Unit 4, Lesson 01: Introduction to Reaction Rates

Reaction rate is a measure of how quickly or slowly the reactants are consumed, or the products are formed, in a chemical reaction.

- a reaction that occurs in a short period of time has a fast reaction rate, eg. explosion of TNT
- a reaction that occurs over a long period of time has a slow reaction rate, eg. oxidation of a copper roof

The study of reaction rates is called "**chemical kinetics**".

To determine reaction rate, we need to measure how the amount of a reactant or product changes over time

$$\text{Reaction Rate} = \frac{\text{change in quantity of reactant or product}}{\text{change in time}}$$

There are many different ways to measure reaction rate, depending on the physical properties of the reactants or products:

Physical Property	Change that is measured	Method or Device used to measure the change	Rate Expressed as:	Sample Units
if a product or reactant is coloured	change in the intensity of colour	carry out the reaction inside a spectrophotometer	$\frac{\Delta \text{ concentration}}{\Delta \text{ time}}$	$\frac{mol/L}{s}$ $\frac{mol/L}{min}$
if a product or reactant is a gas	volume of gas produced or reacted	collect gas in a gas tube and measure its volume	$\frac{\Delta \text{ volume}}{\Delta \text{ time}}$	$\frac{mL}{s}$ $\frac{L}{min}$
if a product or reactant is a gas	pressure of gas produced or reacted	measure pressure inside a sealed container	$\frac{\Delta \text{ pressure}}{\Delta \text{ time}}$	$\frac{kPa}{s}$ $\frac{atm}{min}$
if a product is a gas	mass of an open system	carry out the reaction on an electronic balance; mass will change as gas leaves	$\frac{\Delta \text{ mass}}{\Delta \text{ time}}$	$\frac{mg}{s}$ $\frac{g}{min}$
if a product or reactant is an ion in solution	conductivity of solution	use a conductivity meter to measure changes in the electrical conductivity of a solution	$\frac{\Delta \text{ concentration}}{\Delta \text{ time}}$	$\frac{mol/L}{s}$ $\frac{mol/L}{min}$
if a H+ or OH- ions in solution are produced or consumed	pH of solution	use a pH meter to measure changes in the pH of a solution	$\frac{\Delta \text{ conc'n}}{\Delta \text{ time}}$ $\frac{\Delta \text{ pH}}{\Delta \text{ time}}$	$\frac{mol/L}{s}$ $\frac{mol/L}{min}$

A fast reaction will have a high rate of reaction (a large change per unit time) while slower reactions have a low rate of reaction (a small change per unit time).

By convention, reaction rates are always positive values because they measure how fast the reaction is proceeding or moving forward. A negative sign for a reaction rate is ignored.

The most common method used to express reaction rates in high school chemistry is the change in the concentration of either a reactant or product. Molar concentration (C) is defined as the number of moles of solute per litre of solution. It is calculated using the equation:

$$C = \frac{n}{V}$$

where, n is the number of moles of solute, and
V is the volume of the solution in litres

Molar concentration can be expressed as: 3.0 mol/L = 3.0 mol·L^{-1} = 3.0 M = [A]

eg. [HCl] = 1.5 M = 1.5 mol/L = 1.5 mol·L^{-1}

[OH^{1-}] = 3.00 M = 3.00 mol/L = 3.0 mol·L^{-1}

Reaction rates are commonly expressed as the change in the concentration (in mol/L) of a reactant or product over time.

Pressures or volumes of gases measured under the same conditions can also be used because equal volumes of gases at STP contain equal numbers of particles and will exert equal pressures.

$$\text{reaction rate} = \frac{\Delta[\text{product or reactant}]}{\Delta t}$$

where Δ[product] or Δ[reactant] is the change in the concentration of a solution or pressure of a gas or volume of a gas, and Δt is the change in time

If we graph the molar concentration of a **product** over time, a typical graph has this shape:

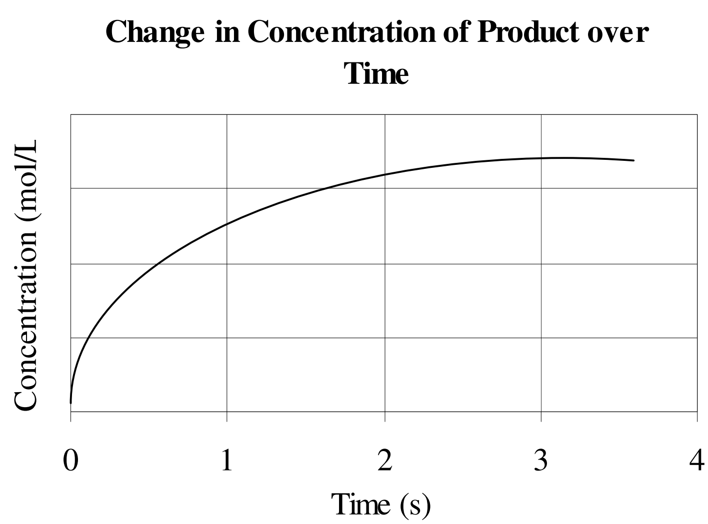
At the beginning of the reaction:

- concentration of the product is low
- product forms very quickly
- slope of line is steep
- reaction rate is fast

As reactant is consumed:

- the concentration of product increases more slowly
- slope of the line becomes less steep
- reaction rate slows down

Eventually, the maximum amount of product has formed, so the slope of the curve flattens out



If we graph the concentration of a **reactant** for the same reaction over time, we would see this shape:

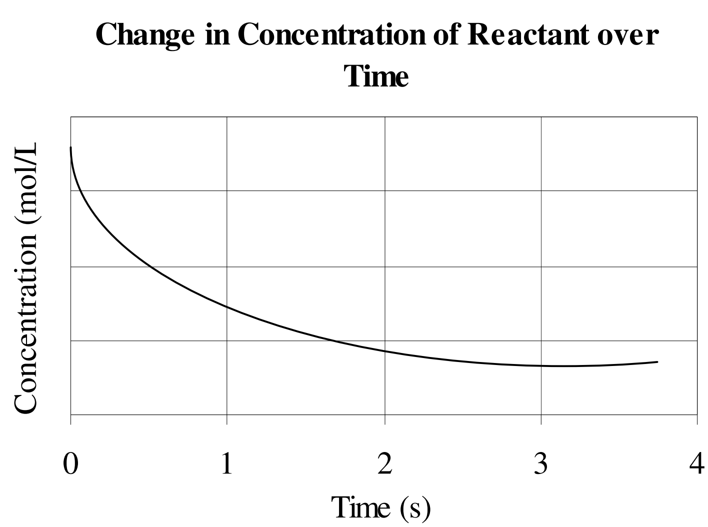
At the beginning of the reaction:

- concentration of the reactants are high
- reactant is used up quickly
- slope of line is steep
- reaction rate is fast

As reactant is consumed:

- the concentration of reactant decreases
- slope of the line becomes less steep
- reaction rate slows down

Eventually, the concentration of the reactant is so low that the slope of the curve flattens out

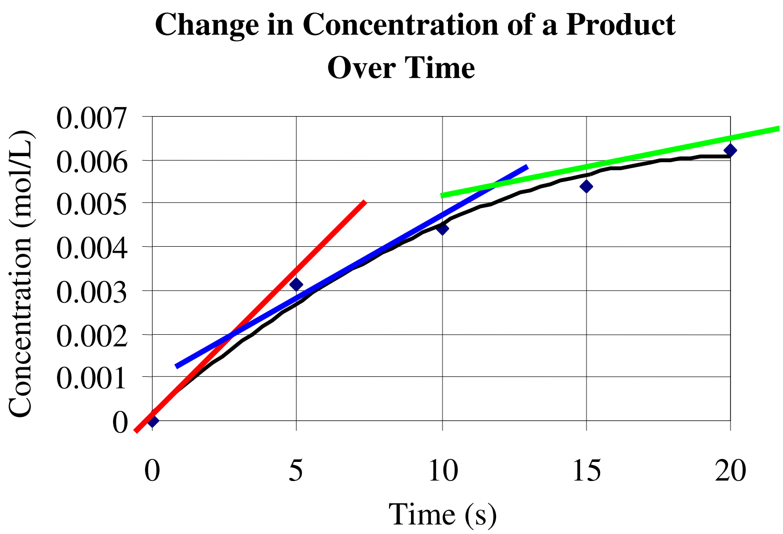


Numerically, reaction rate will change depending on the period of time you study. Common ways of expressing rates are:

1. **<u>Average Reaction Rate</u>** is the change in the quantity of a reactant or product over a specified period of time, for example, over the first 5.0 seconds.

   $$\text{average reaction rate} = \frac{\text{final conc'n of reactant or product} - \text{initial conc'n of reactant or product}}{\text{final time} - \text{initial time}}$$

2. **<u>Instantaneous Rate</u>** is the rate of the reaction at a single moment in time. It is calculated by finding the slope of the tangent to the line at a certain point.



tangent #1: early in the reaction the slope of the tangent is steep indicating a very rapid instantaneous reaction rate

tangent #2: about mid-way through the reaction, the slope of the tangent is flatter, indicating that the rate of the reaction is slowing down

tangent #3: late in the reaction, the slope of the tangent is very flat, indicating that the reaction rate is now very slow

3. **<u>Initial reaction rate</u>** is the rate at the very beginning of the reaction while the concentrations of the reactants are very close to their initial concentrations.

- the first tangent on the above graph represents the initial reaction rate
- initial reaction rate is the most useful for comparing the rates of different reactions

Reaction rate depends on the period of time being studied. It also depends on which reactant or product is being measured.

For example, for the reaction: $2\,NO(g) + O_2(g) \rightarrow 2\,NO_2(g)$

The initial rate of **disappearance** of NO is 8.0 mL/s.
For this same period of time, the rate of disappearance of O_2 would be 4.0 mL/s.

The molar coefficients in the balanced chemical reaction tell us that **two** moles of NO are used up for every **one** mole of O_2 that is used up. NO is used up twice as fast as O_2, so the reaction rate that is calculated using NO will be twice as fast as the reaction rate calculated using O_2.

Therefore, we need to specify the rate of a reaction in terms of which reactant or product is being measured. If we know the rate for one of the reactants or products, we can determine what the reaction rate would be if it was expressed in terms of a different reactant or product using the molar coefficients in the balanced chemical reaction.

eg. Ammonia is created by the Haber reaction:

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

If ammonia gas is produced at the rate of 1.60 mL/s (a change in volume of gas produced with time), how fast is hydrogen gas being consumed?

This can be solved using stoichiometry.

Step 1: Write out the balanced chemical equation and underline the species that we are interested in:

$$N_2\,(g) \quad + \quad \underline{3\ H_2\,(g)} \quad \rightarrow \quad \underline{2\ NH_3\,(g)}$$

Step 2: Set up the mole ratio underneath the reaction, using the molar coefficients in the reaction and any information you are given in the question. Use x to represent the value you need to find:

$$N_2\,(g) \quad + \quad \underline{3\ H_2\,(g)} \quad \rightarrow \quad \underline{2\ NH_3\,(g)}$$

$$\frac{3}{x} \quad = \quad \frac{2}{1.60}$$

Step 3: Cross multiply and solve for x:

$$2\,x = 3 \cdot 1.60$$

$$x = 2.40\ \text{mL/s}$$

So, hydrogen gas is consumed at the rate of 2.40 mL/s

Step 4: Sanity check: does this make sense? Yes- hydrogen is consumed at a faster rate than ammonia gas is produced.