# Reversible Reactions

# Dynamic Equilibrium

# Macroconcept: Systems

# Reversible Reactions – Dynamic Equilibrium

What is a *system*?

- Systems are composed of sub-systems which *interact*.
- Systems follow *rules*.

# Reversible Reactions – Dynamic Equilibrium

A reversible reaction can be considered as a *system* in which changes are taking place. The system is composed of reactants and products that interact with each other following a specific set of *rules*. These rules (set out in Le Chatelier's Principle) allow us to *predict* how the system will respond to changes in concentration, pressure, temperature and the addition of a catalyst. This has important applications in industrial chemistry and the chemistry of biological systems.

# Reversible Reactions – Dynamic Equilibrium

**Guiding questions:**

- What is dynamic equilibrium?
- Why are some reactions reversible while others are not?
- Why can dynamic equilibrium only be reached in a closed system?
- What is happening at the macroscopic (observable) level at equilibrium?
- What is happening at the microscopic (molecular) level at equilibrium?
- How (and why) can changes in variables such as temperature and pressure affect a system that has reached dynamic equilibrium?
- How might the world be different if all reactions were reversible, or if no reactions were reversible?
- How is Le Chatelier’s Principle applied to improve the efficiency of reactions carried out on an industrial scale?

# Reversible Reactions – Dynamic Equilibrium

What do I need to know about *reversible reactions*?

# Reversible Reactions – Dynamic Equilibrium

1. State that some chemical reactions are reversible (*e.g.* manufacture of ammonia).
2. Describe the idea that some chemical reactions can be reversed by changing the reaction conditions.
3. Explain, in terms of rates of the forward and reverse reactions, what is meant by a reversible reaction and dynamic equilibrium.
4. Use Le Chatelier’s principle to deduce qualitatively (from appropriate information) the effects of changes in temperature, concentration, pressure or presence of a catalyst on a system at equilibrium and explain the effects of the changes in terms of rate of reaction.
5. Explain the conditions used in the Haber process, as an example of the importance of an understanding of chemical equilibrium in the chemical industry and the application of Le Chatelier’s principle.

# Reversible Reactions – Dynamic Equilibrium

What is special about *reversible reactions*?

# Reversible Reactions – Dynamic Equilibrium

- **Reversible reaction:** A reversible reaction is a chemical reaction in which the conversion of reactants to products and the conversion of products to reactants occur simultaneously under suitable conditions.

- For example, ethanoic acid is a *weak acid* because its ionisation in water is *reversible*:

$$\text{ethanoic acid} \rightleftharpoons \text{ethanoate ion} + \text{hydrogen ion}$$

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

- Ethanoic acid is a weak acid because it is never fully ionised in aqueous solution. This is because some of the ethanoate ions and some of the hydrogen ions always recombine to form molecules of ethanoic acid.

# Reversible Reactions – Dynamic Equilibrium

- Instead of using a single-headed arrow (→) showing the reaction only proceeds in one direction, the notation for a reversible reaction is ⇌.

- The *forward* reaction is the reaction that goes from *left-to-right*.

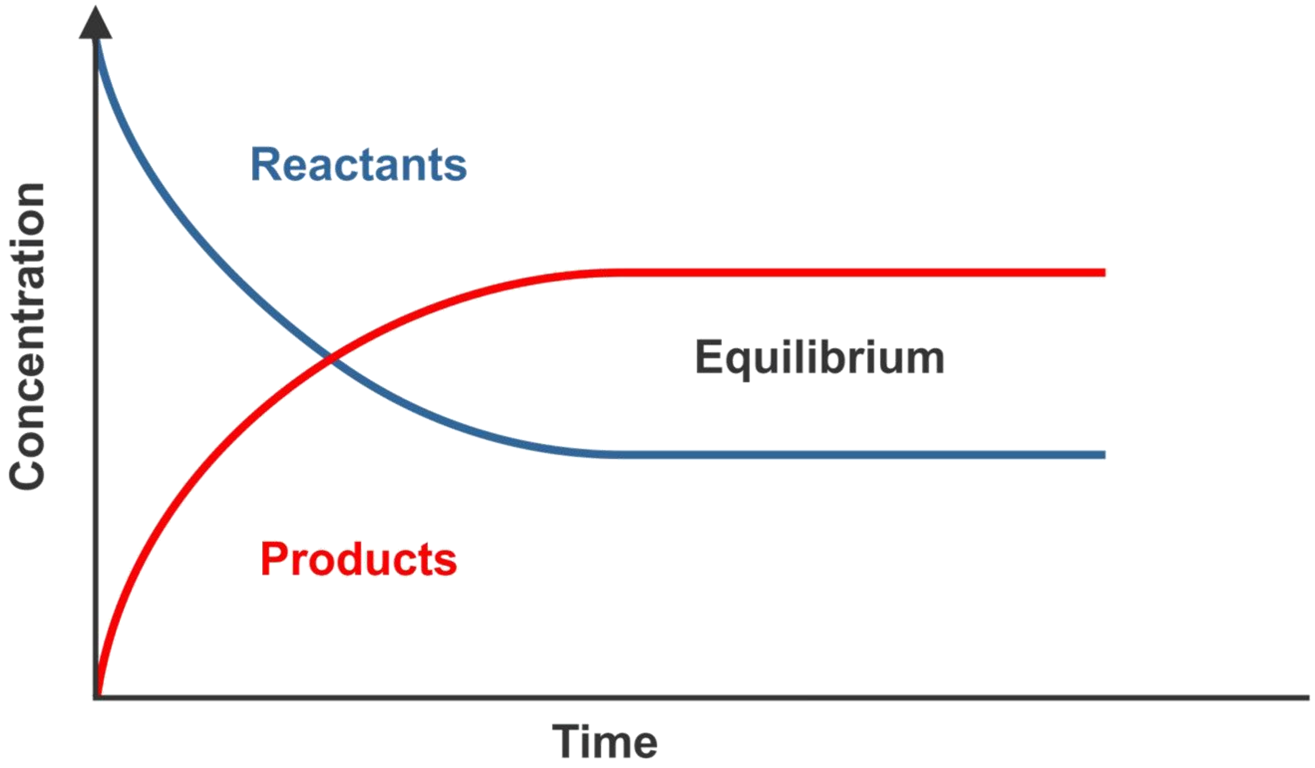
- The *reverse* reaction is the reaction that goes from *right-to-left*.

→ forward reaction →

$$\rightleftharpoons$$

← backward reaction ←

# Reversible Reactions – Dynamic Equilibrium



- **Dynamic equilibrium:** Dynamic equilibrium occurs in reversible reactions when the *rate* of the forward reaction has become equal to the *rate* of the backward reaction. As a result, there is *no change* in the amount / concentration of the reactants and products.

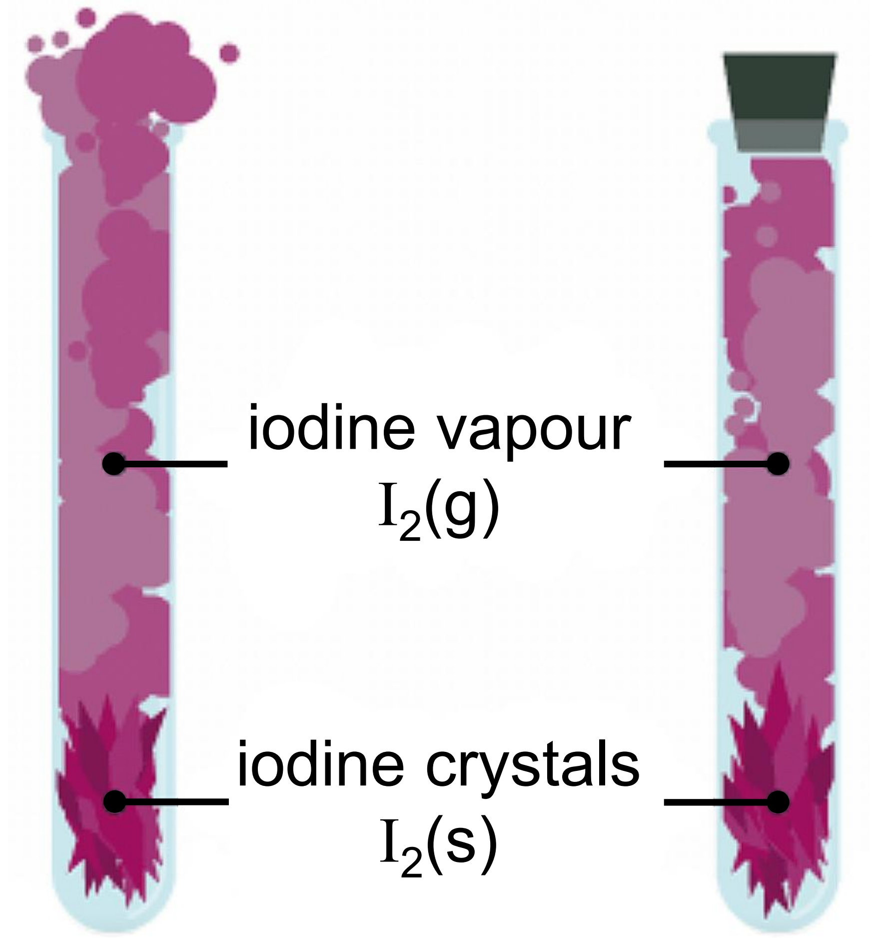
# Reversible Reactions – Dynamic Equilibrium

- Dynamic equilibrium can only be reached in a *closed system* because, in such a system, there is no loss or gain of chemicals to or from the surroundings. This allows the forward and reverse reactions to occur at the same rate, maintaining a constant concentration of reactants and products. In an *open system*, chemicals could escape or enter, disrupting the balance needed for dynamic equilibrium to be established.

# Reversible Reactions – Dynamic Equilibrium

## Open System

Dynamic equilibrium *cannot* be established because the iodine vapour escapes from the system.



$I_2(s) \rightarrow I_2(g)$

$I_2(s) \rightleftharpoons I_2(g)$

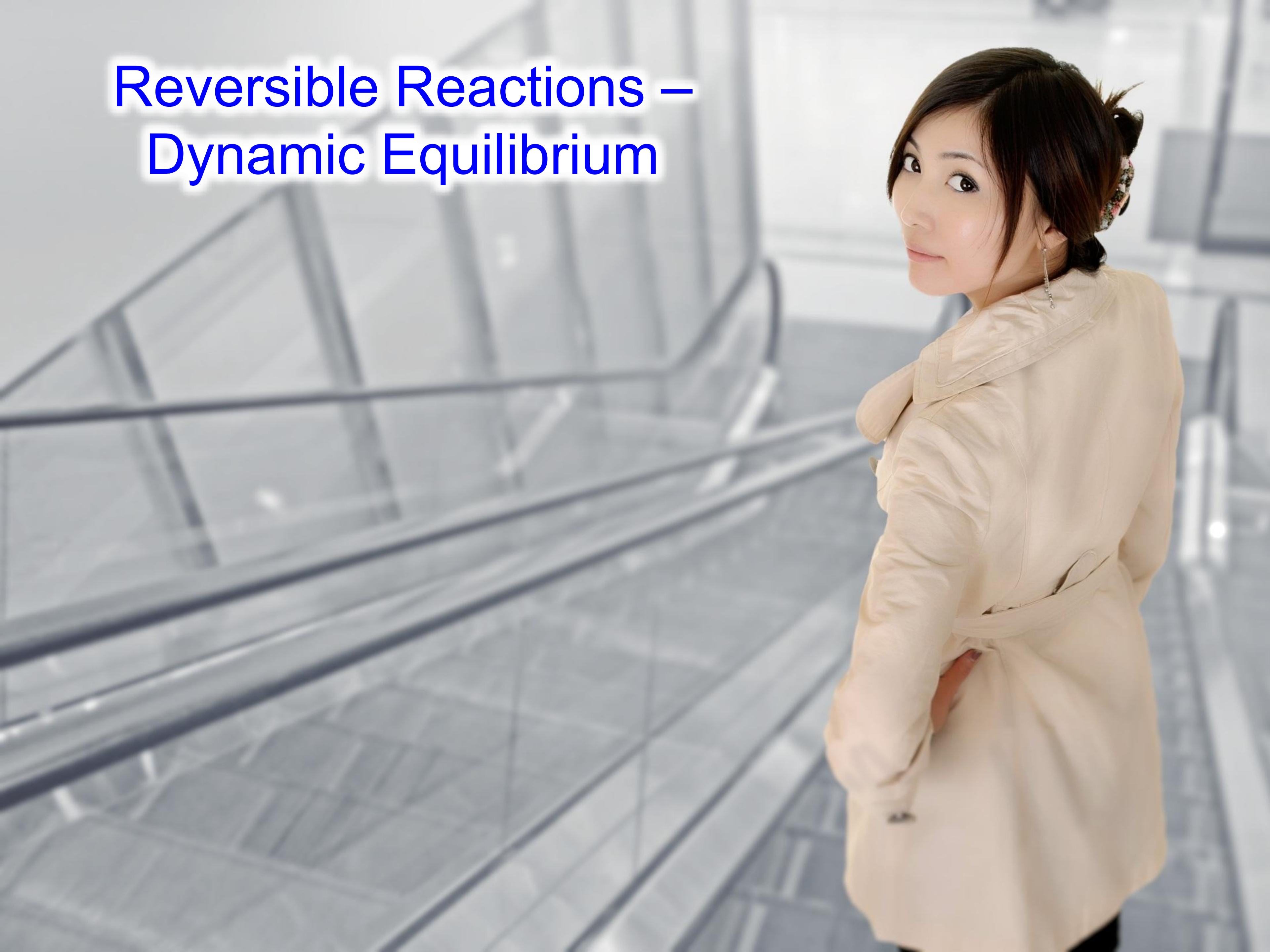
## Closed System

Dynamic equilibrium *can* be established because the iodine vapour remains in the system.

# Reversible Reactions – Dynamic Equilibrium

Could I please have an *analogy* for reversible reactions?

# Reversible Reactions – Dynamic Equilibrium



# Reversible Reactions – Dynamic Equilibrium

- Analogy for a reversible reaction.
- Imagine that you stand at the bottom of escalator and start to move upwards.
- For a chemical reaction, this is the same as the reactants colliding to form the products:

$$A + B \rightarrow C + D$$

# Reversible Reactions – Dynamic Equilibrium

- Analogy for a reversible reaction.
- Now imagine that you turn around and start to *run* back *down* the escalator while it is still *slowly* moving *upwards*.
- This is the same as the products colliding to form the reactants again:

$$A + B \leftarrow C + D$$

# Reversible Reactions – Dynamic Equilibrium

- Analogy for a reversible reaction.
- Now imagine walking back down the escalator at the *same rate* that the escalator is moving upwards.
- You are now at *equilibrium* on the escalator. It appears that you are standing still, although the escalator is still moving, and you are still walking.

# Reversible Reactions – Dynamic Equilibrium

- Analogy for a reversible reaction.
  - A chemical reaction reaches *equilibrium* when the rate of the forward reaction equals the rate of the backward reaction. It *appears* that the reaction has stopped, although chemical changes are still taking place:

$$A + B \rightleftharpoons C + D$$

# Reversible Reactions – Dynamic Equilibrium

- Analogy for a reversible reaction.
- You could reach your *equilibrium position* near the top or bottom of the escalator, not only around the middle.
- Likewise, at equilibrium, the concentrations of reactants and products remain constant *but are not necessarily equal*.

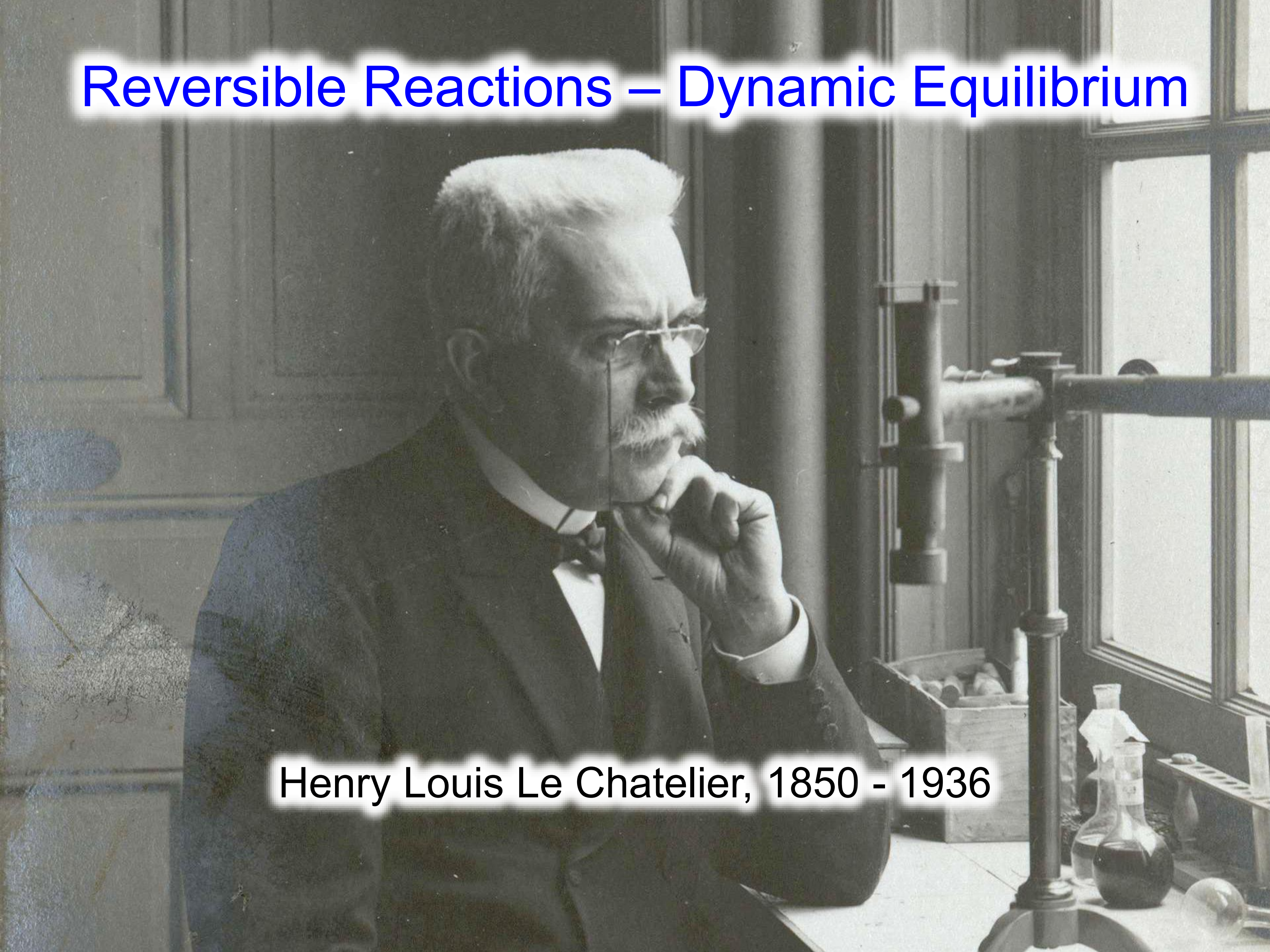
$$A + B \rightleftharpoons C + D$$

*i.e.* $[A] + [B] \neq [C] + [D]$

# Reversible Reactions – Dynamic Equilibrium

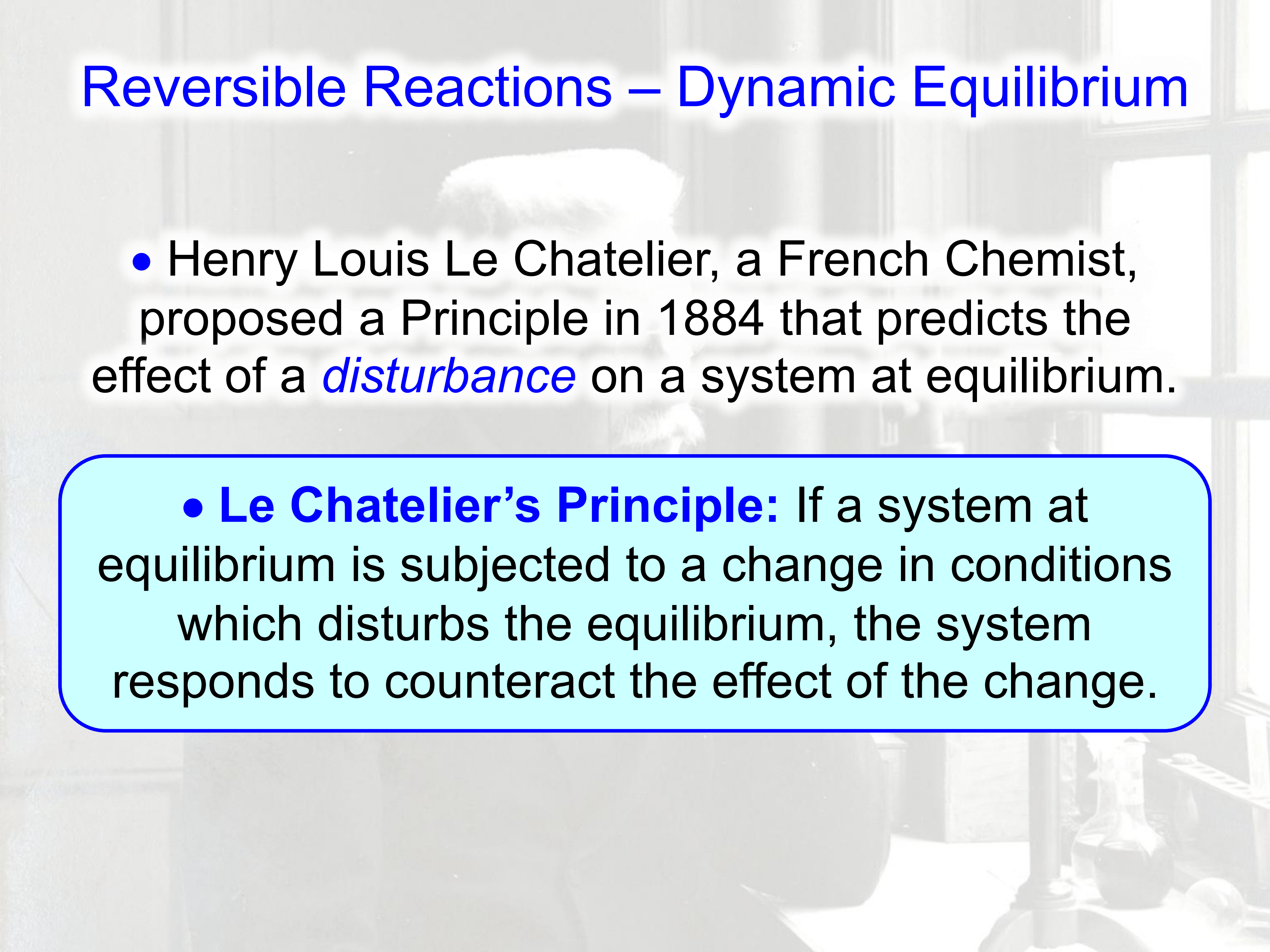
What is
*Le Chatelier's*
*Principle*?

# Reversible Reactions – Dynamic Equilibrium



Henry Louis Le Chatelier, 1850 - 1936

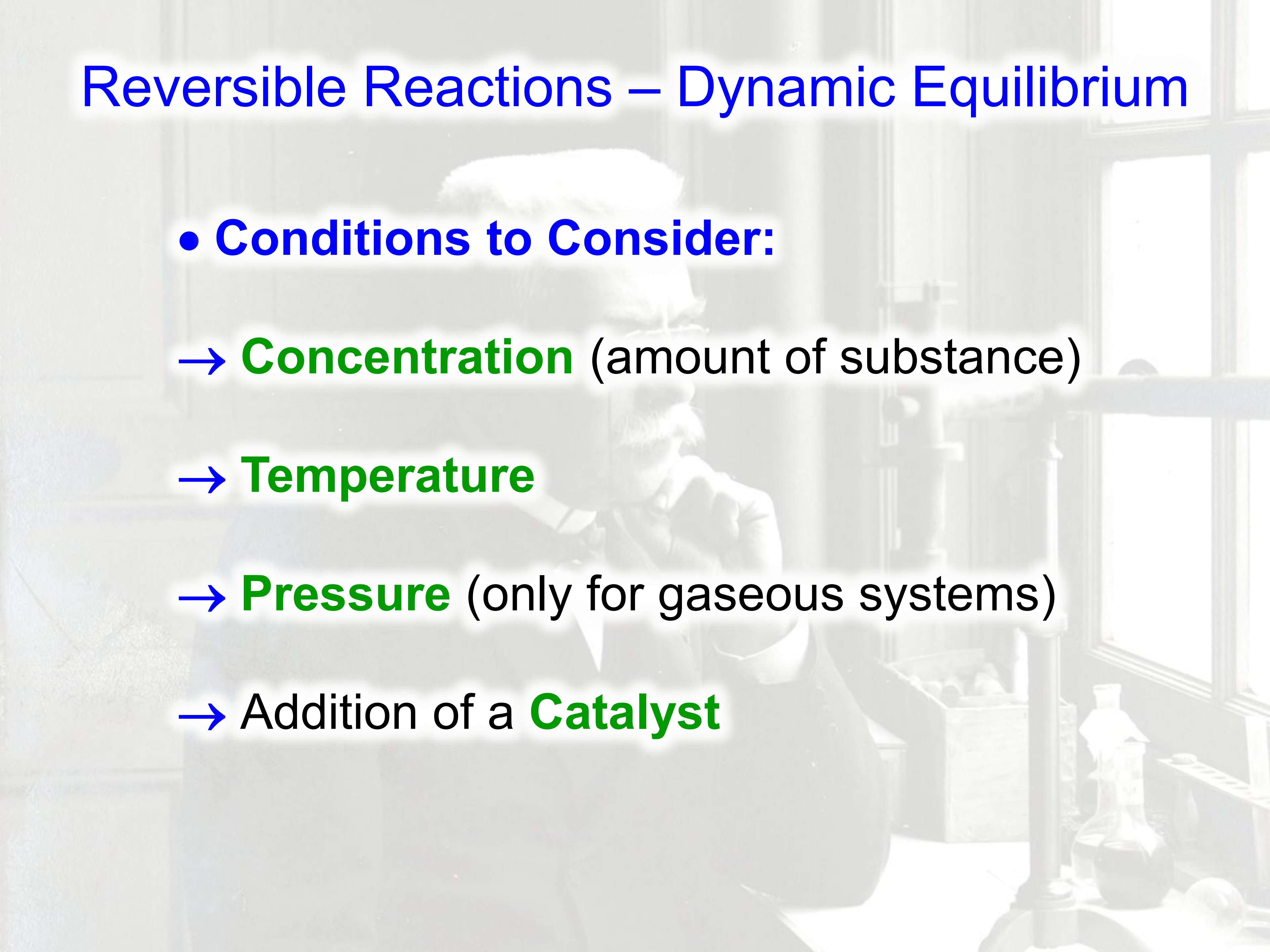
# Reversible Reactions – Dynamic Equilibrium



- Henry Louis Le Chatelier, a French Chemist, proposed a Principle in 1884 that predicts the effect of a *disturbance* on a system at equilibrium.

- **Le Chatelier's Principle:** If a system at equilibrium is subjected to a change in conditions which disturbs the equilibrium, the system responds to counteract the effect of the change.

# Reversible Reactions – Dynamic Equilibrium



- **Conditions to Consider:**

→ **Concentration** (amount of substance)

→ **Temperature**

→ **Pressure** (only for gaseous systems)

→ Addition of a **Catalyst**

# Reversible Reactions – Dynamic Equilibrium

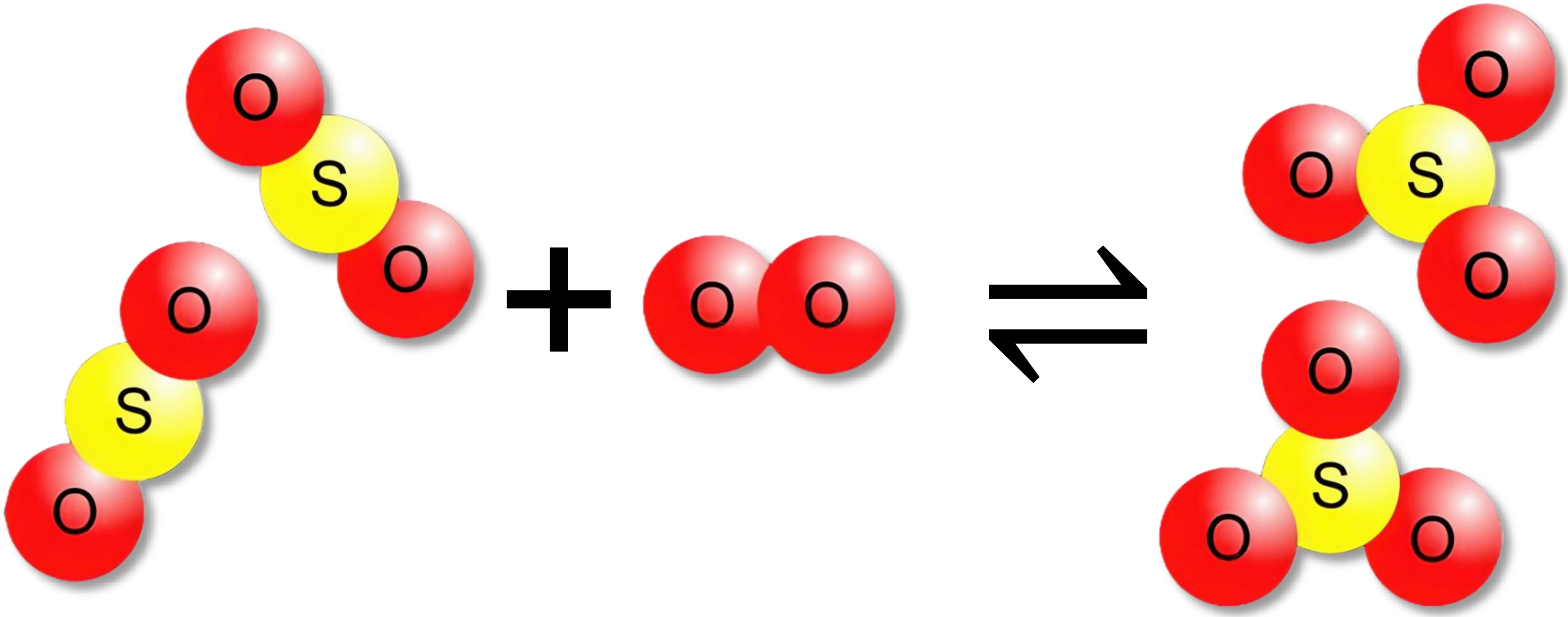
Can I please see an *example* of Le Chatelier’s Principle?

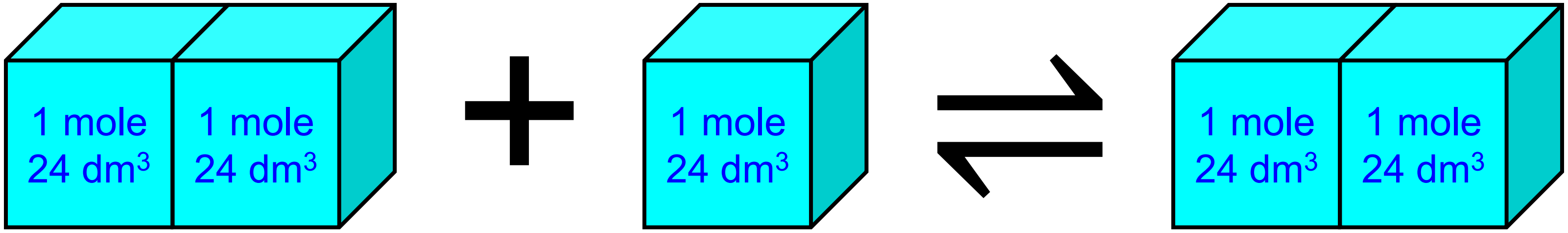
- Consider the following equilibrium mixture:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

# Reversible Reactions – Dynamic Equilibrium

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ $\Delta H = -196\ \text{kJ mol}^{-1}$

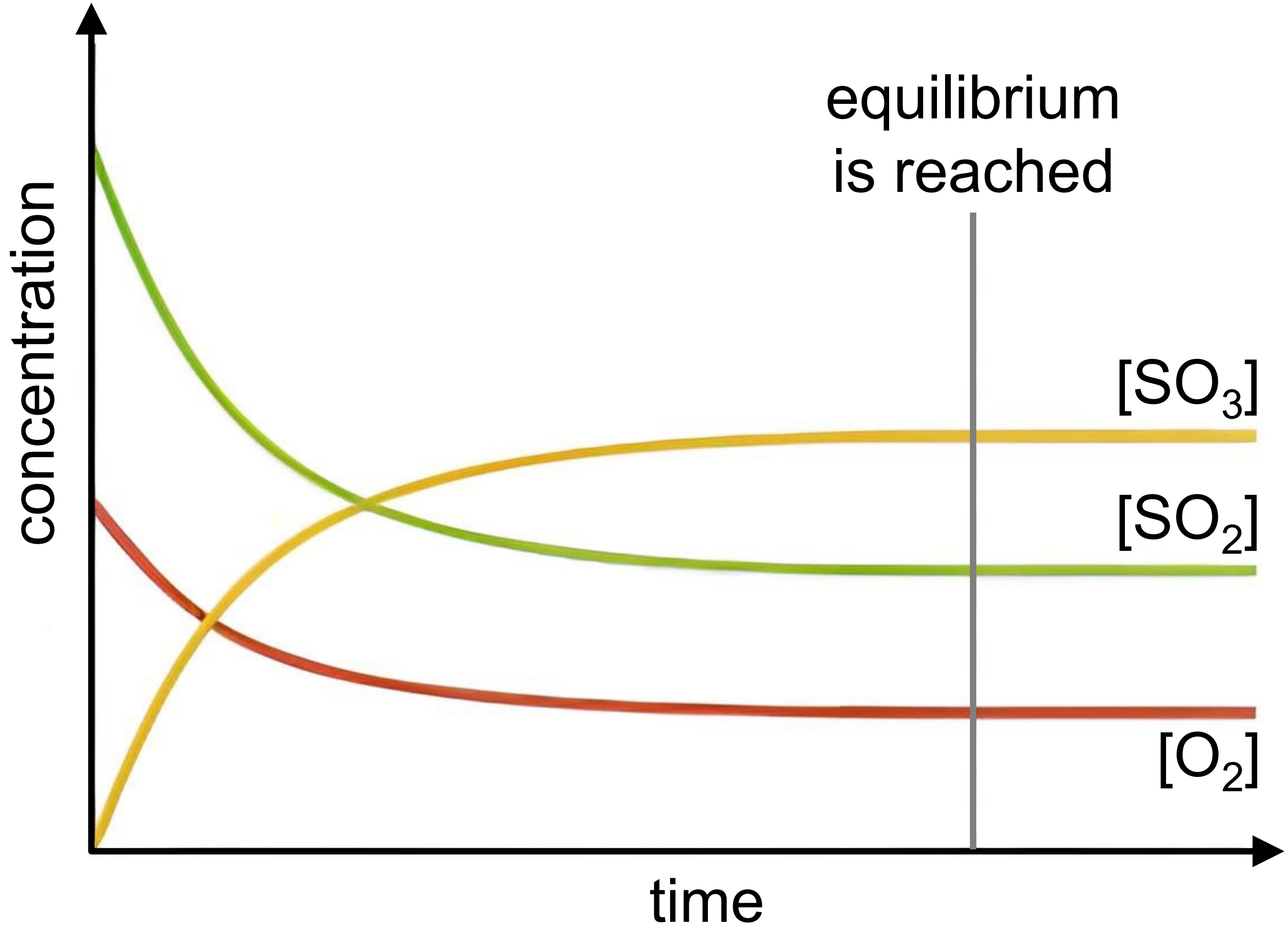




1 mole $24\ dm^3$ | 1 mole $24\ dm^3$ + 1 mole $24\ dm^3$ $\rightleftharpoons$ 1 mole $24\ dm^3$ | 1 mole $24\ dm^3$

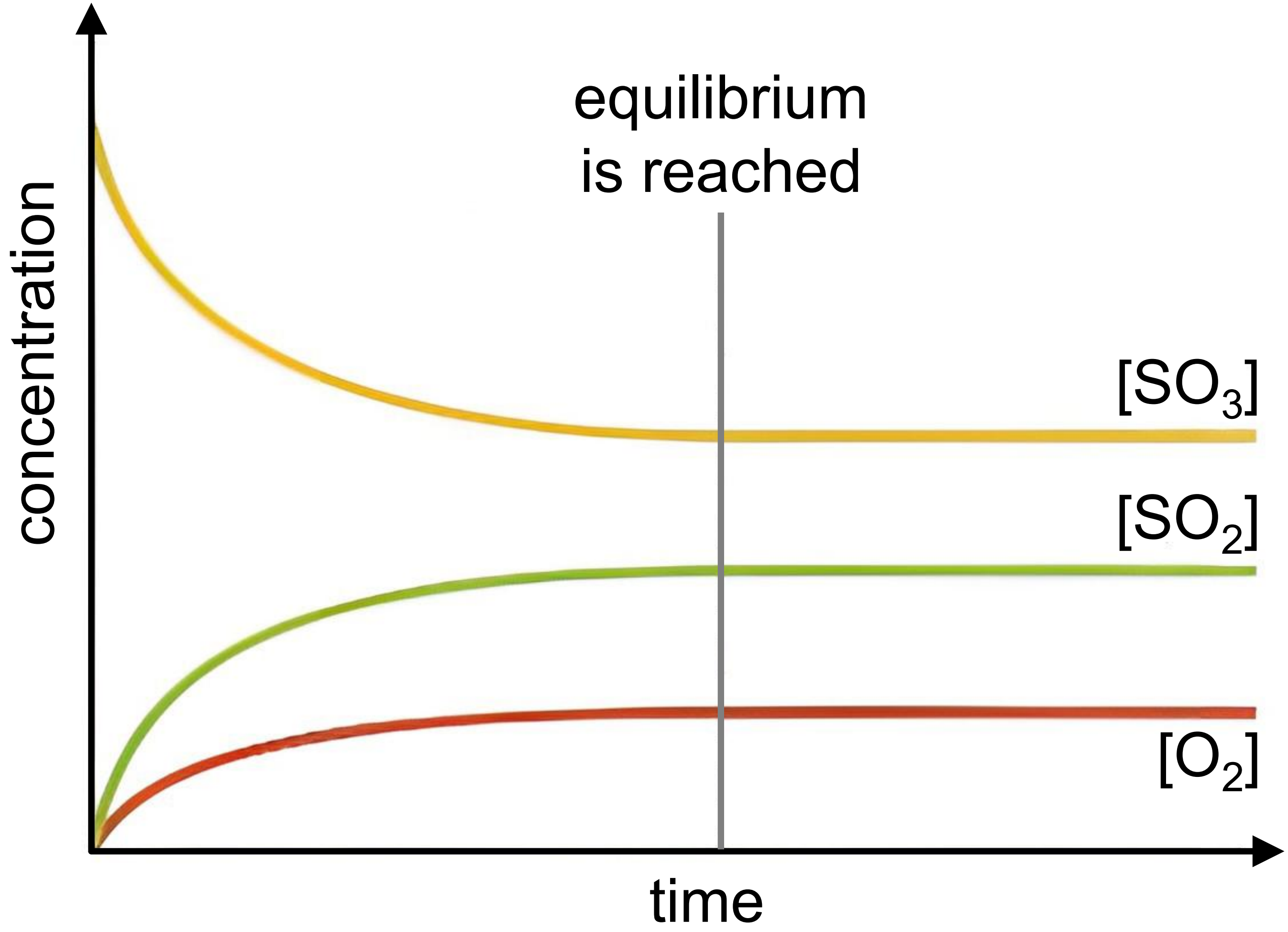
# Reversible Reactions – Dynamic Equilibrium

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ $\Delta H = -196\ \text{kJ mol}^{-1}$



# Reversible Reactions – Dynamic Equilibrium

$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$ $\Delta H = +196\ kJ\ mol^{-1}$



# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Concentration* Affect the Equilibrium Position of the System?

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ $\Delta H = -196\ \text{kJ mol}^{-1}$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Add Sulfur Dioxide, $SO_2$ | Shifts to the ____ to ____ the amount / moles of $SO_2$. | Concentration of $SO_2$ ____ so the rate of the forward reaction ____. |
| Remove Sulfur Trioxide, $SO_3$ | Shifts to the ____ to ____ the amount / moles of $SO_3$. | Concentration of $SO_3$ ____, so the rate of the backward reaction ____. |

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Concentration* Affect the Equilibrium Position of the System?

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ $\Delta H = -196\ kJ\ mol^{-1}$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Add Sulfur Dioxide, $SO_2$ | Shifts to the **right** to **decrease** the amount / moles of $SO_2$. | Concentration of $SO_2$ **increases** so the rate of the forward reaction **increases**. |
| Remove Sulfur Trioxide, $SO_3$ | Shifts to the **right** to **increase** the amount / moles of $SO_3$. | Concentration of $SO_3$ **decreases**, so the rate of the backward reaction **decreases**. |

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Concentration* Affect the Equilibrium Position of the System?

### Generalisation for Concentration

- **Decreasing** the concentration of a chemical favours the reaction that produces **more** moles of that chemical.
- **Increasing** the concentration of a chemical favours the reaction that produces **fewer** moles of that chemical.

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Temperature* Affect the Equilibrium Position of the System?

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H = -196\ \text{kJ mol}^{-1}$$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Decrease Temperature | Shifts to the ____ to ____ thermal energy since the forward reaction is ____. | Rate of the ____ reaction increases more than the rate of the ____ reaction. |
| *Increase Temperature | Shifts to the ____ to ____ thermal energy since the backward reaction is ____. | Rate of the ____ reaction increases more than the rate of the ____ reaction. |

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Temperature* Affect the Equilibrium Position of the System?

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ $\Delta H = -196\text{ kJ mol}^{-1}$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Decrease Temperature | Shifts to the **right** to **release** thermal energy since the forward reaction is **exothermic**. | Rate of the **forward** reaction increases more than the rate of the **backward** reaction. |
| *Increase Temperature | Shifts to the **left** to **absorb** thermal energy since the backward reaction is **endothermic**. | Rate of the **backward** reaction increases more than the rate of the **forward** reaction. |

***Note:** Increasing the temperature increases the rates of both the forward and backward reactions, but the rate of the endothermic reaction increases more that the rate of the exothermic reaction.

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Temperature* Affect the Equilibrium Position of the System?

### Generalisation for Temperature

- **Decreasing** temperature favours the **exothermic** reaction (to release more thermal energy, thereby increasing temperature).
- **Increasing** temperature favours the **endothermic** reaction (to absorb excess thermal energy, thereby decreasing temperature).

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Pressure* Affect the Equilibrium Position of the System?

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ $\Delta H = -196\ \text{kJ mol}^{-1}$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Decrease Pressure (increase volume of reaction vessel) | Shifts to the ____ to ____ pressure since there are ____ moles of gas on the ____ side of the equation. | Rate of the ____ reaction increases more than rate of the ____ reaction. |
| Increase Pressure (reduce volume of reaction vessel) | Shifts to the ____ to ____ pressure since there are ____ moles of gas on the ____ side of the equation. | Rate of the ____ reaction increases more than rate of the ____ reaction. |

# Reversible Reactions – Dynamic Equilibrium

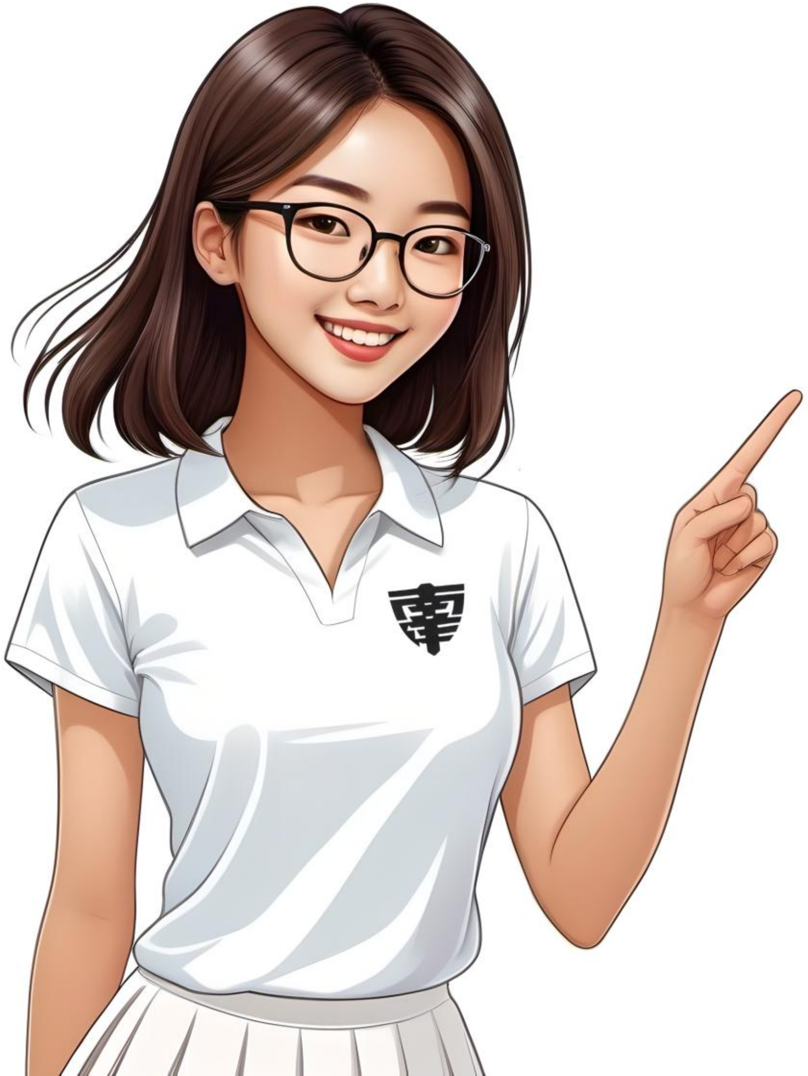
## How Does a Change in *Pressure* Affect the Equilibrium Position of the System?

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$      $\Delta H = -196\ kJ\ mol^{-1}$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Decrease Pressure<br>(increase volume of reaction vessel) | Shifts to the **left** to **increase** pressure since there are **more** moles of gas on the **left** side of the equation. | Rate of the **backward** reaction increases more than rate of the **forward** reaction. |
| Increase Pressure<br>(reduce volume of reaction vessel) | Shifts to the **right** to **decrease** pressure since there are **fewer** moles of gas on the **right** side of the equation. | Rate of the **forward** reaction increases more than rate of the **backward** reaction. |

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Pressure* Affect the Equilibrium Position of the System?



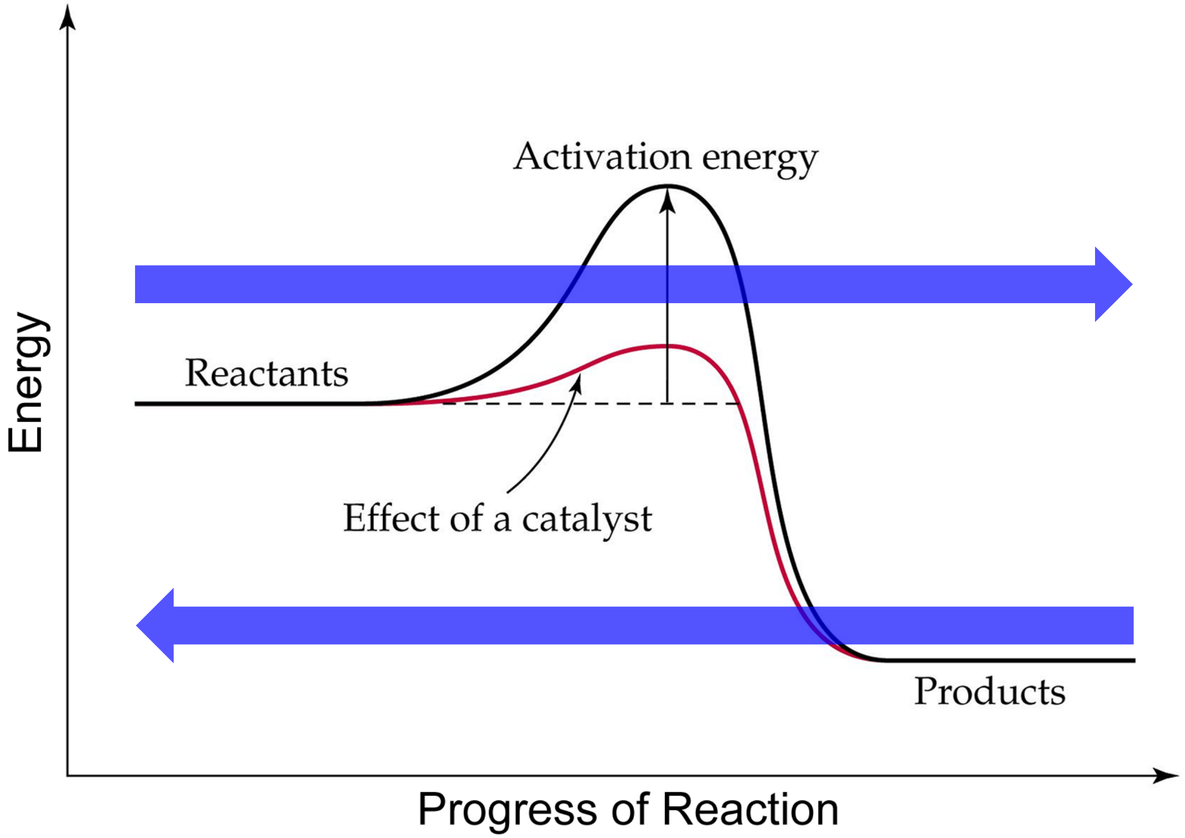
### Generalisation for Pressure

- **Decreasing** pressure favours the reaction that produces **more** moles of gas, thereby increasing pressure.
- **Increasing** pressure favours the reaction that produces **fewer** moles of gas, thereby decreasing pressure.

# Reversible Reactions – Dynamic Equilibrium

## How Does the Addition of a *Catalyst* Affect the Equilibrium Position of the System?

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ $\Delta H = -196\ kJ\ mol^{-1}$



- A catalyst lowers the activation energy of both the *forward* and *backward* reactions.

# Reversible Reactions – Dynamic Equilibrium

## How Does the Addition of a *Catalyst* Affect the Equilibrium Position of the System?

### Generalisation for a Catalyst

- Addition of a catalyst **does not change** the equilibrium position.
- A catalyst **lowers** the **activation energy** of both the forward reaction and backward reaction **by the same extent**.
- Hence, rates of both the forward reaction and backward reaction **increase by the same extent**.
- Therefore, catalysts only allow reactions to reach equilibrium **faster**.

# Reversible Reactions – Dynamic Equilibrium

Could I please see some *more examples* of Chatelier's Principle?

# Reversible Reactions – Dynamic Equilibrium

$2CrO_4^{2-}(aq)$ (yellow) + $2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq)$ (orange) + $H_2O(l)$

- Video: 28 Seconds

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Concentration* Affect the Equilibrium Position of the System?

$2CrO_4^{2-}(aq)$ (**yellow**) + $2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq)$ (**orange**) + $H_2O(l)$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Add Dilute Acid, $H^+(aq)$ | Shifts to the ____ to ____ the concentration of $H^+(aq)$ – solution changes ____ to ____. | Concentration of $H^+(aq)$ ____ so the rate of the forward reaction ____. |
| Add Dilute Alkali, $OH^-(aq)$ | Shifts to the ____ to ____ the concentration of $H^+(aq)$ – solution changes ____ to ____. | $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$<br>Concentration of $H^+(aq)$ ____, so the rate of the forward reaction ____. |

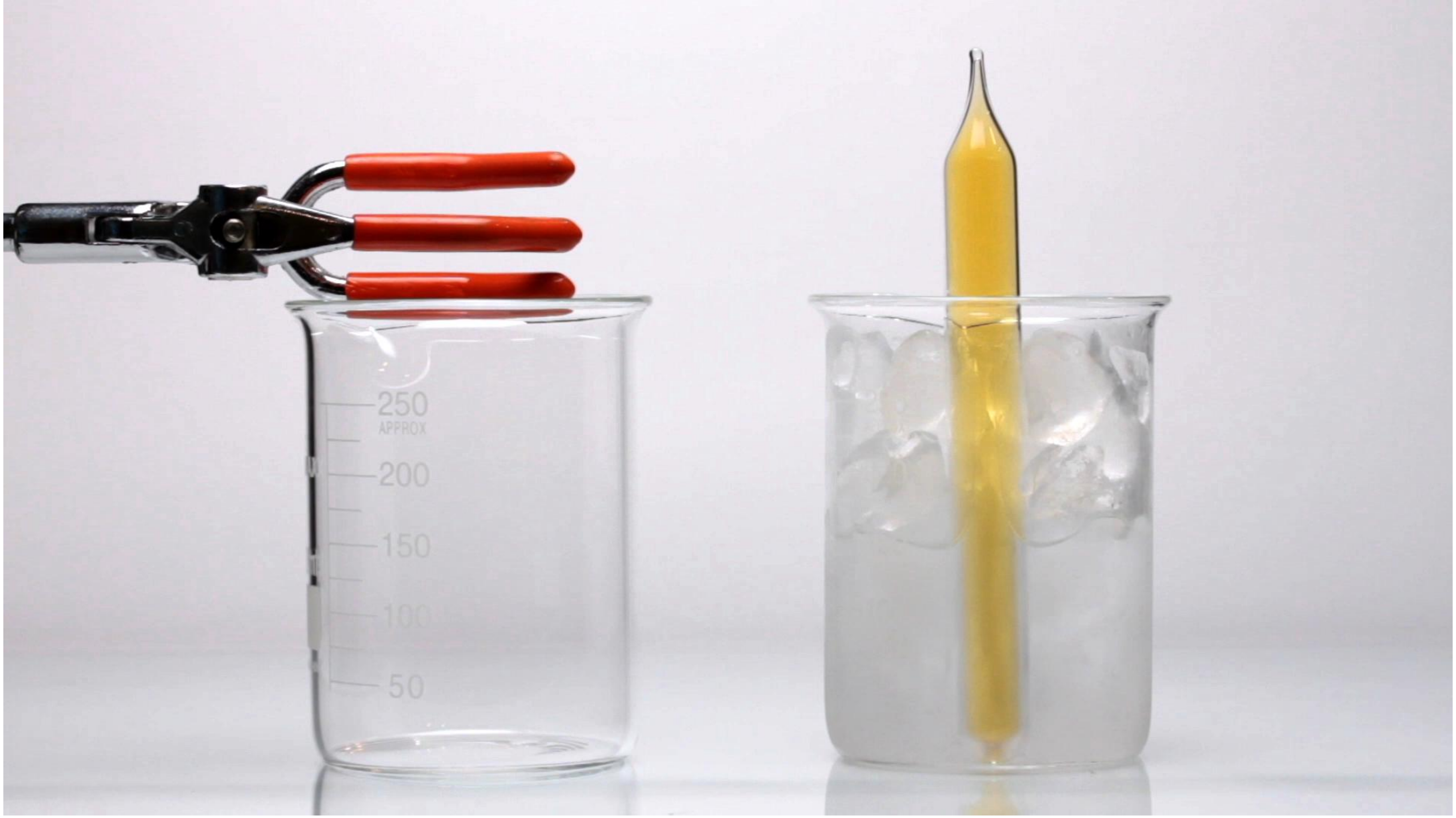
# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Concentration* Affect the Equilibrium Position of the System?

$2CrO_4^{2-}(aq)$ (**yellow**) + $2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq)$ (**orange**) + $H_2O(l)$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Add Dilute Acid, $H^+(aq)$ | Shifts to the **right** to **decrease** the concentration of $H^+(aq)$ – solution changes **yellow** to **orange**. | Concentration of $H^+(aq)$ **increases** so the rate of the forward reaction **increases**. |
| Add Dilute Alkali, $OH^-(aq)$ | Shifts to the **left** to **increase** the concentration of $H^+(aq)$ – solution changes **orange** to **yellow**. | $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$<br>Concentration of $H^+(aq)$ **decreases**, so the rate of the forward reaction **decreases**. |

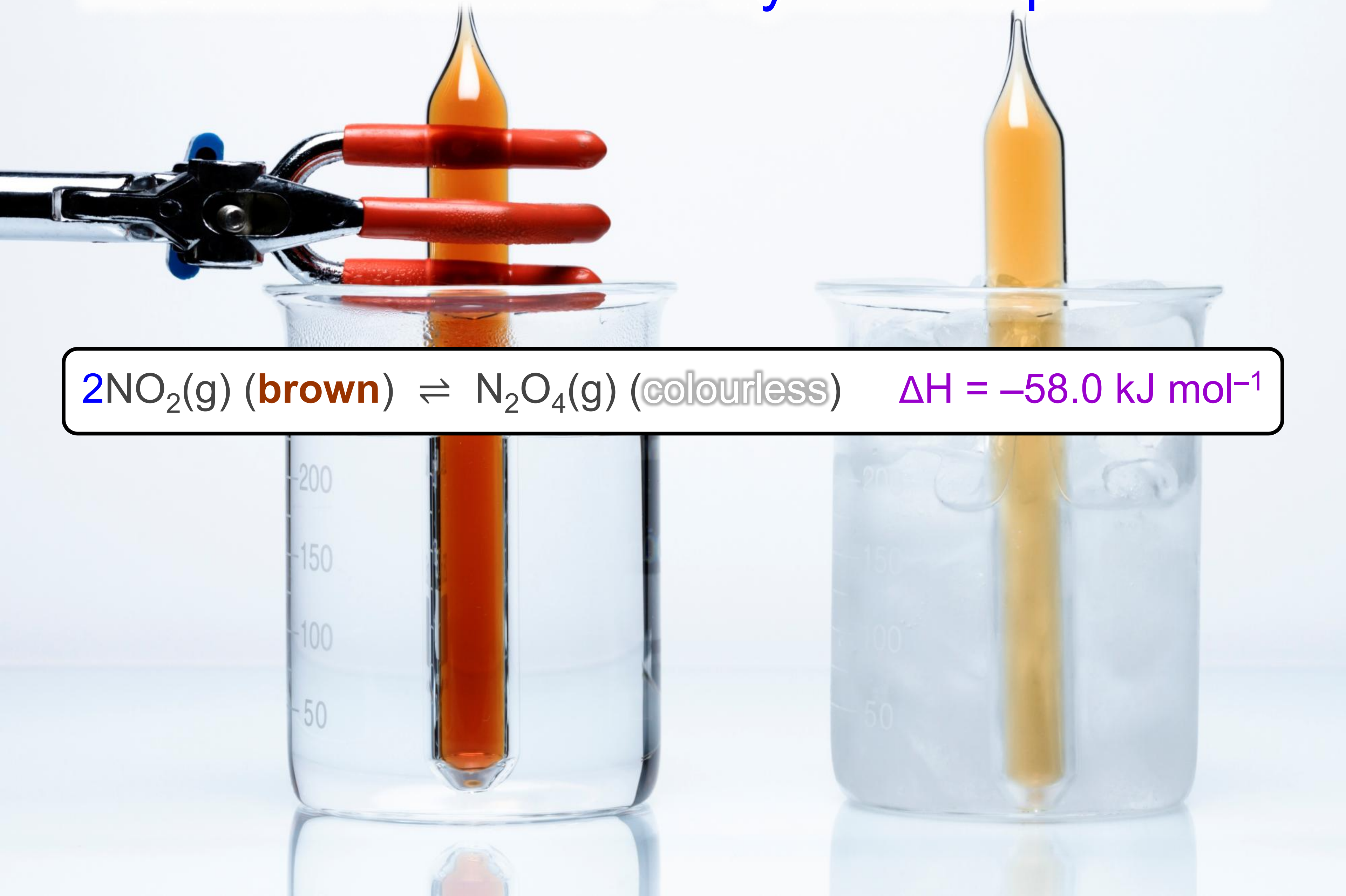
# Reversible Reactions – Dynamic Equilibrium



$2NO_2(g)$ (**brown**) $\rightleftharpoons$ $N_2O_4(g)$ (colourless) $\Delta H = -58.0\ \mathrm{kJ\ mol^{-1}}$

- Video: 64 Seconds

# Reversible Reactions – Dynamic Equilibrium



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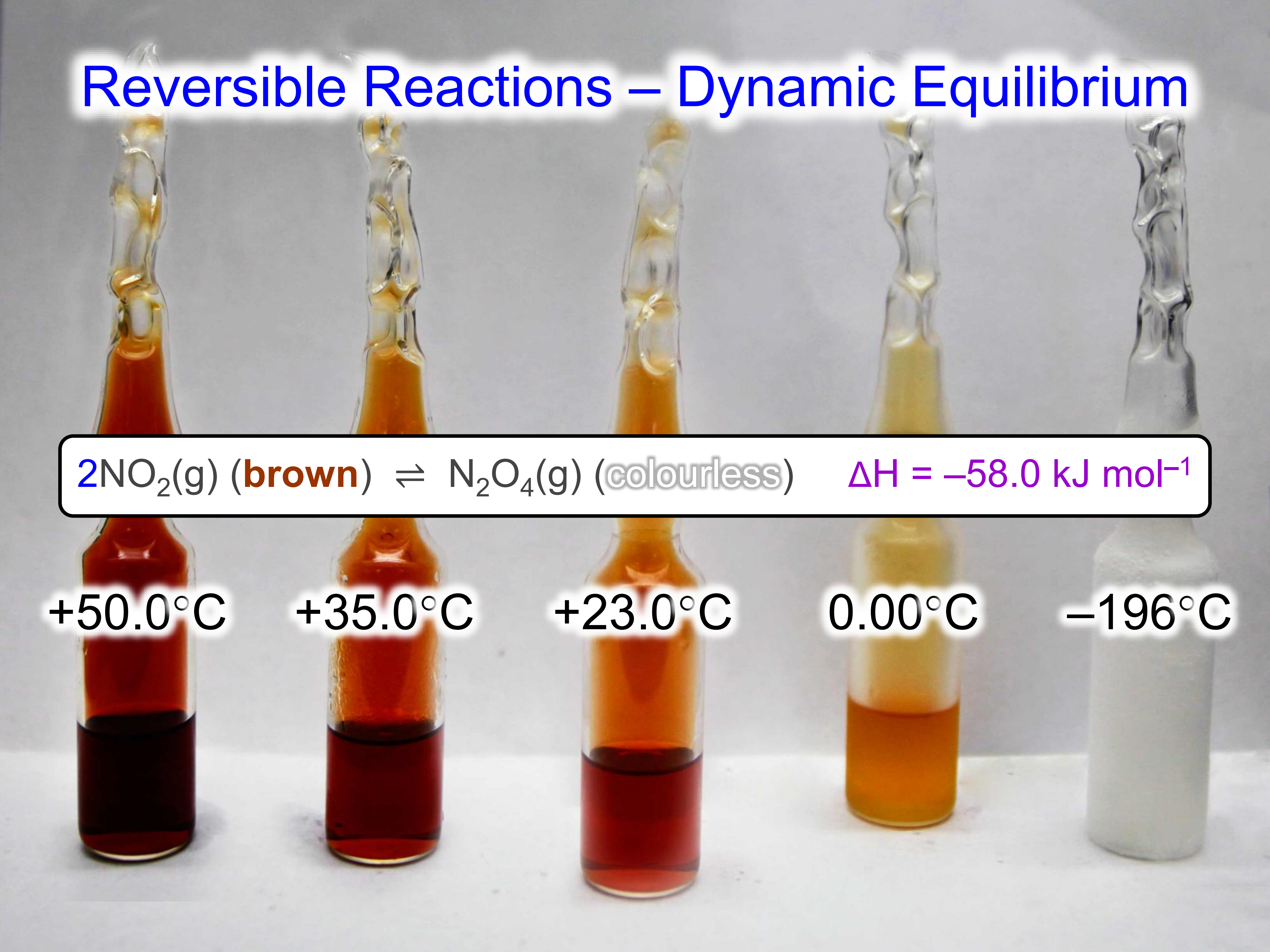
# Reversible Reactions – Dynamic Equilibrium



$2NO_2(g)$ (**brown**) $\rightleftharpoons$ $N_2O_4(g)$ (colourless) $\Delta H = –58.0\ kJ\ mol^{-1}$

+50.0°C +35.0°C +23.0°C 0.00°C –196°C

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Temperature* Affect the Equilibrium Position of the System?

$2NO_2(g)$ (**brown**) $\rightleftharpoons$ $N_2O_4(g)$ (colourless) $\Delta H = -58.0\ \text{kJ mol}^{-1}$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Decrease Temperature | Shifts to the ______ to ______ thermal energy since the forward reaction is ______. System changes from ______ to ______. | Rate of the ______ reaction increases more than the rate of the ______ reaction. |
| *Increase Temperature | Shifts to the ______ to ______ thermal energy since the backward reaction is ______. System changes form ______ to ______. | Rate of the ______ reaction increases more than the rate of the ______ reaction. |

# Reversible Reactions – Dynamic Equilibrium

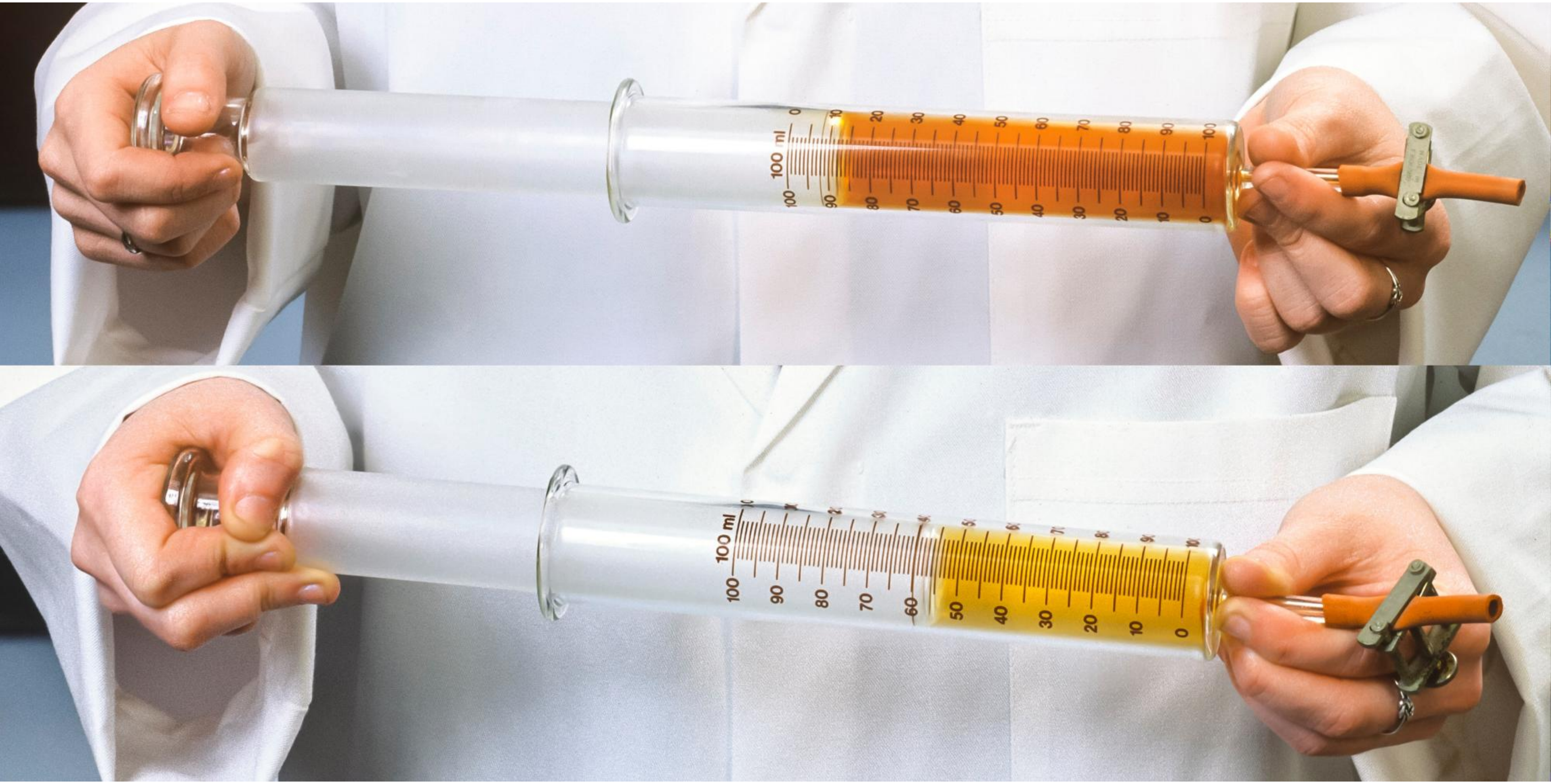
## How Does a Change in *Temperature* Affect the Equilibrium Position of the System?

$2NO_2(g)$ (**brown**) $\rightleftharpoons$ $N_2O_4(g)$ (colourless)   $\Delta H = -58.0\ \text{kJ mol}^{-1}$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Decrease Temperature | Shifts to the **right** to **release** thermal energy since the forward reaction is **exothermic**. System changes from **brown** to **colourless**. | Rate of the **forward** reaction increases more than the rate of the **backward** reaction. |
| *Increase Temperature | Shifts to the **left** to **absorb** thermal energy since the backward reaction is **endothermic**. System changes form **colourless** to **brown**. | Rate of the **backward** reaction increases more than the rate of the **forward** reaction. |

*__Note:__ Increasing the temperature increases the rates of both the forward and backward reactions, but the rate of the endothermic reaction increases more that the rate of the exothermic reaction.

# Reversible Reactions – Dynamic Equilibrium



$2NO_2(g)$ (**brown**) $\rightleftharpoons$ $N_2O_4(g)$ (colourless) $\Delta H = -58.0\ \text{kJ mol}^{-1}$

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Pressure* Affect the Equilibrium Position of the System?

$2NO_2(g)$ (**brown**) $\rightleftharpoons$ $N_2O_4(g)$ (colourless) $\Delta H = -58.0\ \text{kJ mol}^{-1}$

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|---|---|---|
| Decrease Pressure<br>(increase volume of reaction vessel) | Shifts to the ____ to ____ pressure since there are ____ moles of gas on the ____ side of the equation. System changes ____ to ____. | Rate of the ____ reaction increases more than rate of the ____ reaction. |
| Increase Pressure<br>(reduce volume of reaction vessel) | Shifts to the ____ to ____ pressure since there are ____ moles of gas on the ____ side of the equation. System changes ____ to ____. | Rate of the ____ reaction increases more than rate of the ____ reaction. |

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Pressure* Affect the Equilibrium Position of the System?

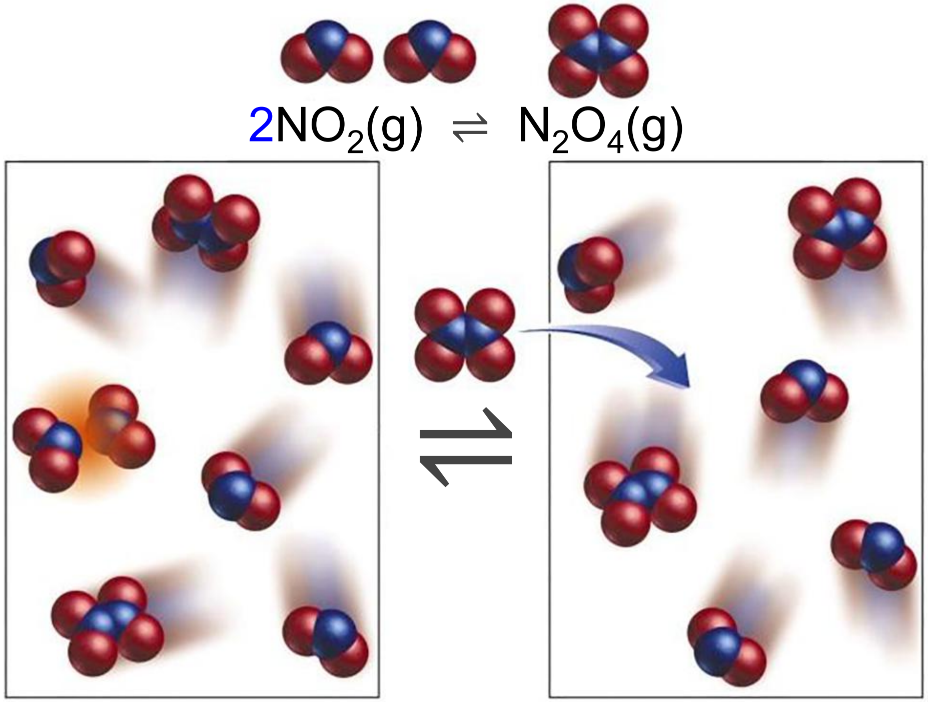
$2NO_2(g)$ (**brown**) $\rightleftharpoons$ $N_2O_4(g)$ (colourless) $\Delta H = -58.0\ \text{kJ mol}^{-1}$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Decrease Pressure<br>(increase volume of reaction vessel) | Shifts to the **left** to **increase** pressure since there are **more** moles of gas on the **left** side of the equation. System changes **colourless** to **brown**. | Rate of the **backward** reaction increases more than rate of the **forward** reaction. |
| Increase Pressure<br>(reduce volume of reaction vessel) | Shifts to the **right** to **decrease** pressure since there are **fewer** moles of gas on the **right** side of the equation. System changes **brown** to **colourless**. | Rate of the **forward** reaction increases more than rate of the **backward** reaction. |

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Concentration* Affect the Equilibrium Position of the System?

$2NO_2(g)$ (**brown**) $\rightleftharpoons$ $N_2O_4(g)$ (colourless)   $\Delta H = -58.0\ kJ\ mol^{-1}$



- What is observed when the amount / moles of $N_2O_4(g)$ is increased or decreased?

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Concentration* Affect the Equilibrium Position of the System?

$2NO_2(g)$ (**brown**) $\rightleftharpoons$ $N_2O_4(g)$ (colourless) $\Delta H = -58.0\ \text{kJ mol}^{-1}$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Increase the Amount of $N_2O_4(g)$ | Shifts to the ____ to ____ the amount of $N_2O_4(g)$ since there are ____ moles of $N_2O_4(g)$ on the ____ side of the equation. System changes ____ to ____. | Rate of the ____ reaction ____ because the amount of $N_2O_4(g)$ on the right-hand-side has ____. |
| Decrease the Amount of $N_2O_4(g)$ | Shifts to the ____ to ____ the amount of $N_2O_4(g)$ since there are ____ moles of $N_2O_4(g)$ on the ____ side of the equation. System changes ____ to ____. | Rate of the ____ reaction ____ because the amount of $N_2O_4(g)$ on the right-hand-side has ____. |

# Reversible Reactions – Dynamic Equilibrium

## How Does a Change in *Concentration* Affect the Equilibrium Position of the System?

$2NO_2(g)$ (**brown**) $\rightleftharpoons$ $N_2O_4(g)$ (colourless) $\Delta H = -58.0\ \text{kJ mol}^{-1}$

| Change | Direction of Equilibrium Shift Based on Le Chatelier's Principle | Explanation in Terms of Rate of Reaction |
|---|---|---|
| Increase the Amount of $N_2O_4(g)$ | Shifts to the **left** to **decrease** the amount of $N_2O_4(g)$ since there are **fewer** moles of $N_2O_4(g)$ on the **left** side of the equation. System changes **colourless** to **brown**. | Rate of the **backward** reaction **increases** because the amount of $N_2O_4(g)$ on the right-hand-side has **increased**. |
| Decrease the Amount of $N_2O_4(g)$ | Shifts to the **right** to **increase** the amount of $N_2O_4(g)$ since there are **more** moles of $N_2O_4(g)$ on the **right** side of the equation. System changes **brown** to **colourless**. | Rate of the **backward** reaction **decreases** because the amount of $N_2O_4(g)$ on the right-hand-side has **decreased**. |

# Reversible Reactions – Dynamic Equilibrium

How should I *answer questions* about reversible reactions?

# Reversible Reactions – Dynamic Equilibrium

## General Guideline to Answering Questions

**a)** State whether the equilibrium position shifts towards the left-hand-side or the right-hand-side.

**b)** With reference to Le Chatelier's Principle, explain *why* the equilibrium position shifts in response to the change.

**c)** State whether the forward or backward reaction is favoured.

**d)** State any consequences due to the change in equilibrium position, *e.g.* change in colour.

# Reversible Reactions – Dynamic Equilibrium

$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(l)$ ΔH is positive

**1.** Add $C_2H_4(g)$

# Reversible Reactions – Dynamic Equilibrium

$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(l)$   ΔH is positive

**1.** Add $C_2H_4(g)$

**a)** As the amount / moles of $C_2H_4$ increases, the equilibrium position will shift to the right.

**b)** By Le Chatelier's Principle, the equilibrium position will shift to the right to reduce the amount / moles of $C_2H_4$.

**c)** This favours the forward reaction (left-to-right).

**d)** Which increases the yield of ethanol ($C_2H_5OH$).

# Reversible Reactions – Dynamic Equilibrium

$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(l)$ ΔH is positive

**2.** Remove $C_2H_5OH$

# Reversible Reactions – Dynamic Equilibrium

$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(l)$     ΔH is positive

**2.** Remove $C_2H_5OH$

**a)** As the amount / moles of $C_2H_5OH$ decreases, the equilibrium position will shift to the right.

**b)** By Le Chatelier's Principle, the equilibrium position will shift to the right to increase the amount / moles of $C_2H_5OH$.

**c)** This favours the forward reaction (left-to-right).

**d)** Which increases the yield of ethanol ($C_2H_5OH$).

# Reversible Reactions – Dynamic Equilibrium

$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(l)$  ΔH is positive

**3.** Increase Temperature

# Reversible Reactions – Dynamic Equilibrium

$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(l)$ ΔH is positive

## **3.** Increase Temperature

**a)** As the temperature increases, the equilibrium position will shift to the right.

**b)** By Le Chatelier’s Principle, as the temperature increases, the equilibrium position will shift in the direction that absorbs thermal energy (endothermic direction) and lowers the temperature.

**c)** This favours the forward reaction (left-to-right).

**d)** Which increases the yield of ethanol ($C_2H_5OH$).

# Reversible Reactions – Dynamic Equilibrium

$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(l)$     ΔH is positive

**4.** Increase Pressure

# Reversible Reactions – Dynamic Equilibrium

$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(l)$ ΔH is positive

## **4.** Increase Pressure

**a)** As the pressure increases, the equilibrium position will shift to the right.

**b)** By Le Chatelier’s Principle, the equilibrium position will shift to the right to reduce the pressure because the right-hand-side contains fewer moles of gas (lower pressure) compared to the left-hand-side (higher pressure).

**c)** This favours the forward reaction (left-to-right).

**d)** Which increases the yield of ethanol ($C_2H_5OH$).

# Reversible Reactions – Dynamic Equilibrium

What conditions are used for the *industrial manufacture of ammonia*?

- What conditions would you use?

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

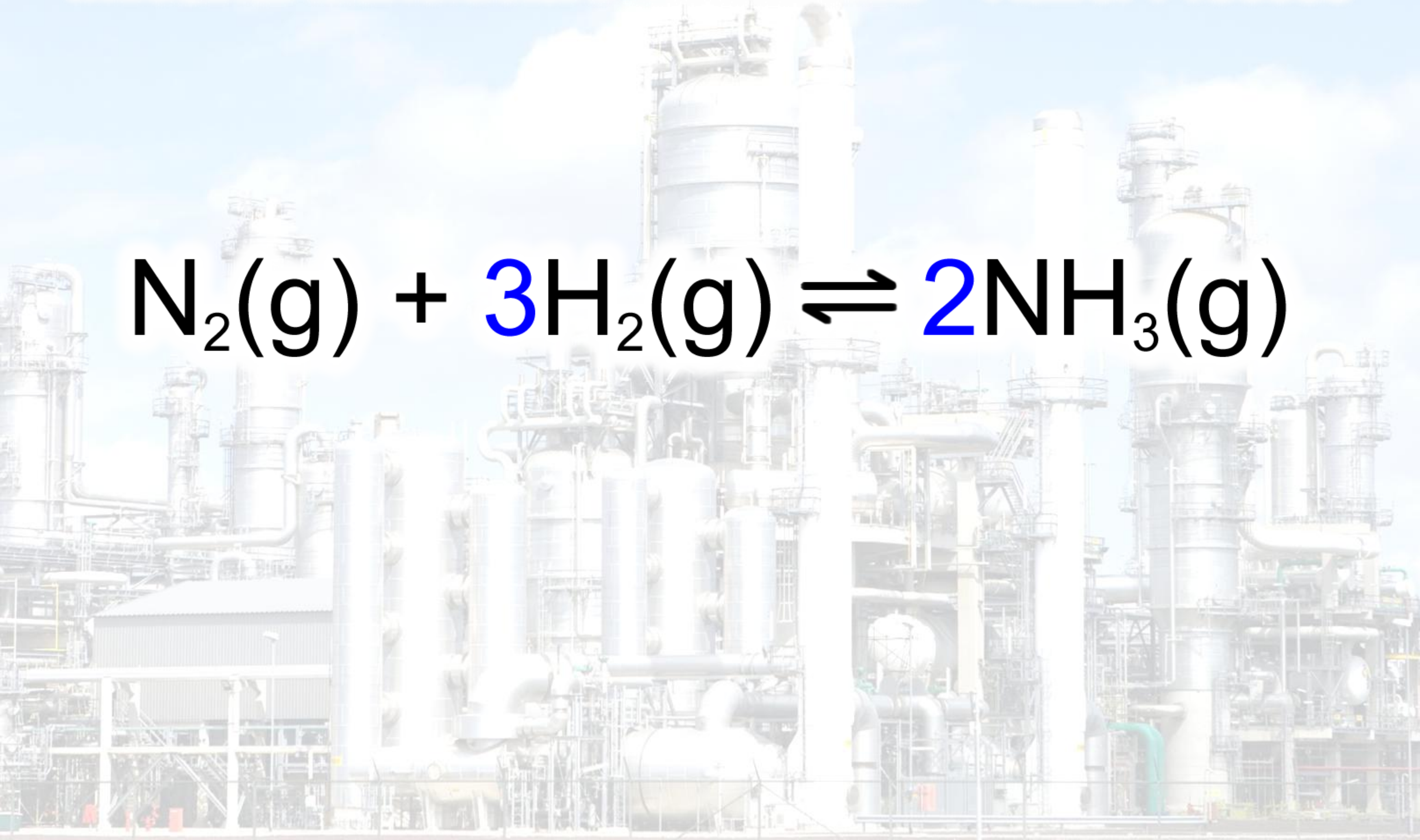
$$\Delta H = -92.4 \text{ kJ mol}^{-1}$$

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

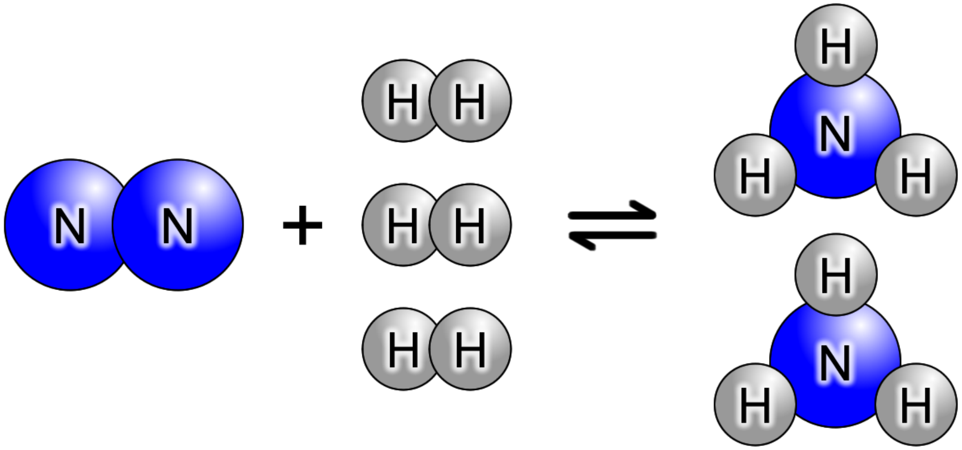


$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

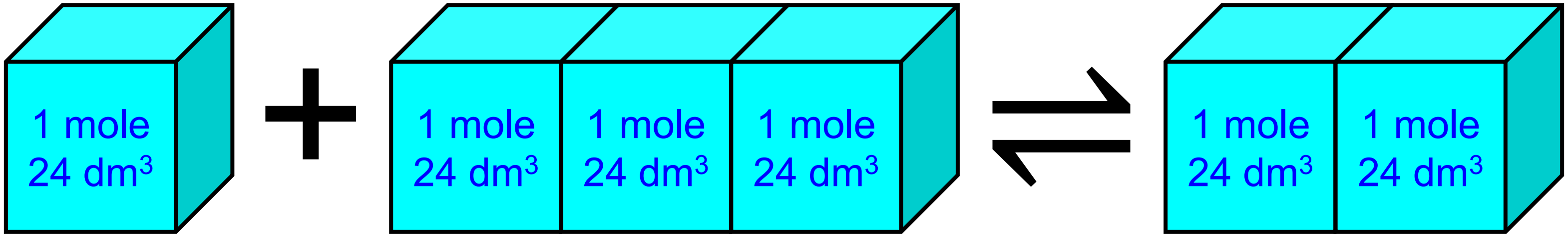
# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

nitrogen + hydrogen ⇌ ammonia



$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$  $\Delta H = -92.4\ \text{kJ mol}^{-1}$

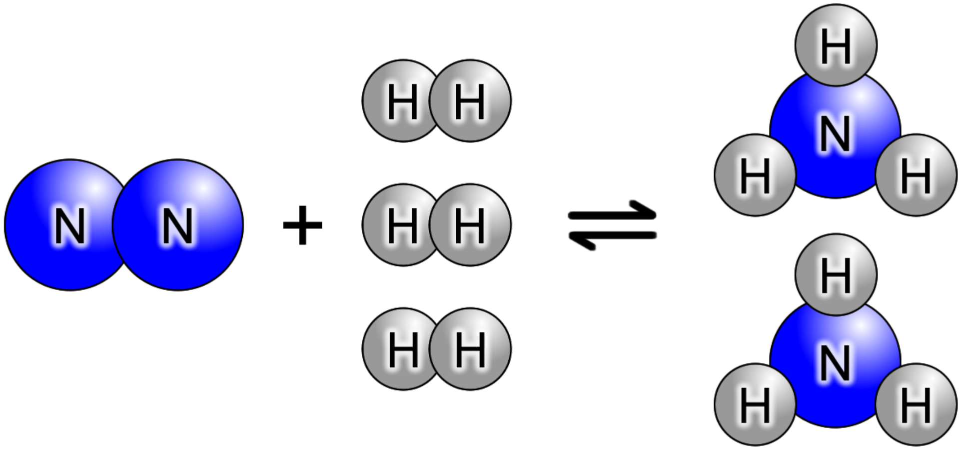


1 mole $24\ dm^3$ + 1 mole $24\ dm^3$ | 1 mole $24\ dm^3$ | 1 mole $24\ dm^3$ ⇌ 1 mole $24\ dm^3$ | 1 mole $24\ dm^3$

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

nitrogen + hydrogen $\rightleftharpoons$ ammonia



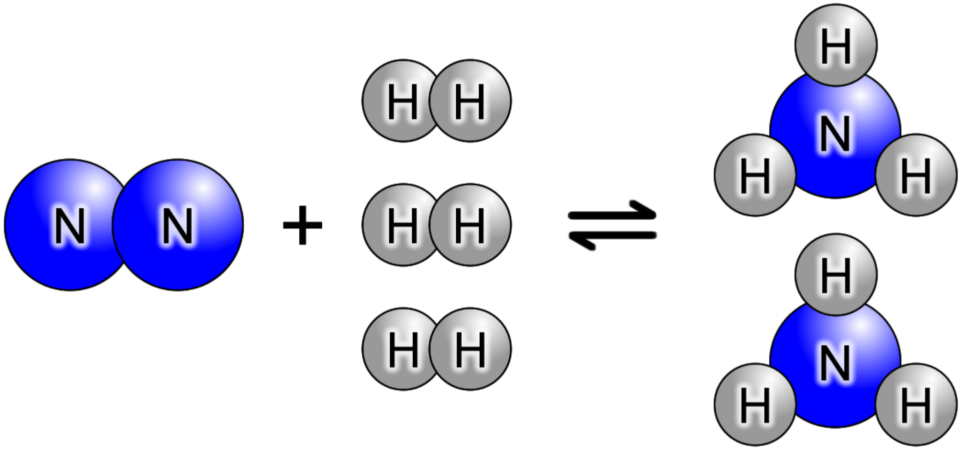
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92.4\ kJ\ mol^{-1}$

- Ammonia is manufactured on an industrial scale by reacting nitrogen gas directly with hydrogen gas under a specialised set of conditions.

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

nitrogen + hydrogen $\rightleftharpoons$ ammonia



$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92.4\ \text{kJ mol}^{-1}$

- The chemical reaction used to manufacture ammonia on an industrial scale is *reversible*. This means that while nitrogen and hydrogen react to form ammonia, ammonia reacts to form nitrogen and hydrogen.

# Reversible Reactions – Dynamic Equilibrium

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.4\ \text{kJ mol}^{-1}$

**1.** Add $N_2(g)$ and $H_2(g)$

# Reversible Reactions – Dynamic Equilibrium

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.4\ \text{kJ mol}^{-1}$$

**1.** Add $N_2(g)$ and $H_2(g)$

**a)** As the amount / moles of $N_2(g)$ and $H_2(g)$ increases, the equilibrium position will shift to the right.

**b)** By Le Chatelier's Principle, the equilibrium position will shift to the right to reduce the amount / moles of $N_2(g)$ and $H_2(g)$.

**c)** This favours the forward reaction (left-to-right).

**d)** Which increases the yield of $NH_3(g)$.

# Reversible Reactions – Dynamic Equilibrium

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.4\ \text{kJ mol}^{-1}$

**2.** Remove $NH_3(g)$

# Reversible Reactions – Dynamic Equilibrium

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.4\ kJ\ mol^{-1}$

**2.** Remove $NH_3(g)$

**a)** As the amount / moles of $NH_3(g)$ decreases, the equilibrium position will shift to the right.

**b)** By Le Chatelier's Principle, the equilibrium position will shift to the right to increase the amount / moles of $NH_3(g)$.

**c)** This favours the forward reaction (left-to-right).

**d)** Which increases the yield of $NH_3(g)$.

# Reversible Reactions – Dynamic Equilibrium

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.4\ \text{kJ mol}^{-1}$$

**3.** Reduce Temperature

# Reversible Reactions – Dynamic Equilibrium

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.4 \text{ kJ mol}^{-1}$

## **3.** Reduce Temperature

**a)** As the temperature decreases, the equilibrium position will shift to the right.

**b)** By Le Chatelier's Principle, as the temperature decreases, the equilibrium position will shift in the direction that releases thermal energy (exothermic direction) and increases the temperature.

**c)** This favours the forward reaction (left-to-right).

**d)** Which increases the yield of $NH_3(g)$.

# Reversible Reactions – Dynamic Equilibrium

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.4 \text{ kJ mol}^{-1}$

**4.** Increase Pressure

# Reversible Reactions – Dynamic Equilibrium

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.4\ \text{kJ mol}^{-1}$$

## **4.** Increase Pressure

**a)** As the pressure increases, the equilibrium position will shift to the right.

**b)** By Le Chatelier's Principle, the equilibrium position will shift to the right to reduce the pressure because the right-hand-side contains fewer moles of gas (lower pressure) compared to the left-hand-side (higher pressure).

**c)** This favours the forward reaction (left-to-right).

**d)** Which increases the yield of $NH_3(g)$.

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

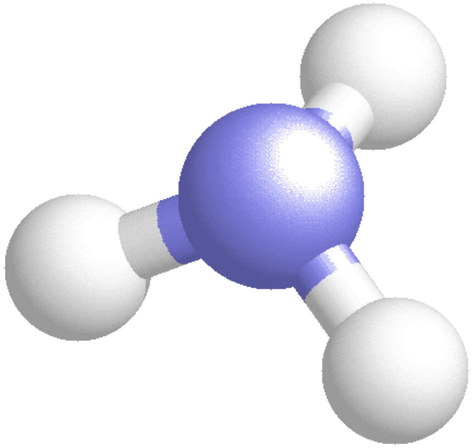
- Hydrogen required for the manufacture of ammonia is obtained by *cracking* long-chain hydrocarbons from crude oil, *e.g.*

$$C_{20}H_{42}(s) \rightarrow C_8H_{16}(l) + C_{12}H_{24}(l) + H_2(g)$$

- Nitrogen required for the manufacture of ammonia is obtained from the *fractional distillation* of liquefied air (remember, the Earth's atmosphere is approximately 78% nitrogen).

# Reversible Reactions – Dynamic Equilibrium

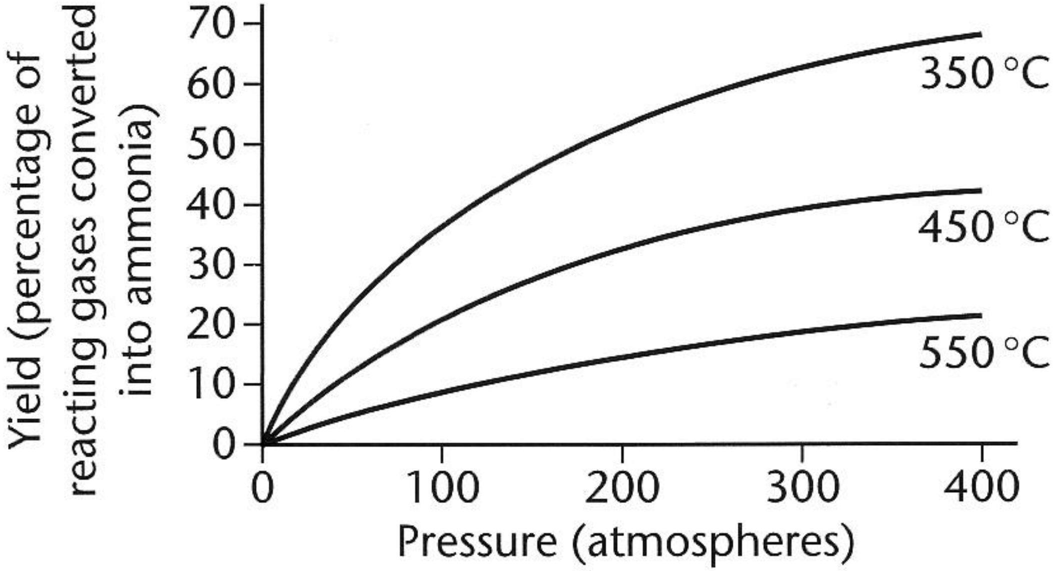
## Industrial Manufacture of Ammonia – Haber Process



The conditions used for the industrial manufacture of ammonia were discovered 100 years ago by a German scientist called *Fritz Haber*. His discovery was so important that he was awarded the 1918 Nobel Prize in Chemistry. The conditions that he discovered are still used in factories around the world today.

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

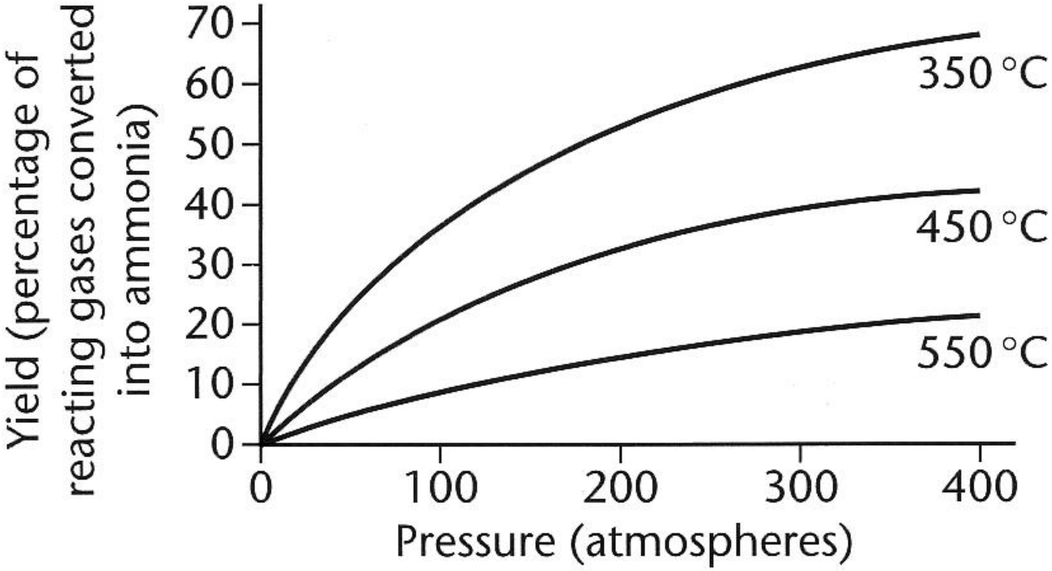


- Based upon the information provided in the graph, what conditions would you use to manufacture ammonia?

- Think about the *yield* of ammonia and the *rate* at which the ammonia is produced.

# Reversible Reactions – Dynamic Equilibrium

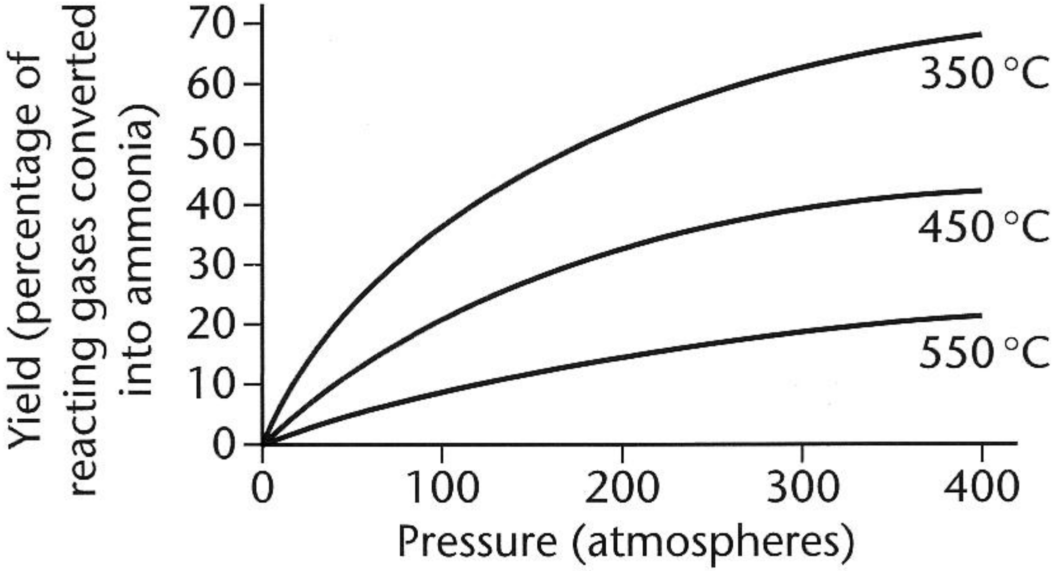
## Industrial Manufacture of Ammonia – Haber Process



- What are the different consequences of using a *high temperature* or a *low temperature*?
- What are the different consequences of using a *high pressure* or a *low pressure*?

# Reversible Reactions – Dynamic Equilibrium

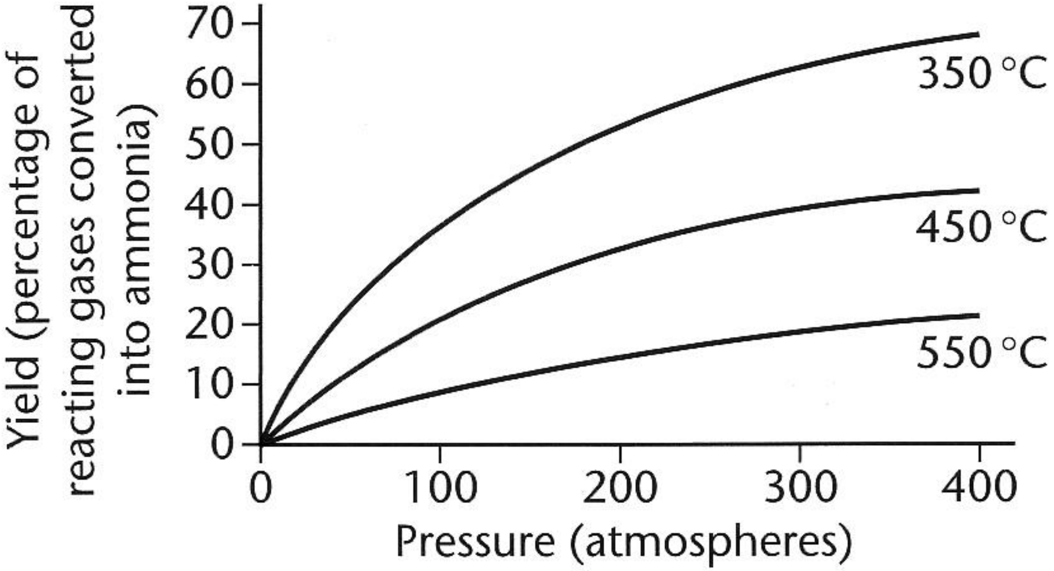
## Industrial Manufacture of Ammonia – Haber Process



- A *low temperature* favours the production of ammonia, but if the temperature is too low, the *rate* at which the ammonia is formed is very *slow*. A temperature of *450 °C* is a compromise, giving a reasonably good yield of ammonia at an acceptable rate.

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process



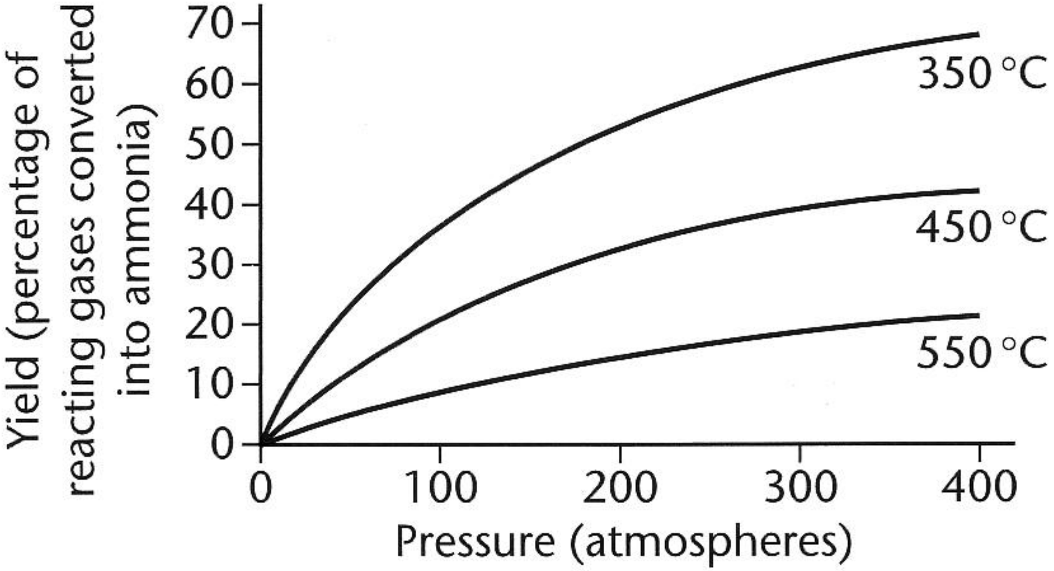
→ → → Low temperature favours the forward reaction → → →

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

←←← High temperature favours the backward reaction ←←←

# Reversible Reactions – Dynamic Equilibrium

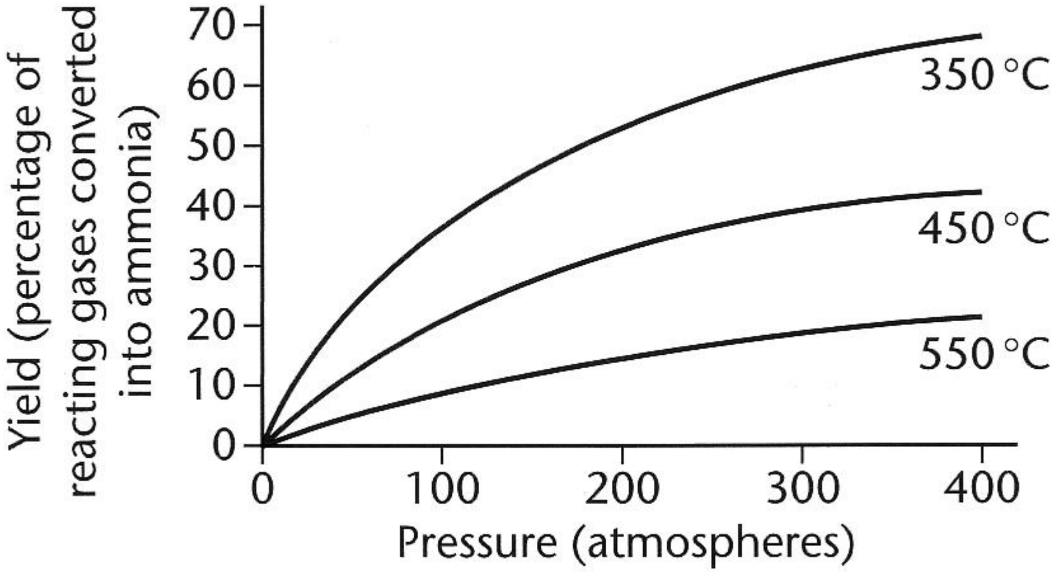
## Industrial Manufacture of Ammonia – Haber Process



- A *high pressure* favours the production of ammonia, but using a very high pressure on an industrial scale is both expensive and potentially dangerous. A pressure of *250 atm.* is a compromise that gives a reasonably good yield of ammonia, that is both cost effective and relatively safe.

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process



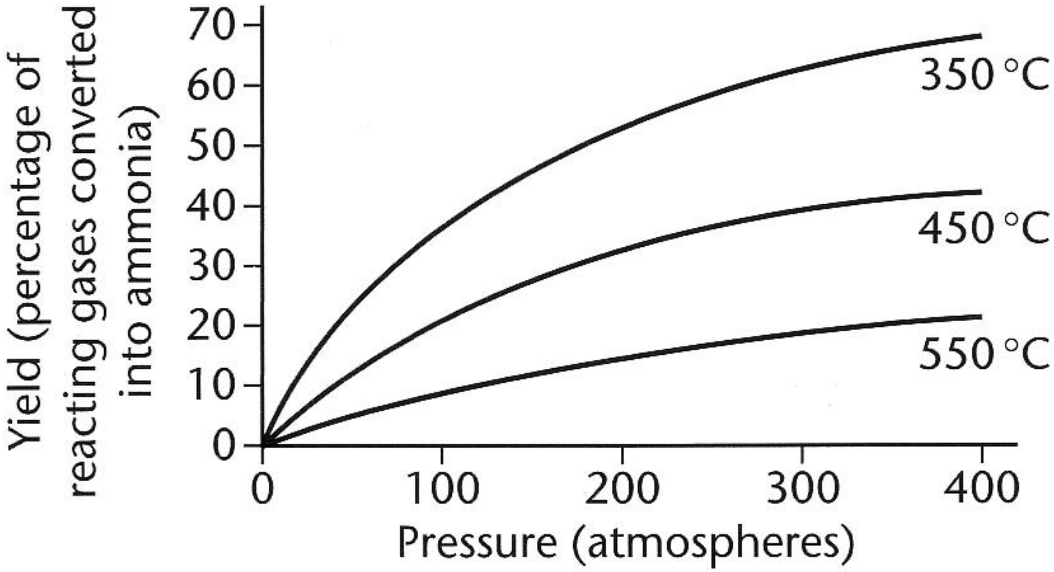
→ → → High pressure favours the forward reaction → → →

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

←←← low pressure favours the backward reaction ←←←

# Reversible Reactions – Dynamic Equilibrium

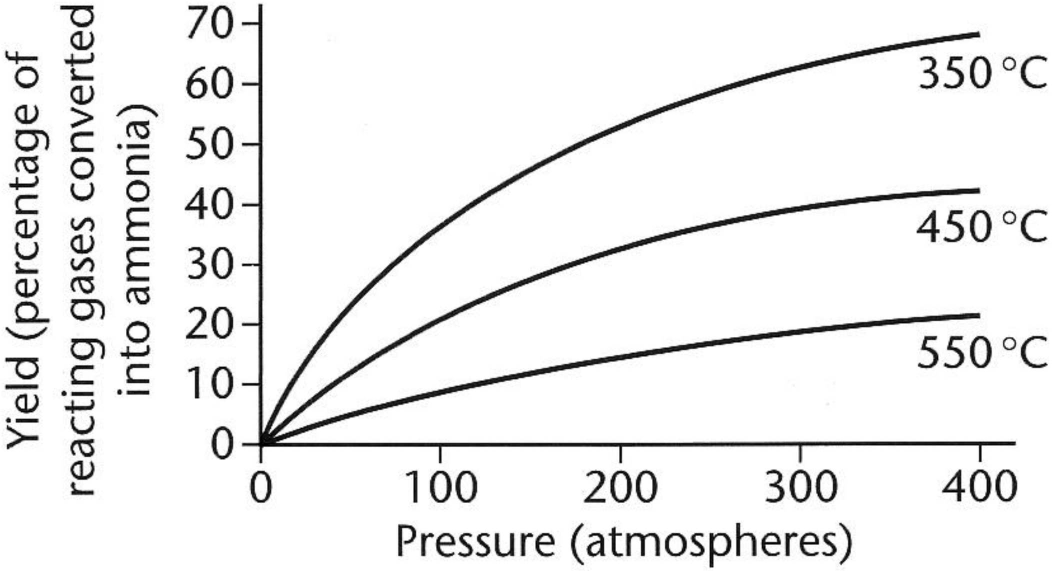
## Industrial Manufacture of Ammonia – Haber Process



- Apart from *temperature* and *pressure*, what else might affect the yield of ammonia and / or the rate at which the ammonia is produced?

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

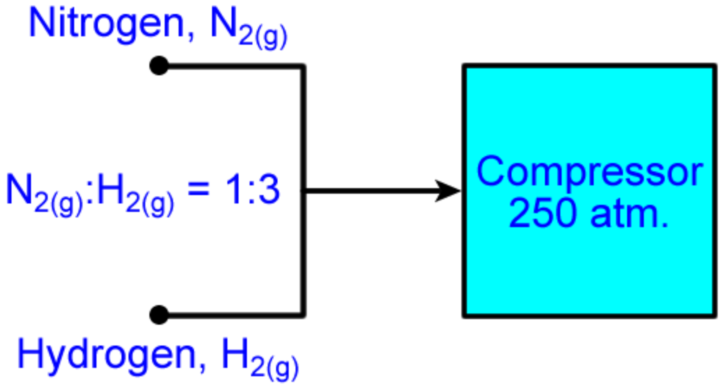


- To increase the rate at which nitrogen and hydrogen react to form ammonia, an *iron catalyst* is used. Note: a catalyst will only increase the *rate* at which the chemical reaction reaches equilibrium, it will not affect the *yield* of the product that is formed.

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

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

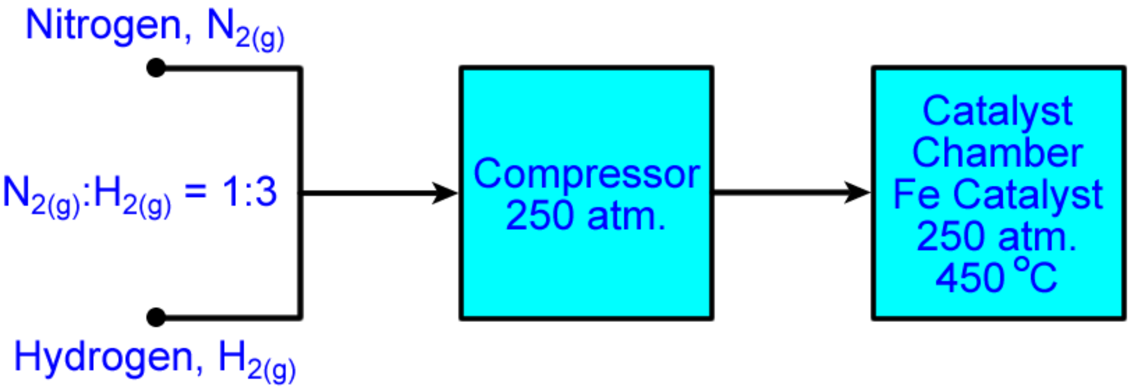


- According to the balanced chemical equation, *1 mol of nitrogen* reacts with *3 mol of hydrogen*, so the nitrogen and hydrogen are combined together in a ratio of *1:3*.

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

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



Nitrogen, $N_{2(g)}$

$N_{2(g)}:H_{2(g)} = 1:3$

Hydrogen, $H_{2(g)}$

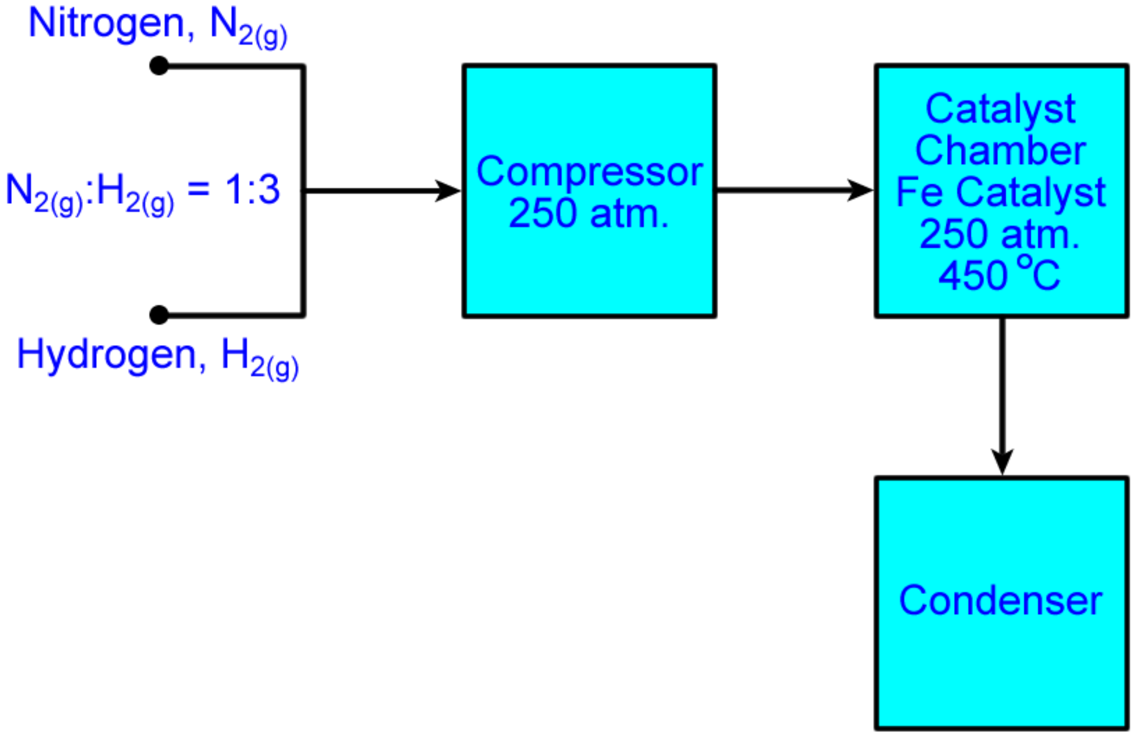
Compressor
250 atm.

Catalyst
Chamber
Fe Catalyst
250 atm.
450 °C

# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

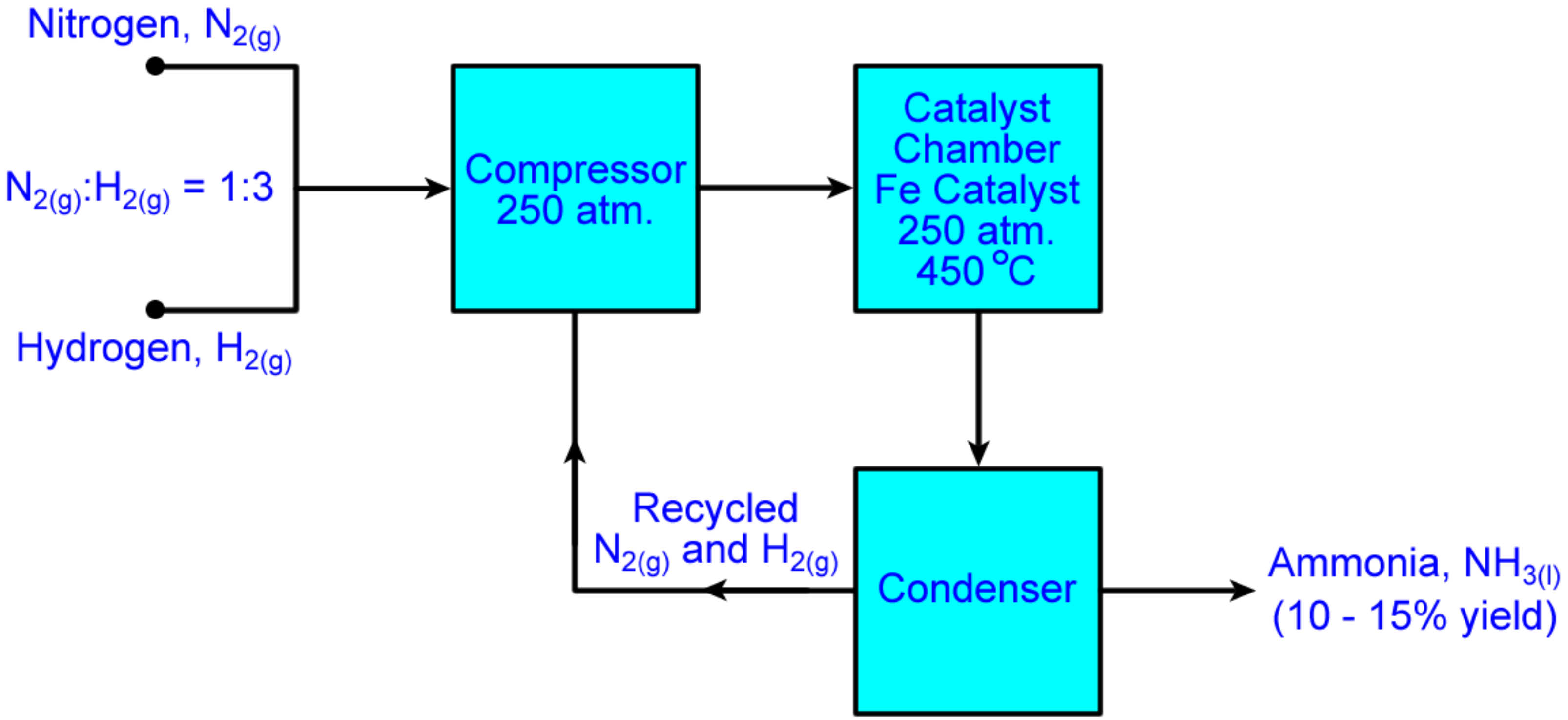
$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$



# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

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



Nitrogen, $N_{2(g)}$

$N_{2(g)}:H_{2(g)} = 1:3$

Hydrogen, $H_{2(g)}$

Compressor
250 atm.

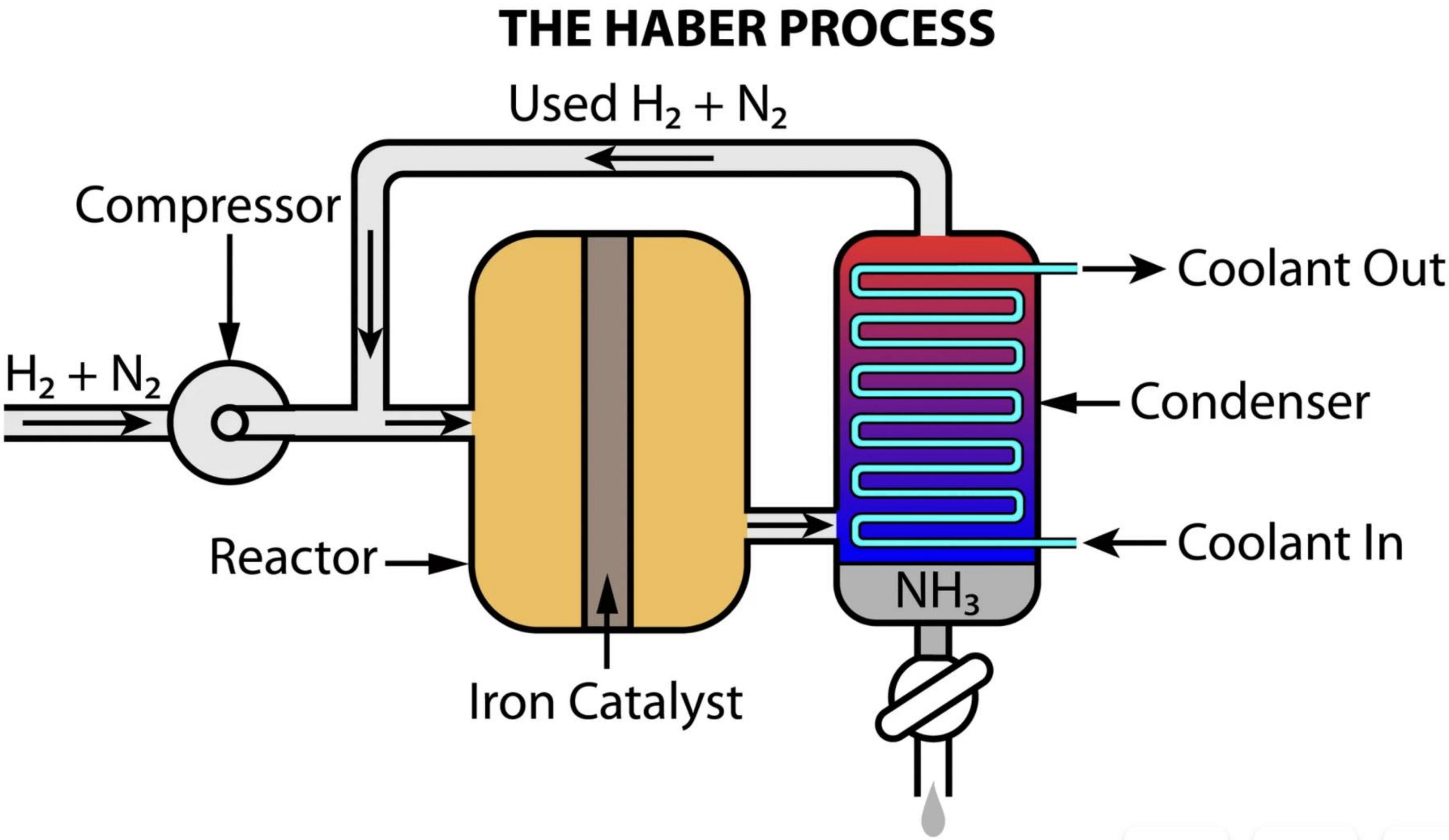
Catalyst
Chamber
Fe Catalyst
250 atm.
450 °C

Condenser

Recycled
$N_{2(g)}$ and $H_{2(g)}$

Ammonia, $NH_{3(l)}$
(10 - 15% yield)

# Reversible Reactions – Dynamic Equilibrium



# Reversible Reactions – Dynamic Equilibrium

## Industrial Manufacture of Ammonia – Haber Process

### Summary

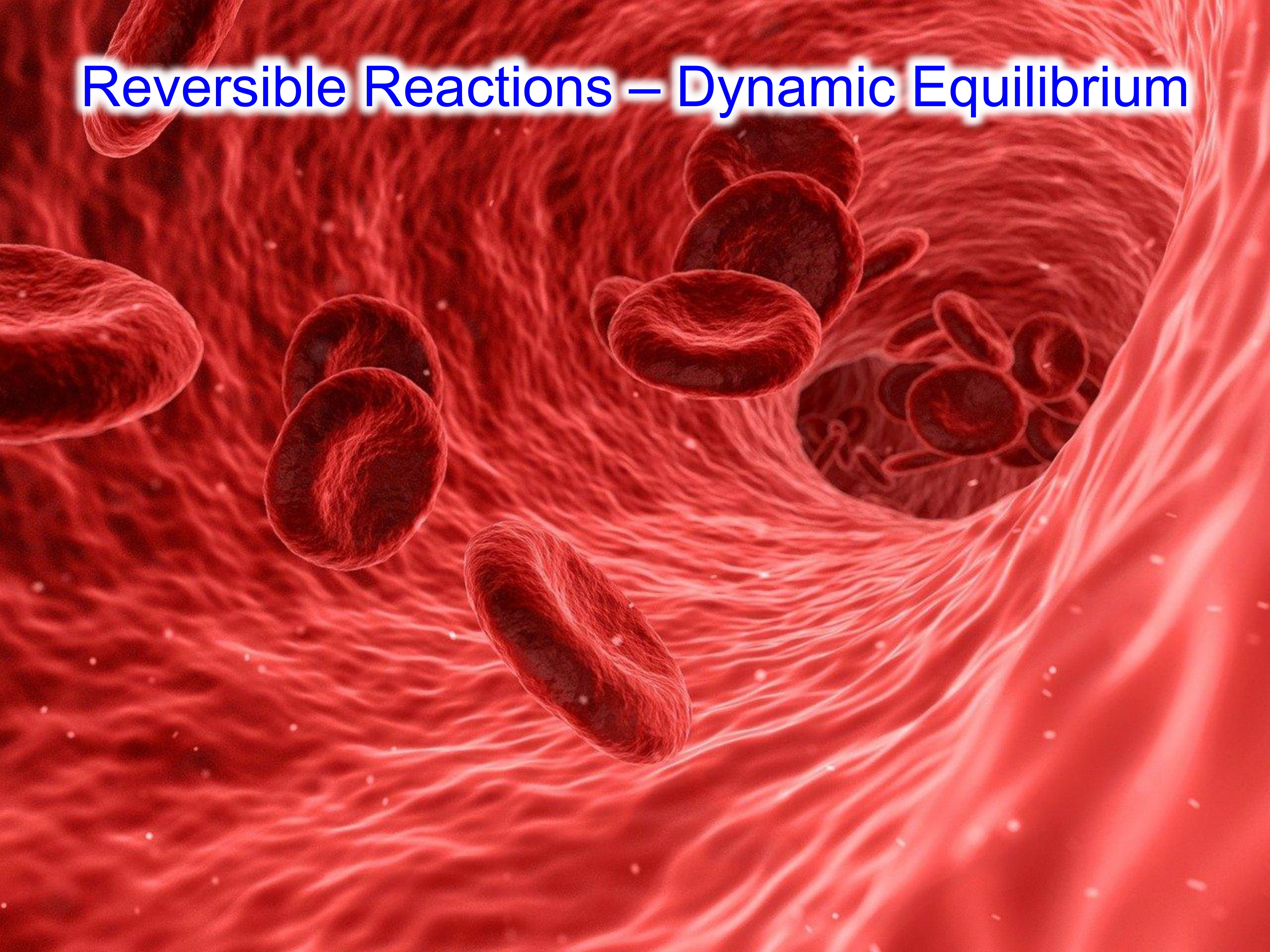
$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

- Hydrogen is obtained from cracking hydrocarbons.
- Nitrogen is obtained from the fractional distillation of liquefied air.
- Catalyst = Iron
- Temperature = 450 °C
- Pressure = 250 atm.

# Reversible Reactions – Dynamic Equilibrium

Are there reversible reactions in *biological systems*?

# Reversible Reactions – Dynamic Equilibrium



# Reversible Reactions – Dynamic Equilibrium

- Red blood cells contain a pigment called *Haemoglobin*. Haemoglobin is responsible for transporting oxygen around the bodies of mammals.

- Haemoglobin contains iron. The iron can change its oxidation state as it bonds with oxygen in the lungs and releases oxygen in the muscles. This change in oxidation state is reversible:

$$Fe^{2+} \rightleftharpoons Fe^{3+} + e^{-}$$

# Reversible Reactions – Dynamic Equilibrium

- In aerobic respiration, oxygen is transported to the cells where it is combined with glucose and metabolised to carbon dioxide which is then transported back to the lungs to be expelled.

- Haemoglobin (symbol Hb) takes up oxygen from the air in which it becomes loosely bound in a complex known as oxyhaemoglobin. The oxygen concentration is reduced by 50% at the ends of the capillaries delivering blood to the tissues. Equilibrium is shifted to the left, releasing the oxygen so it can diffuse into the cells:

$$\text{haemoglobin} + \text{oxygen} \rightleftharpoons \text{oxyhaemoglobin}$$

$$Hb + 4O_2 \rightleftharpoons Hb \cdot 4O_2$$

# Reversible Reactions – Dynamic Equilibrium

Are there *Calculations* for reversible reactions?

# Reversible Reactions – Dynamic Equilibrium

**1.** In a reaction vessel there were *initially* 1.0 mol of $SO_2$, 1.0 mol of $O_2$ and 0.0 mol of $SO_3$.
If the amount of $SO_3$ at equilibrium was 0.4 mol, calculate the amount of $SO_2$ and $O_2$ at equilibrium.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

| | $SO_2$ | $O_2$ | $SO_3$ |
|---|---|---|---|
| Initial moles / mol | 1.0 | 1.0 | 0.0 |
| Change | | | |
| Amount present at equilibrium / mol | | | 0.4 |

# Reversible Reactions – Dynamic Equilibrium

1. In a reaction vessel there were *initially* 1.0 mol of $SO_2$, 1.0 mol of $O_2$ and 0.0 mol of $SO_3$.
   If the amount of $SO_3$ at equilibrium was 0.4 mol, calculate the amount of $SO_2$ and $O_2$ at equilibrium.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

| | $SO_2$ | $O_2$ | $SO_3$ |
|---|---|---|---|
| Initial moles / mol | 1.0 | 1.0 | 0.0 |
| Change | –0.4 | –0.2 | +0.4 |
| Amount present at equilibrium / mol | 0.6 | 0.8 | 0.4 |

# Reversible Reactions – Dynamic Equilibrium

**2.** There were initially 4.0 mol of $NH_3$ in a reaction vessel. If the amounts of $N_2$, $H_2$ and $NH_3$ *at equilibrium* were 2.0 mol, 4.0 mol and 2.0 mol respectively, calculate the initial amounts of $N_2$ and $H_2$ before equilibrium was established.

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

| | $N_2$ | $H_2$ | $NH_3$ |
|---|---|---|---|
| Initial moles / mol | | | 4.0 |
| Change | | | |
| Amount present at equilibrium / mol | 2.0 | 4.0 | 2.0 |

# Reversible Reactions – Dynamic Equilibrium

**2.** There were initially 4.0 mol of $NH_3$ in a reaction vessel. If the amounts of $N_2$, $H_2$ and $NH_3$ *at equilibrium* were 2.0 mol, 4.0 mol and 2.0 mol respectively, calculate the initial amounts of $N_2$ and $H_2$ before equilibrium was established.

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

| | $N_2$ | $H_2$ | $NH_3$ |
|---|---|---|---|
| Initial moles / mol | 1.0 | 1.0 | 4.0 |
| Change | +1.0 | +3.0 | –2.0 |
| Amount present at equilibrium / mol | 2.0 | 4.0 | 2.0 |

# Reversible Reactions – Dynamic Equilibrium

Advanced concepts for enrichment: *equilibrium constants*.

# Reversible Reactions – Dynamic Equilibrium

## Equilibrium Constant, $K_c$

- The position of equilibrium can be estimated by the equilibrium constant $K_c$, which is the ratio between the amount of reactant and the amount of product. It is used to determine the chemical behaviour of the system.

- Example: For the reaction, $A + 2B \rightleftharpoons C + 3D$

$$K_c = \frac{[C][D]^3}{[A][B]^2}$$

- Larger values of $K_c$ will indicate that the equilibrium lies more to the right.

# Reversible Reactions – Dynamic Equilibrium

## Acid Dissociation Constant, $K_a$

- Equilibrium constants can also be defined for the ionization of weak acids. This is known as the acid dissociation constant, $K_a$.

- Example: For the reaction, $HA \rightleftharpoons H^+ + A^-$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

- Larger values of $K_a$ indicate that the acid dissociates to a greater extent and is a stronger acid.

# Reversible Reactions – Dynamic Equilibrium

## Solubility Product Constant, $K_{sp}$

- Equilibrium constants can also be defined for substances dissolving in water, known as solubility product constant, $K_{sp}$.

- Example: For the reaction, $AB(s) \rightleftharpoons A^+(aq) + B^-(aq)$

$$K_{sp} = [A^+][B^-]$$

- Larger values of $K_{sp}$ indicate that the solid is more soluble due to having higher concentrations of ions.

# Reversible Reactions – Dynamic Equilibrium

Could I please have some *questions* to check my understanding?

# Reversible Reactions – Dynamic Equilibrium

1. Nitrogen reacts with oxygen in an equilibrium reaction.

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g) \quad \Delta H = +170 \text{ kJ/mol}$$

When the reaction is at equilibrium, which statement is correct?

**A** The concentration of nitrogen present will change with time.

**B** The forward and backward reactions are taking place at the same rate.

**C** The forward reaction releases heat energy.

**D** There are more molecules on the left hand side of the equation than on the right.

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# Reversible Reactions – Dynamic Equilibrium

2. In a closed flask, gases **Q** and **R** reach a dynamic equilibrium.

$$\mathbf{Q}(g) \rightleftharpoons 2\mathbf{R}(g) \qquad \Delta H \text{ is positive}$$

Which change will move equilibrium to the right?

**A** Adding a catalyst.

**B** Decreasing the temperature.

**C** Increasing the pressure.

**D** Increasing the volume of the reaction flask.

# Reversible Reactions – Dynamic Equilibrium

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# Reversible Reactions – Dynamic Equilibrium

**3.** Elements **X** and **Y** react together in a reversible reaction to form $\mathbf{XY_2}$.

$$\mathbf{X} + 2\mathbf{Y} \rightleftharpoons \mathbf{XY_2}$$

1.0 mol of **X** is mixed with 1.0 mol of **Y** and the mixture is left to react until an equilibrium position is reached.

Which statement(s) about this reaction is / are correct?

1 After the equilibrium position has been reached, the reaction stops.

2 At equilibrium there is more than 0.5 mol of X present.

3 At equilibrium there is less than 0.5 mol of $XY_2$ present.

**A** 1, 2 and 3
**B** 2 only
**C** 3 only
**D** 2 and 3 only

# Reversible Reactions – Dynamic Equilibrium

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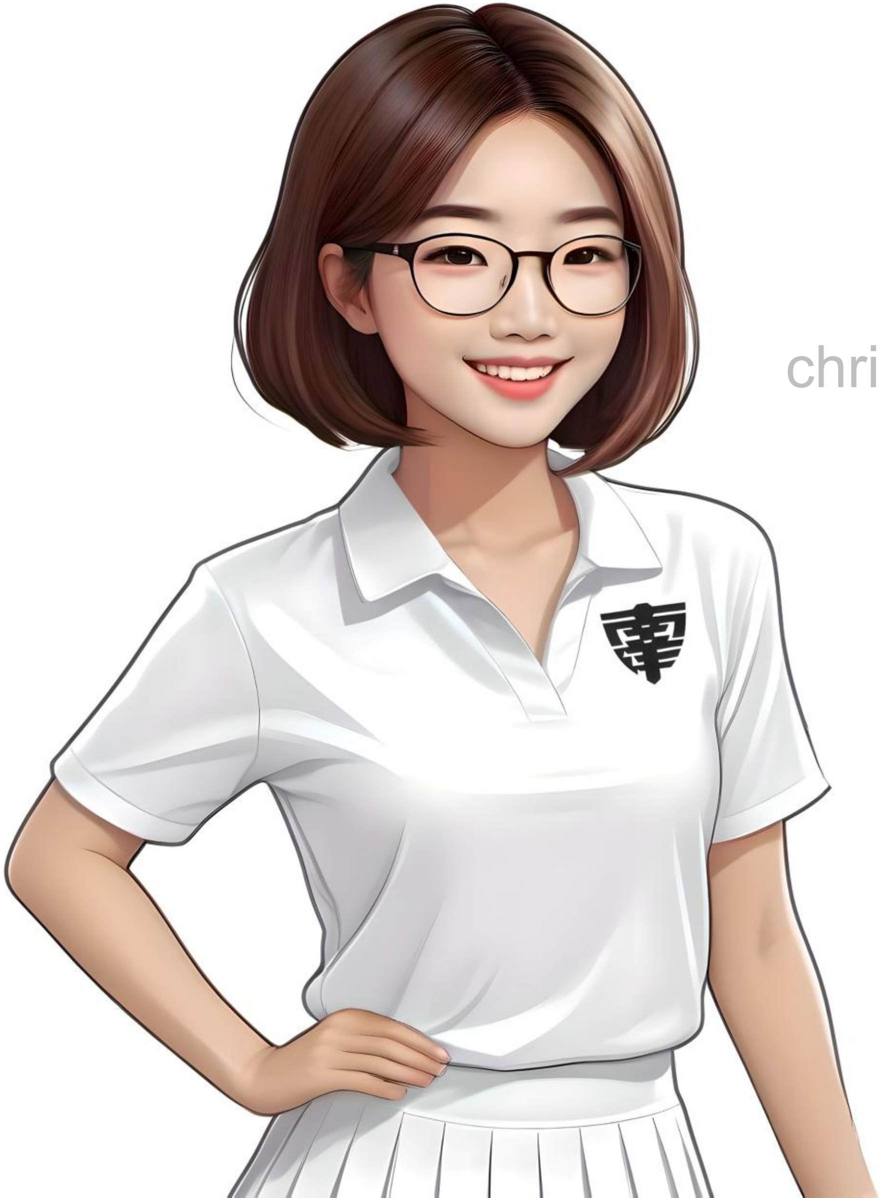
# Reversible Reactions – Dynamic Equilibrium

- 30 more multiple-choice questions on reversible reactions:

http://www.chemist.sg/ammonia_equilibrium/equilibrium_mcq_qu.pdf

# Reversible Reactions – Dynamic Equilibrium



Presentation on
Reversible Reactions
by Dr. Chris Slatter
christopher_john_slatter@nygh.edu.sg

Nanyang Girls' High School
2 Linden Drive
Singapore
288683

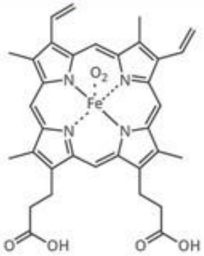
12th May 2024

# THE CHEMISTRY OF THE DIFFERENT COLOURS OF BLOOD

## Red

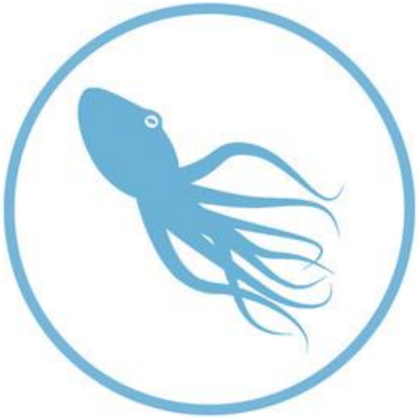
**HUMANS AND THE MAJORITY OF OTHER VERTEBRATES**

### HAEMOGLOBIN



HAEM B
*(oxygenated form)*

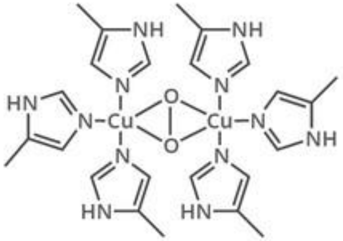
Haemoglobin is a protein found in blood, built up from subunits containing 'haems'. These haems contain iron, and their structure gives blood its red colour when oxygenated. Deoxygenated blood is a deep red colour - not blue!



## Blue

**SPIDERS, CRUSTACEANS, SOME MOLLUSCS, OCTOPUSES & SQUID**

### HAEMOCYANIN



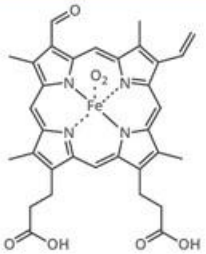
HAEMOCYANIN
*(oxygenated form)*

Unlike haemoglobin, which is bound to red blood cells, haemocyanin floats free in the blood. Haemocyanin contains copper instead of iron. When deoxygenated, the blood is colourless, but when oxygenated, it gives a blue colouration.

## Green

**SOME SEGMENTED WORMS, SOME LEECHES, & SOME MARINE WORMS**

### CHLOROCRUORIN



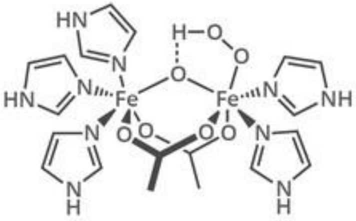
CHLOROCRUORIN
*(oxygenated form)*

Chemically similar to haemoglobin; the blood of some species contains both haemoglobin & chlorocruorin. Light green when deoxygenated, it is green when oxygenated, although when more concentrated it appears light red.

## Violet

**MARINE WORMS INCLUDING PEANUT WORMS, PENIS WORMS & BRACHIOPODS**

### HAEMERYTHRIN



HAEMERYTHRIN
*(oxygenated form)*

Haemerythrin is only 1/4 as efficient at oxygen transport when compared to haemoglobin. In the deoxygenated state, haemerythtin is colourless, but it imparts a violet-pink colour when oxygenated.