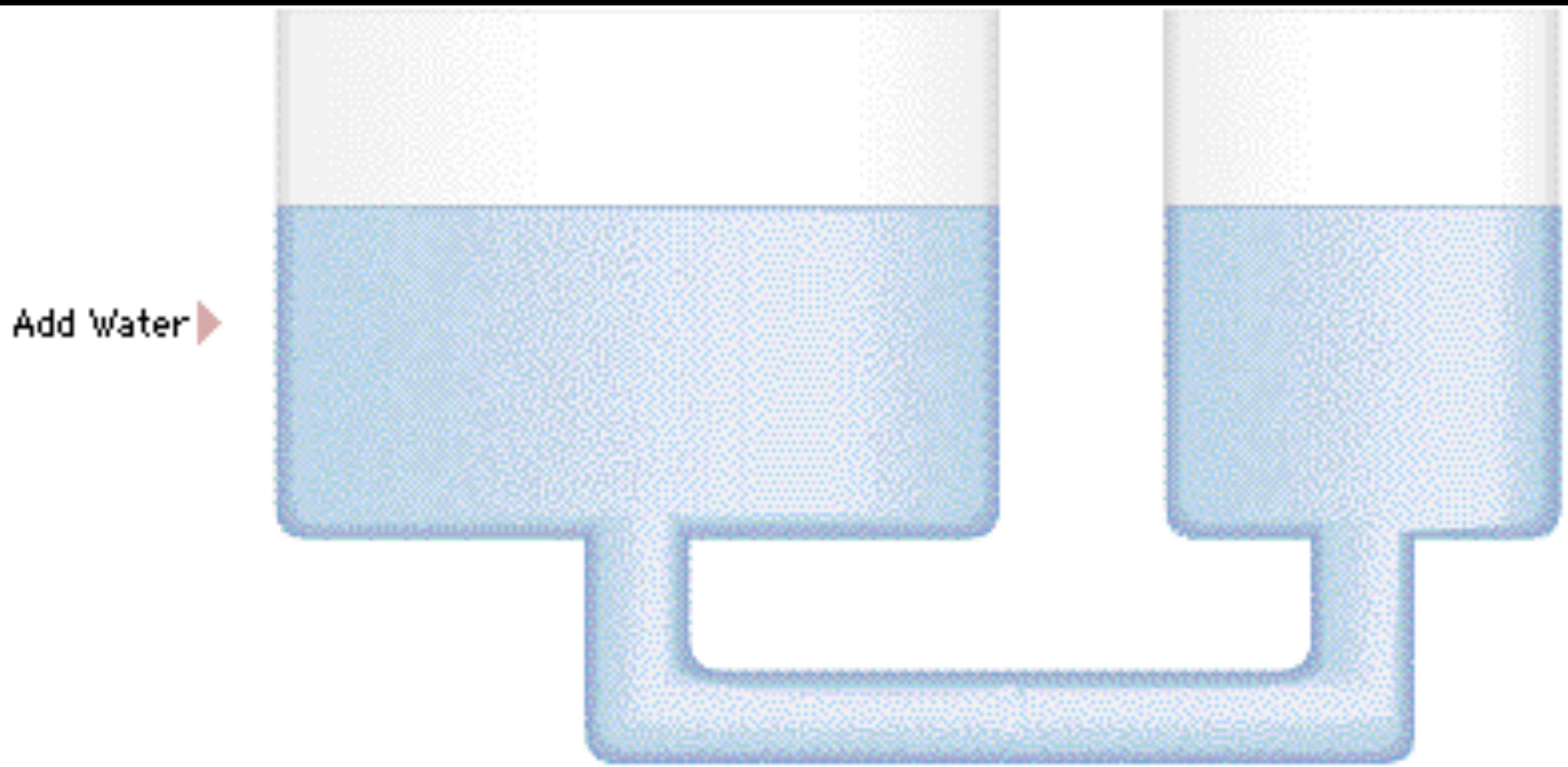
but an infinite number of possible equilibrium positions!

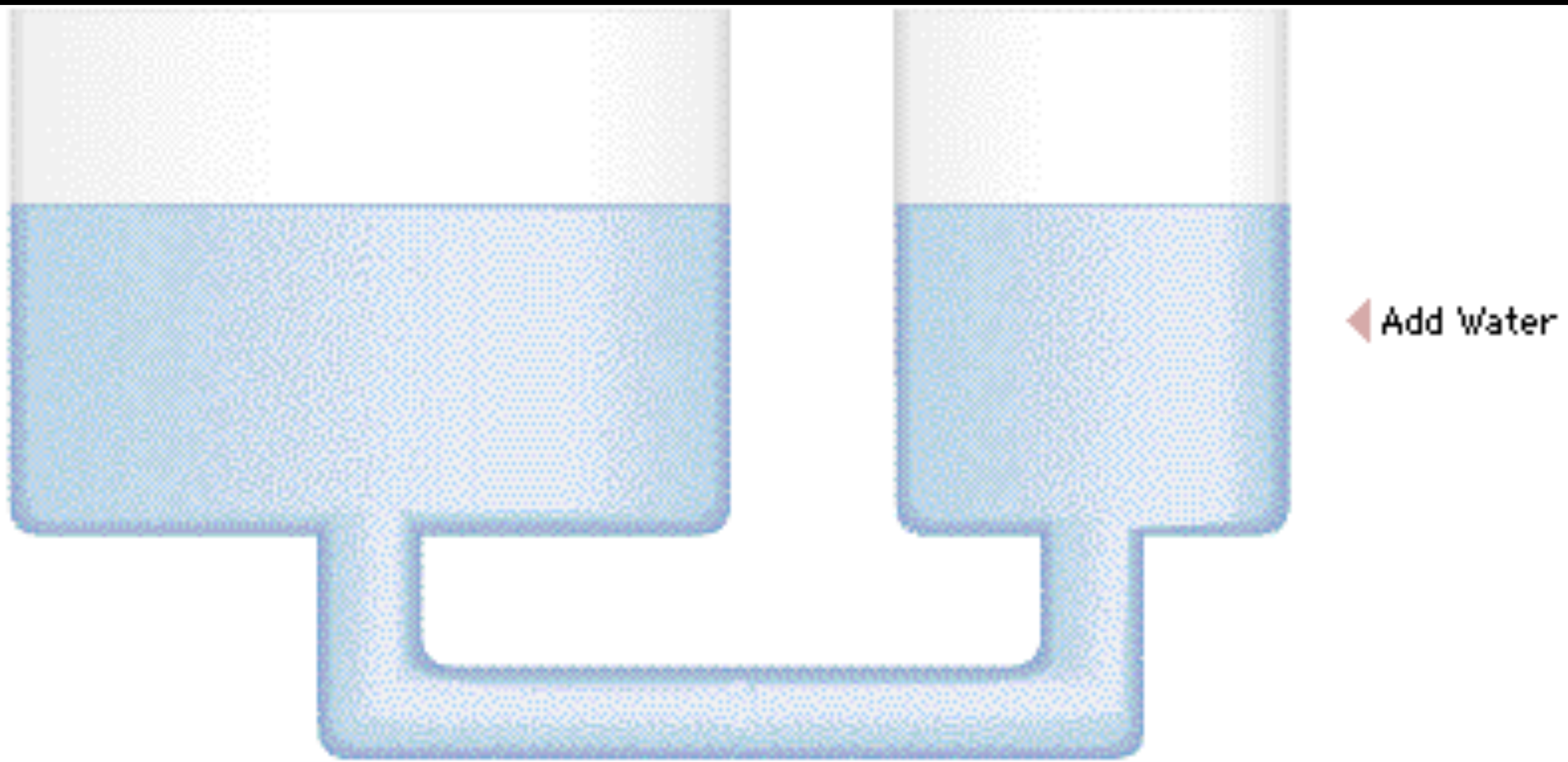
[practice problems done in class]

Le Châtelier's Principle:

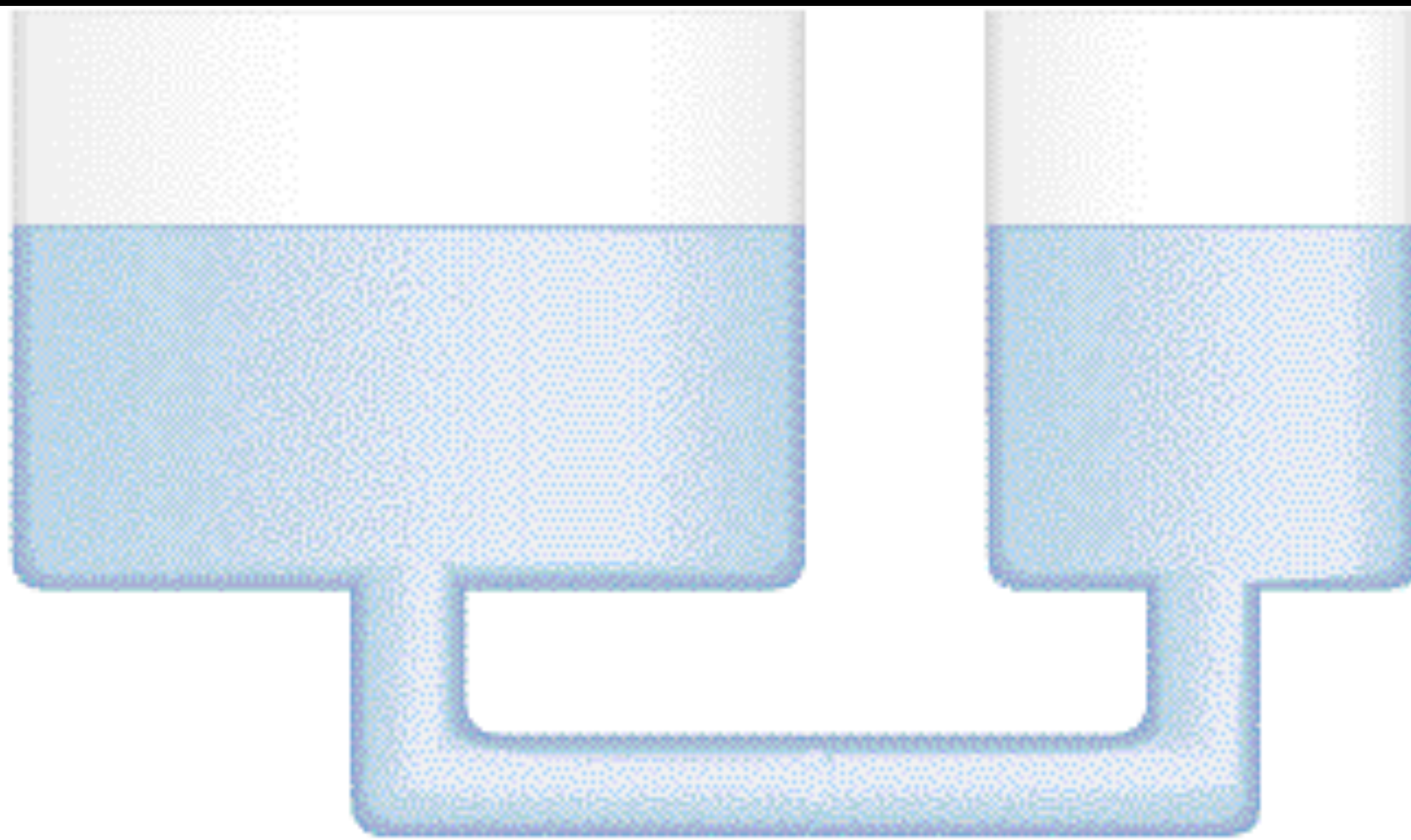
If a change is imposed on a system at equilibrium, the reaction will "shift" in the direction that tends to reduce that change.

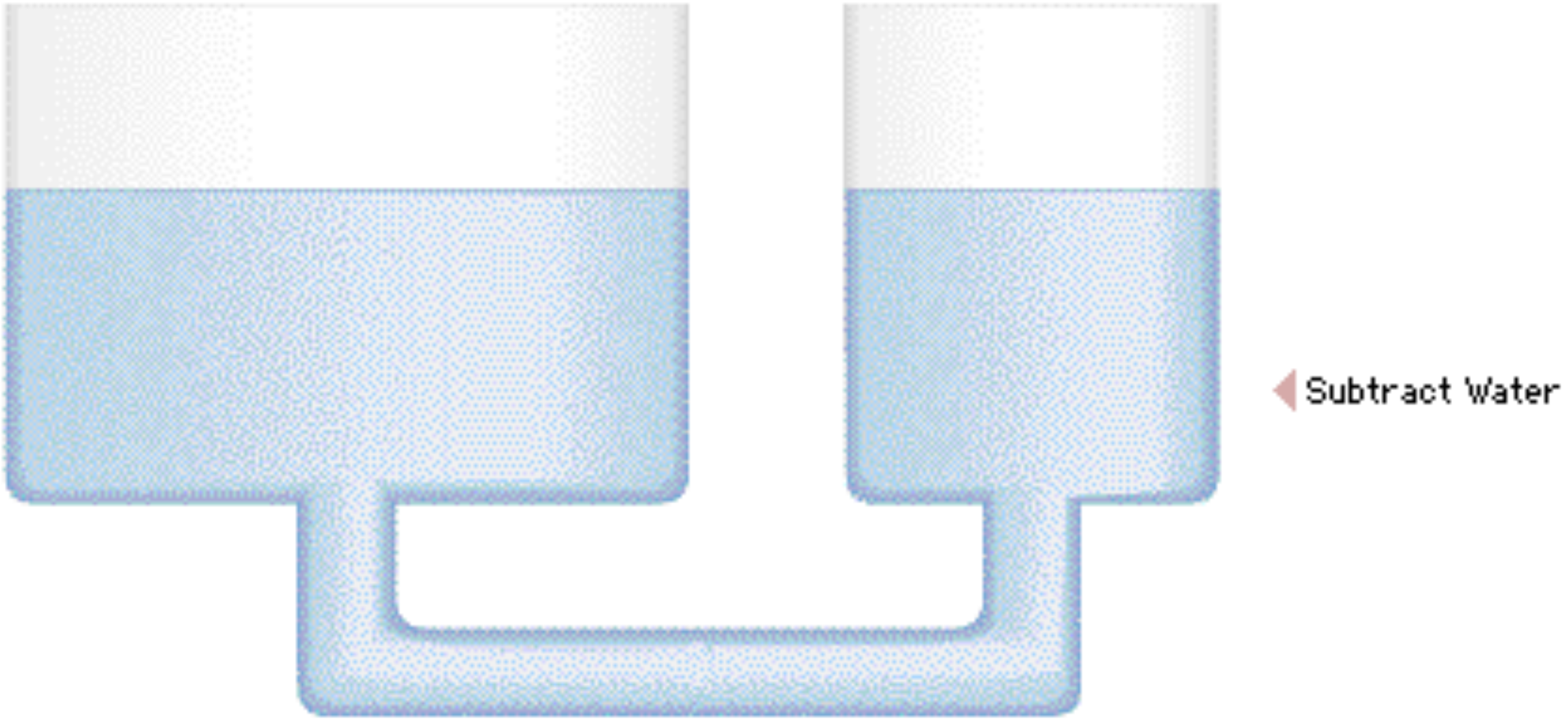



Add Water


Add Water

Subtract Water




Subtract Water

The primary factors that can shift the equilibrium of a reaction are:

- Concentration
- Temperature
- Pressure (gases only)

When increasing the pressure of a gaseous equilibrium system, the reaction will shift towards side with the least moles of gas...

thereby decreasing the overall pressure.

Hints for using La Chatelier's principle

- Treat heat as you would any other reactant or product.
- Unless the system is "closed," assume gases will escape.
- A catalyst has no effect on equilibrium.

The Reaction Quotient, Q

Used to determine how the initial concentrations of a reversible reaction will shift as equilibrium is reached

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

Three possibilities:

$K_c > Q$, (meaning: [products] > $[\text{products}]_i$)

Therefore to reach equilibrium, the reaction must shift to the right.

$K_c \dashrightarrow Q$

Three possibilities:

$K_c < Q$,   (meaning: [products] < [products]$_i$)

Therefore to reach equilibrium, the reaction must shift to the left.

$K_c \leftarrow Q$

Three possibilities:

$K_c = Q$, (meaning: [products] = [products]$_i$)

Therefore the reaction is already at equilibrium (no shift).

Industrial processes: Equilibrium and kinetics

IB
CHEMISTRY
VIDEO
REVIEW
COMPLETE

Industrial applications:

Process	Catalyst	Product	Product uses
Haber	Fe	ammonia	to make fertilizers, plastics, drugs, explosives...
Contact	V_2O_5	sulfuric acid	to make explosives, drugs, other chemicals... ; the electrolyte in the lead-acid storage (car) batteries

Industrial applications:

Process	Catalyst	Product	Product uses
Haber	Fe	ammonia	to make fertilizers, plastics, drugs, explosives...
Contact	V_2O_5	sulfuric acid	The most produced chemical in the world... It's production is a strong indicator of the strength of a country's economy.

Optimum conditions: Haber Process

Pressure	Temperature
200 atm	450 ℃
Shifts the reaction to the right.	Though the reaction is exothermic, high temperatures are needed increase the rate (to break the N_2 triple bond).

Optimum conditions: Contact process

Pressure	Temperature
2 atm	450 ℃
Results in a high enough yield, so higher (both dangerous & expensive) pressures are not needed.	Reaction is exothermic and low temperatures would increase yield, but high temperatures are needed to increases the rate of reaction.

The Common Ion Effect

The phenomenon in which the addition of an ion common to two solutes brings about a shift in an equilibrium system.