Unit-9

REVERSIBLE REACTIONS

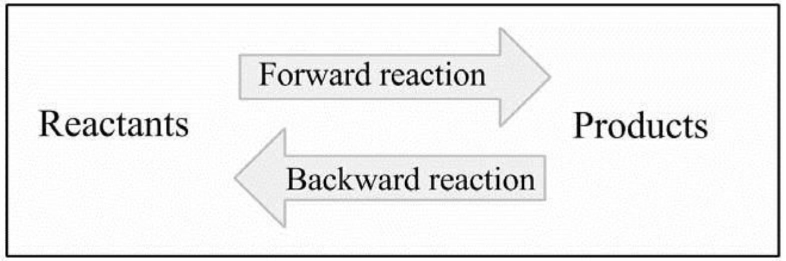
Reversible Reactions

In some chemical reactions, the **products** of the reaction can react together to produce the original **reactants**. These reactions are called **reversible reactions**. They can be represented in the following way:

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

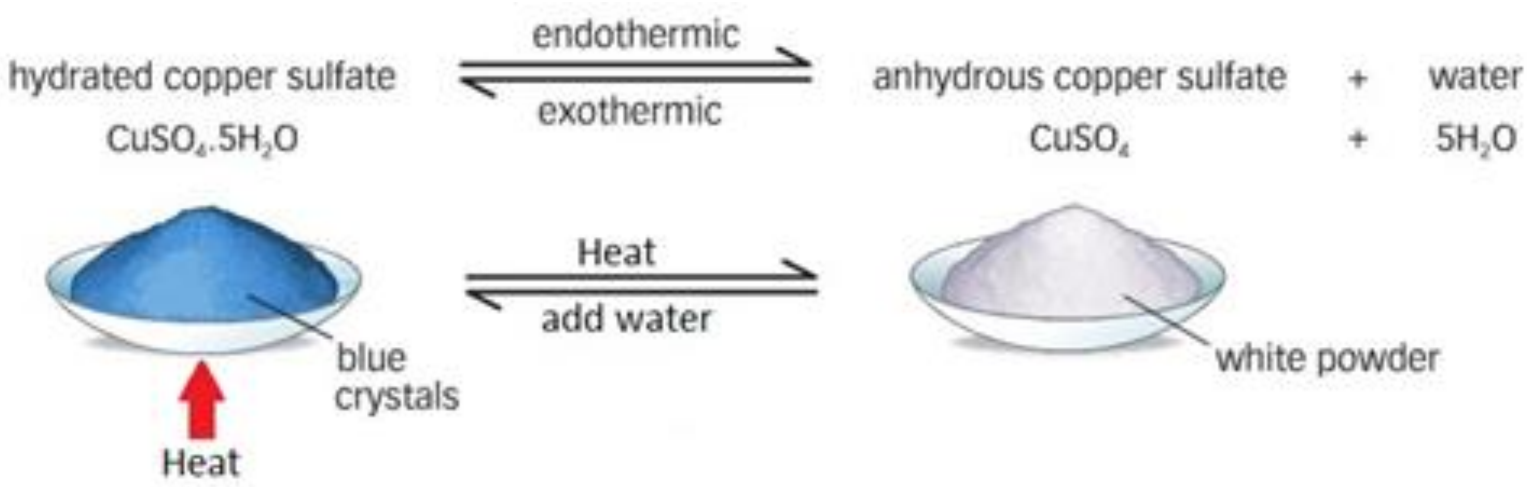
The symbol $\rightleftharpoons$ has two half arrowheads, one pointing in each direction. It is used in equations that show reversible reactions:

- the forward reaction is the one that goes to the right
- the backward reaction is the one that goes to the left



Chemical reactions that undergo reactions in both directions are reversible reactions.

Example for reversible reaction:



Forward reaction: The blue crystals of *hydrated copper(II) sulphate* on heating turn into a white powder. This is *anhydrous copper(II) sulfate*.

Backward reaction: The reaction is easy to reverse; on adding water, the *anhydrous copper(II) sulfate* turn into blue *hydrated copper(II) sulfate*.

A reversible reaction is endothermic in one direction, and exothermic in the other. The same amount of energy is transferred each time.

Some examples for reversible reactions:

N_2 Nitrogen	+	$3H_2$ Hydrogen	$\rightleftharpoons$	$2NH_3$ Ammonia		
N_2 Nitrogen	+	O_2 Oxygen	$\rightleftharpoons$	$2NO$ Nitric oxide		
		NH_4Cl Ammonium chloride	$\rightleftharpoons$	NH_3 Ammonia	+	HCl Hydrochloric acid
$2SO_2$ Sulphur dioxide	+	O_2 Oxygen	$\rightleftharpoons$	$2SO_3$ Sulphur trioxide		
H_2 Hydrogen	+	I_2 Iodine	$\rightleftharpoons$	$2HI$ Hydrogen iodide		
		PCl_5 Phosphorus pentachloride	$\rightleftharpoons$	PCl_3 Phosphorus trichloride	+	Cl_2 Chlorine

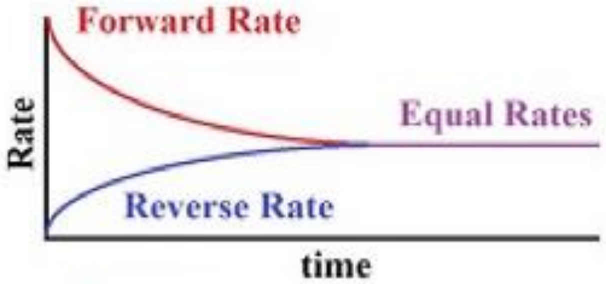
Chemical equilibrium:

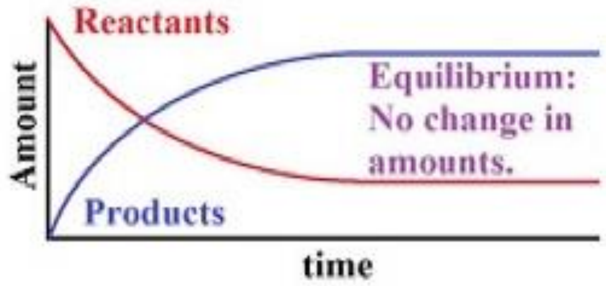
Reversible chemical process, where **concentrations** of reactants and products remain ***constant*** *over time* ***because*** *the* ***rates*** of the forward reaction ***equal*** *the reverse reaction.*

At equilibrium (t_e):

1. **Rate** of **forward** reaction **EQUALS** **rate** of **reverse** reaction – ***indefinitely.***

2. **Concentrations** of both **reactants** and **products** will be **CONSTANT** – *not **same**, but constant.*

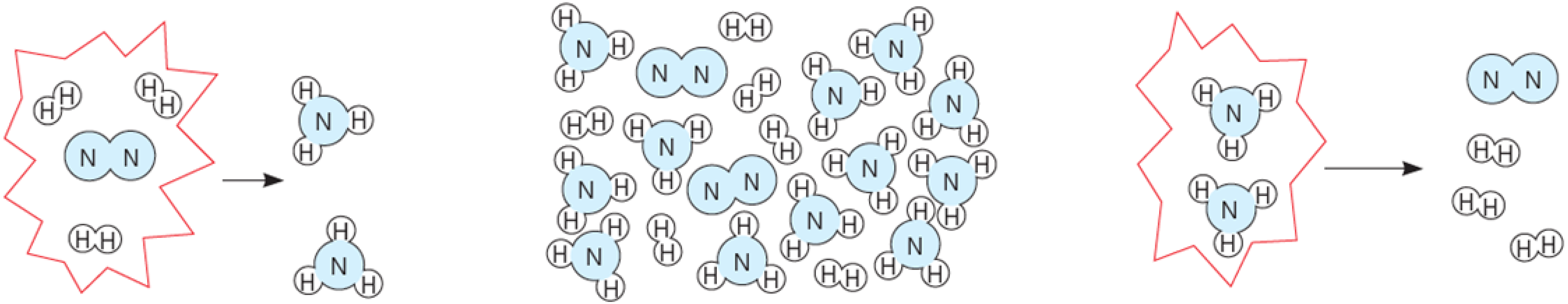




Reversible reactions and Equilibrium:

Reaction between Nitrogen and Hydrogen to make Ammonia is reversible.

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



In this example, Nitrogen reacts with Hydrogen to form ammonia. This is a reversible reaction. In a closed container, where nitrogen and hydrogen are mixed, 3 molecules of hydrogen react with 1 molecule of nitrogen to form 2 molecules of ammonia. Once a certain amount of ammonia is formed, the system reaches a state of equilibrium, from then on, every time, two ammonia molecules form, another two breakdown into nitrogen and hydrogen. So level of ammonia remains unchanged.

Equilibrium means there is *no overall change*.

The amounts of nitrogen, hydrogen and ammonia remain steady. But **dynamic** means there is continual change. Ammonia molecules continually break down, while new ones form.

In a closed system, a reversible reaction reaches a state of dynamic equilibrium, where the forward and back reactions take place at the same rate. So there is no overall change.

The term **dynamic equilibrium** is usually shortened to **equilibrium**.

Shifting the equilibrium

When a reversible reaction is in equilibrium and you make a change, the system acts to oppose the change, and restore equilibrium.

A reversible reaction *always* reaches equilibrium, in a closed system. But by changing conditions, you can *shift equilibrium, so* that the mixture contains more product

$$N_2 + 3H_2 \rightarrow 2NH_3$$

Exothermic
(gives out heat energy)

$$N_2 + 3H_2 \leftarrow 2NH_3$$

Endothermic
(takes in heat energy)

$$N_2 + 3H_2 \underset{\text{Endothermic}}{\overset{\text{Exothermic}}{\rightleftharpoons}} 2NH_3$$

Shifting Equilibrium – by changing concentration

$$N_2 + 3H_2 \underset{\text{Endothermic}}{\overset{\text{Exothermic}}{\rightleftharpoons}} 2NH_3$$

Change	How the equilibrium shifts
Increase in concentration	Equilibrium shifts to the **right** to reduce the effect of increase the concentration of a reactant
Decrease in concentration	Equilibrium shifts to the **left** to reduce the effect of a decrease in concentration of reactant

On increasing the concentration of Nitrogen and Hydrogen (that is reactants), the equilibrium will shift to the right, resulting in more ammonia (that is product) formation.

On decreasing the concentration of Nitrogen and Hydrogen (that is reactants), the equilibrium will shift to the left, resulting in more formation. That means, ammonia breaks down to nitrogen and hydrogen resulting less product formation.

Shifting Equilibrium – by changing temperature

$$N_2 + 3H_2 \underset{\text{Endothermic}}{\overset{\text{Exothermic}}{\rightleftharpoons}} 2NH_3$$

Change	How the equilibrium shifts
Increase in temperature	Equilibrium moves in the **endothermic** direction
Decrease in temperature	Equilibrium moves in the **exothermic** direction

Forward reaction in ammonia formation is exothermic (gives out heat), and the backward reaction is endothermic (absorbs heat).

To get more product (ammonia) the temperature should be kept low. In other words, if the temp is high, then backward reaction favours resulting in more ammonia converting into hydrogen and nitrogen.

Shifting Equilibrium – by changing pressure

Exothermic

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

Endothermic

1 molecule　　3 molecules　　　　2 molecules

Change	How the equilibrium shifts
Increase in pressure	Equilibrium shifts in the direction that produces the **smaller** number of molecules of gas
Decrease in pressure	Equilibrium shifts in the direction that produces the **larger** number of molecules of gas

By increasing the pressure, the equilibrium will be shifted towards less molecule side. By decreasing pressure, the equilibrium will be shifted to more molecule side.

In this example, reactants are 4 molecules (1-nitrogen & 3-hydrogen) and product is 2 molecules (2-ammonia). Hence, increase in the pressure results in more ammonia formation.

Effect of Catalyst on Equilibrium

$$N_2 + 3H_2 \underset{\text{Endothermic}}{\overset{\text{Exothermic}}{\rightleftharpoons}} 2NH_3$$

The presence of a catalyst does not affect the position of equilibrium, but it increases the rate at which the equilibrium is reached. This is because the catalyst increases the rate of both the forward and backward reactions by the same amount.

In manufacturing of Ammonia, iron is used as a catalyst. **A catalyst speeds up the forward and back reactions equally.** So the reaction reaches equilibrium faster, which saves you time. But *the amount of ammonia does not change*.

Choosing the optimum conditions:

In order to get best yield of Ammonia, it is best to:

- ✓ use high concentration of reactants (Nitrogen and Hydrogen),
- ✓ use high pressure,
- ✓ use a moderate (low) temperature,
- ✓ and a catalyst.