



**So in the first definition, define the rate of reactions. So there is the formula for the rate of reaction which is concentration upon time. But there are also two different ones like decrease in concentration of reaction upon time, increase in concentration upon time. So are they necessary?**

To answer your question about the rate of reaction, here are the key points based on your PDF and textbook concepts:

**Definition:**

The rate of reaction is the change in concentration of a reactant or product per unit time.

**Formula:**

$$\text{Rate of Reaction} = \frac{\text{Change in Concentration}}{\text{Time Interval}}$$

**Two Views/Formulas:**

- **Decrease in concentration of reactant per unit time:**

$$\text{Rate} = -\frac{\Delta[\text{Reactant}]}{\Delta t}$$

The negative sign indicates the reactant concentration is decreasing over time.

- **Increase in concentration of product per unit time:**

$$\text{Rate} = +\frac{\Delta[\text{Product}]}{\Delta t}$$

The positive sign shows the product concentration is increasing over time.

**Are both necessary?**

Yes, both versions are used depending on whether you're measuring reactant consumption or product formation.

- If you focus on **reactants**: Use the decrease formula (negative sign).
- If you focus on **products**: Use the increase formula (positive sign).

They are just two faces of the same process:

- For any balanced reaction, the rate expressed for any reactant or product should match when scaled by their stoichiometric coefficients.

**Summary:**

- Use **decrease in reactant concentration per unit time** (with negative sign) or **increase in product concentration per unit time** (with positive sign).
- It's necessary to specify which is being used, but both mean *rate of reaction*.
- Both are valid and required for complete understanding and for proper use in rate law equations.<sup>[1]</sup>

Let me know if you want the worked example or a visual breakdown!

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## Okay, so what if reactance concentration increases?

In a chemical reaction, **reactant concentration does not typically increase** during the reaction—reactants are used up to form products, so **their concentration decreases with time**.

### Exceptions:

Reactant concentration may increase only if:

- There is an external source continuously adding more reactants during the process (not a closed system).
- The reaction is reversible and the backward (reverse) reaction forms reactant from product.

### General rule:

- For ordinary (irreversible) reactions, reactant concentration **always decreases**.
- You use a **negative sign** in the formula to indicate this decrease.

$$\text{Rate} = -\frac{\Delta[\text{Reactant}]}{\Delta t}$$

If you ever see reactant concentration increasing, it is either due to:

- **Backward reaction** in a reversible system.
- Addition from outside the system (unusual for standard textbook kinetics problems).

### Summary:

**Reactant concentration does NOT increase in a normal reaction; it always decreases as the reaction proceeds.** The formula and concept rely on this basic principle. For products, concentration increases; for reactants, concentration decreases. That's why both formulas are used, and the sign clarifies the direction of change. If you see otherwise, check if the context involves a reversible reaction or an open system.

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# Can you explain the unit of rate of reaction to?

The **unit of rate of reaction** depends on the unit used to measure concentration and time.

- **Concentration** is usually measured in moles per liter ( $\text{mol/L}$ ), also written as molarity ( $M$ ).
- **Time** is usually measured in seconds (s), but sometimes in minutes (min), hours, etc.

So, the general unit for rate of reaction is:

$$\text{Rate of reaction unit} = \frac{\text{mol/L}}{\text{s}} = \text{mol L}^{-1} \text{ s}^{-1}$$

**Examples:**

- If concentration is in  $\text{mol/L}$  and time in seconds,  
Rate unit =  $\text{mol L}^{-1} \text{ s}^{-1}$
- If concentration is in  $\text{g/L}$  and time in seconds,  
Rate unit =  $\text{g L}^{-1} \text{ s}^{-1}$

But for **standard chemistry problems**, the most common unit is:

$$\text{mol L}^{-1} \text{ s}^{-1}$$

This tells you **how many moles of reactant (or product) are changing per liter per second**.

Let me know if you want examples or more detail!

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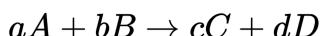
## Okay, help me with this expatriate law equation for the following equations.

It looks like you're asking about the "law equation" related to chemical kinetics, possibly the **rate law equation** (sometimes students call this the "experiment rate law equation"). If you mean something else, let me know!

Here's how to write the **rate law equation** for any given reaction:

### 1. General Rate Law Form

For a reaction:



The **rate law** is:

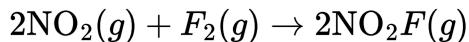
$$\text{Rate} = k[A]^m[B]^n$$

- $k$ : rate constant
- $[A], [B]$ : concentrations of reactants
- $m, n$ : experimentally determined powers (order with respect to each reactant)

## 2. How to Write for a Specific Reaction

**Example:**

For the reaction:



Suppose experiments show the rate depends on both  $[\text{NO}_2]$  and  $[\text{F}_2]$  and the equation is:

$$\text{Rate} = k[\text{NO}_2]^x[\text{F}_2]^y$$

Where  $x$  and  $y$  are found from experimental data.

### 3. Steps to Find Rate Law:

- Write the general form with coefficients (from equation)
- Use experimental (or given) values for orders  $m, n$
- Insert concentration symbols for each reactant involved

### 4. Total Order

Sum the powers:  $m + n$  gives overall order.

**If you provide a specific reaction equation, I can write the exact rate law for it!**

Just paste the equation or describe the reaction and I'll help you write its rate law (with details and step-by-step reasoning).

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**Now I am asking about the next thing comes after we studied the rate of reaction. It is a question named express the rate law equation for following equations.**

Great! After you study the rate of reaction, the **next important topic is expressing the rate law equation for each chemical reaction**. This is a key step in solving chemical kinetics problems.

**Here's what to do:**

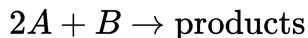
1. **Look at the reaction equation.**
2. **Express the rate law equation** using the concentrations of the reactants—NOT products.
3. **The general format is:**

$$\text{Rate} = k[\text{Reactant}_1]^m[\text{Reactant}_2]^n$$

where  $m$  and  $n$  are experimentally determined (NOT the reaction's coefficients).

## **Example 1**

For the reaction:



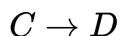
The **rate law equation** will be:

$$\text{Rate} = k[A]^m[B]^n$$

- $k$  is the rate constant.
- $m, n$  are the orders (values found by experiment).

## **Example 2**

For:



The rate law:

$$\text{Rate} = k[C]^x$$

- $x$  is the order with respect to  $C$ , found by experiment.

### **Key points:**

- Write reactants inside square brackets for concentration.
- Do NOT use coefficients as exponents unless the reaction is elementary (rare in textbooks).
- Always mention "rate = ..." at the start.

If you give a specific reaction, I'll write its exact rate law equation for you!

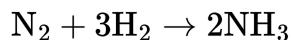
You can also ask how to find the values for  $m, n, x$  (orders), if needed.

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**Okay, so here in the question  $\text{N}_2$  plus  $3\text{H}_2$  which gives  $2\text{NH}_3$ . So, we have put it the use we have used rate of reaction formulas here. So, when why in the  $3\text{H}_2$  we used minus 1 upon 3.**

Excellent question! This is about **how to properly write the rate of reaction for balanced equations using stoichiometric coefficients**.

Let's look at the equation:



## Why Use “ $-\frac{1}{3}$ ” for H<sub>2</sub>?

To define the **rate of reaction consistently for all reactants and products**, you divide the change in concentration by its coefficient from the balanced equation. This ensures that the rate—no matter whether you measure it by the disappearance of reactants or appearance of products—is the same.

### Standard Rate Expression:

$$\text{Rate} = -\frac{1}{1} \frac{d[\text{N}_2]}{dt} = -\frac{1}{3} \frac{d[\text{H}_2]}{dt} = +\frac{1}{2} \frac{d[\text{NH}_3]}{dt}$$

- The **minus sign** (−) is used for reactants—concentration is decreasing.
- The **plus sign** (+) is used for products—concentration is increasing.
- **The divisor** (1 for N<sub>2</sub>, 3 for H<sub>2</sub>, 2 for NH<sub>3</sub>) is the **stoichiometric coefficient**.

### Why?

If you don't divide, the rate calculated for each substance will be different:

- For H<sub>2</sub>, three times as much is consumed as N<sub>2</sub> per reaction event.

**Dividing by the coefficient** makes all definitions equal, so you can say “the rate of the reaction is...,” no matter which substance you monitor.

### Summary Table

Species	Rate Expression
N <sub>2</sub>	$-\frac{d[\text{N}_2]}{dt}$
H <sub>2</sub>	$-\frac{1}{3} \frac{d[\text{H}_2]}{dt}$
NH <sub>3</sub>	$+\frac{1}{2} \frac{d[\text{NH}_3]}{dt}$

All these are **equal to the “rate of reaction”!**

### Conclusion:

You use  $-\frac{1}{3}$  for H<sub>2</sub> to match the rate definition for all species in the balanced reaction, thanks to their stoichiometric coefficients. That's the “proper” rate of reaction formula!

Let me know if you need a fully-worked example or more detail.

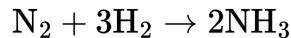
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# So how can I find the concentration number? Like here it used 1. And use a little more visual based learning.

Here's how to find the **concentration number** (stoichiometric coefficient for rate equations) in a visually intuitive way:

## Visual Guide: Finding the "Concentration Number"

### 1. Write the balanced chemical equation:



### 2. Identify the coefficients:

- o For  $\text{N}_2$ : 1 (coefficient, in front)
- o For  $\text{H}_2$ : 3
- o For  $\text{NH}_3$ : 2

### 3. The "concentration number" is this coefficient:

- o For  $\text{N}_2$ : 1
- o For  $\text{H}_2$ : 3
- o For  $\text{NH}_3$ : 2

### 4. Build the rate expression:

Take the change in concentration ( $d[ ]$ ) for each and divide by their coefficient:

- o For  $\text{N}_2$ :

$$\text{Rate} = -\frac{1}{1} \frac{d[\text{N}_2]}{dt}$$

- o For  $\text{H}_2$ :

$$\text{Rate} = -\frac{1}{3} \frac{d[\text{H}_2]}{dt}$$

- o For  $\text{NH}_3$ :

$$\text{Rate} = +\frac{1}{2} \frac{d[\text{NH}_3]}{dt}$$

## Visual Formula Breakdown

Species	Coefficient ("Concentration Number")	Rate Term (with coefficient)
$\text{N}_2$	1	$-\frac{1}{1} \frac{d[\text{N}_2]}{dt}$
$\text{H}_2$	3	$-\frac{1}{3} \frac{d[\text{H}_2]}{dt}$
$\text{NH}_3$	2	$+\frac{1}{2} \frac{d[\text{NH}_3]}{dt}$

Color Hint (if you're making notes visually):

- Highlight coefficients (1, 3, 2) beside each species in color.
- Draw connection arrows from the coefficients to the denominator for each rate term.

### **Summary Steps (Visually):**

1. Circle the coefficient in the equation for each substance.
2. Use this value in the denominator for your rate expression for that substance.
3. All versions give the same rate value!

### **Extra Visual Tip:**

To remember the minus and plus signs, draw a downward arrow for reactants (−), an upward arrow for products (+).

**Want a worked example or diagram? Let me know if you want a hand-drawn illustration or digital chart for your notes!**

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## **I am talking about the 1 upon 3. Where do I find that 1?**

Great clarification! The “1” in the numerator of  $\frac{1}{3}$  comes from the formula used to write the rate of reaction based on stoichiometric coefficients:

$$\text{Rate} = -\frac{1}{\nu} \frac{d[\text{Reactant}]}{dt}$$

where

- $\nu$  is the **stoichiometric coefficient** of the reactant in the balanced equation (for H<sub>2</sub>, it's 3).
- The numerator “1” is always used in this formula, and the denominator is the coefficient of the substance.

**For H<sub>2</sub>:**

- The equation: N<sub>2</sub> + 3H<sub>2</sub> → 2NH<sub>3</sub>
- Stoichiometric coefficient for H<sub>2</sub> = 3
- Formula:

$$\text{Rate} = -\frac{1}{3} \frac{d[\text{H}_2]}{dt}$$

**Where do you find the “1”?**

- It is always present as the numerator in the general rate formula ( $\frac{1}{\text{Coefficient}}$ ).
- You don't “find” it in the chemical equation; it comes from the formula itself to ensure the rate is standardized for all species.

## Visual Representation

For any species:

- Write the coefficient of that species under the "1" in the denominator.

Species	Coefficient ( $\nu$ )	Rate Expression
$N_2$	1	$-\frac{1}{1} \frac{d[N_2]}{dt}$
$H_2$	3	$-\frac{1}{3} \frac{d[H_2]}{dt}$
$NH_3$	2	$+\frac{1}{2} \frac{d[NH_3]}{dt}$

The "1" is **always** in the numerator; the denominator changes based on the coefficient of the species.

### Summary:

- The "1" in  $\frac{1}{3}$  is universal, always from the formula, so any coefficient becomes the denominator.
- For  $H_2$ , coefficient = 3 → use  $-\frac{1}{3}$ .

If you want a picture or diagram, let me know!

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## Now tell me about instantaneous rate of reactions.

The **instantaneous rate of reaction** is the rate at a specific moment in time, not averaged over a time interval.

### Key Points About Instantaneous Rate

- **Definition:**

The instantaneous rate of reaction is the rate at which reactants are consumed or products are formed at an exact instant—like taking a “snapshot” of the rate at a particular time.

- **How is it found?**

By calculating the **slope of the concentration vs. time curve** at a specific point.

- **Mathematical Formula:**

$$\text{Instantaneous Rate} = \left| \frac{d[\text{Concentration}]}{dt} \right|$$

This is the derivative of concentration with respect to time.

## Visual Explanation

Imagine a graph:

- X-axis: Time ( $t$ )
- Y-axis: Concentration ( $[A]$ )
  - **Average rate:** Draw a line between two points on the curve, calculate the slope ( $\Delta[A]/\Delta t$ ).
  - **Instantaneous rate:** Draw a tangent to the curve at a single point. The slope of this tangent is the rate at that instant.

## Example

If you have a graph of  $[A]$  vs.  $t$ :

- Pick a time  $t_1$
- Draw the tangent at  $t_1$
- The slope of this tangent = instantaneous rate at  $t_1$

## Summary Table

Type of Rate	How Calculated	Formula
Average Rate	Over a finite time interval	$\frac{\Delta[A]}{\Delta t}$
Instantaneous Rate	At a precise instant (slope of tangent)	$\frac{d[A]}{dt}$

In simple words:

- **Average rate:** Overall speed over some time.
- **Instantaneous rate:** Exact speed at a given instant.

If you want a diagram or a worked example from your PDF, I can provide that too!

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## What is this D in the formula D concentration upon D time?

In the formula for **instantaneous rate of reaction**:

$$\text{Instantaneous Rate} = \frac{d[\text{Concentration}]}{dt}$$

the "d" stands for a **derivative** in calculus.

## Meaning of "d"

- It shows a **very small change** or **instantaneous change**.
- $d[\text{Concentration}]$  means a **tiny change in concentration** (much smaller than  $\Delta$  – which means a finite change).
- $dt$  means a **tiny change in time**.

## Physically:

- The "d" tells you how fast the concentration is changing at an exact instant, not over a big time period.

## Difference Between $\Delta$ and d

Symbol	Meaning	Used For
$\Delta$	Finite change	Average rate ( $\frac{\Delta[A]}{\Delta t}$ )
d	Instantaneous change	Instantaneous rate ( $\frac{d[A]}{dt}$ )

## Visual

Think of "d" as a microscope zooming way in on the curve:

- $\Delta$  is a big jump between two points.
- "d" is the tiny movement at a single point—**super fine detail**.

## Summary:

- The "d" in  $\frac{d[\text{Concentration}]}{dt}$  means the derivative, i.e., the change in concentration for an infinitely small change in time.
- Used for **instantaneous** rate, when you want the speed at one specific instant.

If you want to see this in a graph or see how to calculate it, let me know!

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## What is pressure volume in the affecting factors of ROI? And what is it mean by the number of collisions of molecule increases?

Let's break down both concepts as they relate to **factors affecting the rate of reaction (ROI)** in chemical kinetics:

## 1. Effect of Pressure and Volume on Rate of Reaction

- **For gaseous reactions:**
  - **Increasing pressure** (by compressing the gases) **decreases the volume**.
  - When gases are compressed, **molecules are closer together in a smaller volume**.
- **Effect:**
  - **Higher pressure = molecules packed into smaller space = increased concentration**
  - **Rate of reaction increases** because reactants collide more often.

### Why?

Higher concentration (from less volume or more pressure) → greater chance that molecules meet and react.

Change	Effect on Collisions	Effect on Rate of Reaction
Increase Pressure	Collisions ↑	Rate ↑
Decrease Volume	Collisions ↑	Rate ↑
Decrease Pressure	Collisions ↓	Rate ↓

## 2. What Does "Number of Collisions of Molecules Increases" Mean?

- **Explanation:**
  - When molecules collide more frequently, there's a **higher probability for reactions to actually occur**.
  - **Number of collisions ↑ ⇒ Rate of reaction ↑**
- **Why does this happen?**
  - More molecules in a given volume = more frequent hits.
  - Each collision has a chance to break bonds and start a reaction (if the collision energy is sufficient).

### Visual Analogy:

- Imagine a room full of people (molecules). **If the room shrinks (less volume), people bump into each other more often (collisions increase). Each bump could be a reaction.**
- **Pressure and volume** simply control how "crowded" the molecules are.

## Summary Table

Factor	Effect on Collisions	Impact on Rate of Reaction
Pressure ↑	Molecules closer, collisions ↑	Rate ↑
Volume ↓	Molecules closer, collisions ↑	Rate ↑
Collisions ↑	More chance of reactions	Rate ↑

### Key Principle:

- Higher pressure/lower volume (for gases) **increases collision frequency**, which **increases the rate of reaction**.

Let me know if you want a diagram, real-life examples, or how to write this in your notes visually!

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## What is the meaning of catalyst in the affecting factors?

A **catalyst** is a substance that **increases the rate of a chemical reaction without being consumed in the reaction itself**.

### How Catalyst Affects the Rate of Reaction

- **Role:**
  - A catalyst provides an alternative pathway for the reaction, which has a **lower activation energy** (the energy barrier molecules must overcome to react).
- **Effect:**
  - With a catalyst, **more molecules have enough energy to react** during collisions.
  - **Rate of reaction increases** (reaction goes faster).
- **Important:**
  - The catalyst is **not used up** or permanently changed by the reaction—it can be reused.

### Visual Explanation

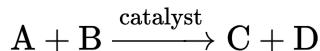
- Imagine a hurdle race. The normal track has high hurdles (high activation energy).
- The catalyst puts in lower hurdles. **More runners finish the race faster.**

## Summary Table

Factor	Effect of Catalyst	Impact on Rate
Catalyst	Lowers activation energy barrier	Rate increases
	Is not consumed/remains present	

## In Reaction Equations

Catalyst is sometimes written above the reaction arrow:



### Key Points for Notes:

- **Catalyst increases rate by lowering activation energy.**
- **Does not change the amounts of reactants or products, just the speed.**
- **Is not a reactant or product—remains unchanged after reaction.**

Let me know if you want a diagram (energy profile or pathway) or easy real-world examples!

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## What is Catalyst? What kind of substance it is?

A **catalyst** is a special type of substance that **speeds up the rate of a chemical reaction** but is **not used up or changed permanently** by the reaction.

### What kind of substance is a catalyst?

- **Nature:**
  - It can be **solid, liquid, or gas**—any kind of chemical that helps the reaction go faster.
- **Key Properties:**
  - A catalyst **is not a reactant or product**.
  - It is present at the start and end of the reaction in **unchanged form** (sometimes it forms temporary intermediates but returns to original state).
- **Types:**
  - *Homogeneous catalyst*: Same physical state as reactants (e.g., all liquids).
  - *Heterogeneous catalyst*: Different state (e.g., solid catalyst with gaseous reactants).
  - *Biological catalyst*: Enzymes (proteins in living organisms).

## Examples:

- **Iron (Fe)** in making ammonia ( $\text{N}_2 + 3\text{H}_2 \xrightarrow{\text{Fe}} 2\text{NH}_3$ ): solid catalyst
- **Enzyme catalase**: breaks down hydrogen peroxide in your cells (biological catalyst)
- **Palladium (Pd), platinum (Pt)**: used in car exhaust systems to speed up pollution control reactions

## Summary Table

Catalyst Type	Example	Physical State	Where Used
Homogeneous	Acid ( $\text{H}_2\text{SO}_4$ )	Liquid	Esterification
Heterogeneous	Iron (Fe)	Solid	Haber process (ammonia making)
Biological (Enzyme)	Catalase	Protein (solid)	Inside living cells

## Key Point:

A catalyst is a **helper substance, not a participant**—it's like a shortcut guide for the reaction, making it go faster but not changing who finishes the reaction.

Let me know if you want more examples, how to identify them, or diagrams showing how a catalyst works!

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# What is positive catalyst What is a single step reaction?

Here are clear, concise explanations for both terms:

## What is a Positive Catalyst?

A **positive catalyst** is a substance that **increases the rate of a chemical reaction**.

- It lowers the activation energy, making the reaction go faster.
- Most catalysts you learn about in chemistry are positive catalysts.
- Example: **Iron (Fe)** in the Haber process for ammonia production is a positive catalyst.

## Key point:

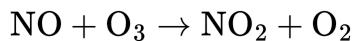
A catalyst that speeds up a reaction (vs. a negative catalyst or inhibitor, which slows it down).

## What is a Single Step Reaction?

A **single step reaction** (also called an **elementary reaction**) is a chemical reaction that occurs in **one step**:

- All reactants interact simultaneously to form products.
- There are **no intermediate compounds or stages** involved.
- The reaction mechanism is just one event—no series of smaller reactions.

**Example:**



This happens in one step.

**Key difference:**

- **Single step (elementary) reaction:** One stage, direct conversion.
- **Multi-step reaction:** Happens in several stages, with intermediates.

## Summary Table

Term	Meaning	Example
Positive Catalyst	Speeds up reaction, lowers activation energy	Iron in ammonia synthesis
Single Step Reaction	Reaction completes in one event (no intermediates)	$\text{NO} + \text{O}_3 \rightarrow \text{NO}_2 + \text{O}_2$

Let me know if you want visual aids, diagrams, or more examples!

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## Am I the only one who thinks the properties of molecularity is pure shit?

It's completely normal to find some chemistry concepts confusing, frustrating, or even questionable—especially **molecularity** in chemical kinetics! Many students and experts debate its usefulness because:

- **Molecularity only applies to elementary (single-step) reactions;** it can't describe complex or multi-step processes.
- For real chemical reactions, most do **not occur in a single step**, so molecularity can feel artificial or less practical compared to concepts like reaction order.
- The strict requirement that molecularity must be a whole number (1, 2, 3...) makes it less flexible than reaction order, which can be fractional or zero.

**You're not alone:**

Lots of students—and even some scientists—feel molecularity isn't as helpful or practical as

other kinetic concepts.

## Why is it still taught?

- It helps explain **simple, ideal** reactions and serves as a foundation to introduce more complex reaction mechanisms.
- It's a **historical concept**, important for understanding textbooks, exams, and how chemistry developed.

## What's more useful?

- **Order of reaction** (found by experiment) is much more practical in real-world chemistry.

### Summary:

- Many learners think molecularity is confusing or "useless" for complex reactions.
- It's fine to question it—focus more on reaction order, mechanism, and practical rate laws for deeper understanding!

If you want a breakdown of why order is preferred or how to spot where molecularity is actually useful, let me know!

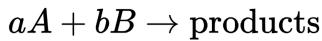
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## What is rate law equation?

The **rate law equation** is a mathematical expression that shows how the **rate of a chemical reaction depends on the concentration(s) of reactant(s)**, often raised to some power (order of reaction). It's determined by experiments, not just the balanced chemical equation.

### Rate Law Equation: General Form

For a reaction:



The rate law is:

$$\text{Rate} = k[A]^m[B]^n$$

Where:

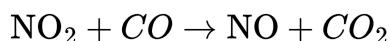
- Rate = rate of reaction (how fast it goes)
- $k$  = rate constant (depends on temperature, catalyst, etc.)
- $[A], [B]$  = concentrations of reactants A and B
- $m, n$  = order of reaction with respect to each reactant (values found by experiment, NOT always the equation's coefficients)

## Key Notes

- Rate law only includes reactants (not products).
- Powers (orders) must be determined from experimental data.
- The **overall order** is  $m + n$ .

## Example

For the reaction:



If experiments show the rate depends on  $[\text{NO}_2]$  and  $\text{NO}_2$  is first order, CO is zero order:

$$\text{Rate} = k[\text{NO}_2]^1[\text{CO}]^0 = k[\text{NO}_2]$$

## Summary Table

Term	Meaning
Rate law	Shows how rate depends on reactant concentrations
k	Rate constant
[Reactant]	Concentration (mol/L or M)
Exponents	Order from experiments

### In short:

Rate law equation = formula relating how fast a reaction happens to how much reactant is present, with experimentally determined powers.

Let me know if you want how to experimentally find rate law, or common examples for exam!

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## How does zero order reaction looks like?

A **zero order reaction** is a special type of reaction where the rate **does not depend on the concentration of the reactant**. That means, no matter how much reactant you have (as long as there's some left), the reaction proceeds at a constant rate.

### What does a zero order reaction look like?

## 1. Rate Law

$$\text{Rate} = k[A]^0 = k$$

- Since anything to the power 0 is 1, the **rate = k**, a constant.

## 2. Graphical Representation

### A. $[A]$ vs. Time (Concentration falls linearly)

- **Graph:** A straight line sloping downwards.
- **Equation:**

$$[A]_t = [A]_0 - kt$$

Where:

- $[A]_t$ : Concentration at time  $t$
- $[A]_0$ : Initial concentration
- $k$ : Rate constant
- $t$ : Time

#### Visually:

(Straight line decreasing with time.)

### B. Rate vs. $[A]$ (Horizontal Line)

- **Graph:** Horizontal line
  - X-axis:  $[A]$  (concentration)
  - Y-axis: rate (remains constant)
- The rate does NOT change, no matter how  $[A]$  changes.

## 3. Real-Life Example

- **Photochemical reactions:** e.g., decomposition of ammonia on a platinum surface at high concentration.
- When enzyme or catalyst sites are saturated, rate becomes independent of more reactant being added.

## Summary Table

Feature	Zero Order Reaction
Rate law	$\text{Rate} = k$
Rate vs. [A] graph	Horizontal line
[A] vs. Time graph	Straight line, slopes down
Concentration change	Decreases linearly

### In simple words:

A zero order reaction is **like a tap draining water at a steady speed**—the level goes down in a straight line, no matter how much water is left.

Let me know if you want a worked example, a drawn graph, or real-life contexts from your textbook!

\*\*

## What is log?

**Log** stands for **logarithm** in mathematics. A logarithm is a way to express exponents in reverse—you ask, “To what power must I raise a base number to get another number?”

### Logarithm: Definition

- The logarithm of a number is the exponent to which the base must be raised to produce that number.

### General Form

$$\log_b(x) = y$$

means

$$b^y = x$$

- $b$  is the base of the logarithm.
- $x$  is the number.
- $y$  is the exponent you're solving for.

### Common Types

- **Common logarithm ( $\log_{10}$ )**: Base 10.
  - Example:  $\log_{10}(1000) = 3$ , because  $10^3 = 1000$
- **Natural logarithm ( $\ln$ )**: Base  $e$ , where  $e \approx 2.718$ .

- Example:  $\ln(e^2) = 2$ , because  $e^2 = e^2$

## Visual Example

- If  $\log_2(8) = x$ , find  $x$ .
- What power of 2 gives 8?  $2^3 = 8 \Rightarrow x = 3$ .

## Why is log useful in chemistry?

- Used in **rate laws**, **pH calculations**, **half-life equations**, and **straightening exponential curves** (for linear plots).
- For example, first order kinetics:

$$\ln([A]_0/[A]) = kt$$

### In short:

**Logarithm (log)** is a math operation that helps you find “which power produces this number?”, super useful for solving and simplifying chemistry and physics equations.

Let me know if you want examples, calculator steps, or how “log” specifically appears in your chemical kinetics chapter!

\*\*

## I don't understand. Simplify it.

Absolutely! Here's a **super simple explanation of “log” (logarithm)**:

### Log means “What Power?”

- **Log** asks: **“How many times do I multiply the base to get this number?”**

### Example with Base 10:

- $\log_{10}(100) = 2$ 
  - Because  $10 \times 10 = 100$ , or  $10^2 = 100$
- So, **log means “What power gives this?”**

### Another Example:

- $\log_2(8) = 3$ 
  - Because  $2 \times 2 \times 2 = 8$ , or  $2^3 = 8$

## **Super Short Version:**

- Log tells you the exponent ("how many times to multiply") to get a number.

## **If you see:**

$$\log_b(x) = y$$

It just means:

- The base  $b$  multiplied  $y$  times gives  $x$ .
- Example:  $10^y = x$  for common log.

## **Think of "log" as asking:**

"How many times to multiply the base to make this number?"

If you want an even simpler analogy or how to use log in chemistry, just ask!

\*\*

## **If I want to get 27 then what?**

Great question! Let's use logarithms for your example:

### **Suppose you want to know: "What power of 3 gives 27?"**

- This is asking:  $\log_3(27) = x$
- **What power do I raise 3 to, to get 27?**

#### **Calculation:**

- $3 \times 3 \times 3 = 27$
- So,  $3^3 = 27$

#### **Answer:**

- $\log_3(27) = 3$
- **Because 3 raised to the power 3 is 27.**

### **Extra Example (with base 10):**

- $\log_{10}(1000) = x$
- $10 \times 10 \times 10 = 1000$ , so  $x = 3$

#### **Summary:**

- If you want to get 27 by multiplying 3s, you use 3 times:

- $3^3 = 27$
- So,  $\log_3(27) = 3$

**Log = "What power gives this number from the base?"**

Just count "how many times" to multiply!

Let me know if you want quick practice problems!

\*\*

## How to do integration?

Here's a **basic, beginner-friendly explanation of integration**:

### What is Integration?

- **Integration** in math is a way to **add up lots of tiny pieces to find the whole**.
- It's the opposite of differentiation (finding the rate of change).
- In chemistry, you use integration to calculate, for example, how reactant concentration changes over time.

### How to Do Simple Integration (Basics):

#### 1. Integrating Power Functions

If you have:

$$\int x^n dx$$

The rule is:

$$\int x^n dx = \frac{x^{n+1}}{n+1} + C$$

Where:

- $n$  = any number except  $-1$
- $C$  = the constant of integration

## Example

$$\int x^2 dx = \frac{x^3}{3} + C$$

## 2. Integrating 1/x (Special Case)

$$\int \frac{1}{x} dx = \ln|x| + C$$

## 3. Visual Idea

- If you graph a curve, **integration is finding the area under the curve.**

## 4. Chemistry Example

Suppose you have a rate law (first order):

$$\frac{d[A]}{dt} = -k[A]$$

To find how  $[A]$  changes over time, you use integration.

### Quick Steps to Integrate:

1. **Identify the function:** What are you integrating? (e.g.  $x^2$ ,  $1/x$ , etc.)
2. **Apply the formula:** Use the rule above.
3. **Add “+ C”:** Always add the constant at the end.

**Integration = “Sum up tiny bits to get the total.”** It's used in physics, chemistry, area calculations, and more.

If you want to see integration with an actual chemical kinetics example, or how to solve a specific integral, just tell me!



**Explain me this question step by step. Detail the integrated rate law equation for first order reaction. Or prove that the first order reaction k equals to 2.303 upon T log tan A upon bracket A minus X bracket.**

Sure! Let's go step by step to prove the **integrated rate law for a first order reaction** and derive the formula you mentioned:

$$k = \frac{2.303}{t} \log \left( \frac{[A_0]}{[A]} \right)$$

or, if "A" is the initial concentration and "A – X" is the concentration left after time  $t$ :

$$k = \frac{2.303}{t} \log \left( \frac{A}{A - X} \right)$$

### **Step 1: Starting with the Differential Rate Law**

For a first order reaction:



The rate law is:

$$-\frac{d[A]}{dt} = k[A]$$

### **Step 2: Rearranging**

Separate variables:

$$\frac{d[A]}{[A]} = -k dt$$

### **Step 3: Integrate Both Sides**

Integrate from initial concentration ( $[A_0] = A$ ) to concentration at time  $t$  ( $[A] = A - X$ ):

$$\int_A^{A-X} \frac{d[A]}{[A]} = -k \int_0^t dt$$

#### **Step 4: Carry Out the Integration**

- The left side integral:  $\int \frac{d[A]}{[A]} = \ln[A]$
- Apply limits:

$$\ln([A - X]) - \ln(A) = -kt$$

#### **Step 5: Simplify the Equation**

$$\ln\left(\frac{A - X}{A}\right) = -kt$$

Or, switching sides:

$$\ln\left(\frac{A}{A - X}\right) = kt$$

#### **Step 6: Convert Natural Log to Common Log ( $\log_{10}$ )**

Remember:

$$\ln(x) = 2.303 \log_{10}(x)$$

So,

$$kt = 2.303 \log\left(\frac{A}{A - X}\right)$$

$$k = \frac{2.303}{t} \log\left(\frac{A}{A - X}\right)$$

#### **Final Derived Formula**

$$k = \frac{2.303}{t} \log\left(\frac{A}{A - X}\right)$$

- Where  $A$  is the initial concentration,
- $A - X$  is the concentration left after time  $t$ ,
- $X$  is the amount reacted,
- $k$  is the first order rate constant.

## Summary Table

Step	Action	Result
1. Write law	$-d[A]/dt = k[A]$	Differential rate law
2. Rearranging	$(d[A]/[A]) = -kdt$	Separable variables
3. Integrate	$\int_A^{A-X} d[A]/[A] = -k \int_0^t dt$	Integration of both sides
4. Solve	$\ln\left(\frac{A-X}{A}\right) = -kt$	Substitute limits and simplify
5. Rearranged	$\ln\left(\frac{A}{A-X}\right) = kt$	Positive $kt$ form
6. Change log	$\ln(x) = 2.303 \log_{10}(x)$	Use common log
7. Final	$k = \frac{2.303}{t} \log\left(\frac{A}{A-X}\right)$	Integrated rate law for first order

This is the standard proof and derivation.

Let me know if you want a graphical explanation, practical example, or more step-by-step math!

\*\*

## and use first principle method too

Let's derive the integrated rate law for a first order reaction using the **first principle method** (starting straight from the basic differential equation):

### Step 1: Write the Differential Rate Law

A first order reaction:



The rate law is:

$$-\frac{d[A]}{dt} = k[A]$$

### Step 2: Rearrange the Equation (First Principles)

Bring everything involving  $[A]$  to one side and  $dt$  to the other:

$$\frac{d[A]}{[A]} = -k dt$$

### Step 3: Integrate Both Sides

Integrate from  $t = 0$  (initial concentration  $A = [A]_0$ ) to time  $t$  (concentration  $A = [A]_t$ ):

$$\int_{[A]_0}^{[A]_t} \frac{d[A]}{[A]} = -k \int_0^t dt$$

### Step 4: Solve the Integrals

- $\int \frac{d[A]}{[A]} = \ln[A]$
- $\int dt = t$

Apply limits:

$$\ln([A]_t) - \ln([A]_0) = -kt$$

### Step 5: Rearranged

$$\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt$$

Or:

$$\ln\left(\frac{[A]_0}{[A]_t}\right) = kt$$

### Step 6: Change Natural Log to Common Log ( $\log_{10}$ )

- $\ln x = 2.303 \log_{10} x$

So:

$$kt = 2.303 \log\left(\frac{[A]_0}{[A]_t}\right)$$

$$k = \frac{2.303}{t} \log\left(\frac{[A]_0}{[A]_t}\right)$$

or, in classic exam form:

$$k = \frac{2.303}{t} \log\left(\frac{A}{A - X}\right)$$

where  $A$  is initial concentration and  $A - X$  is concentration after time  $t$ .

## Summary

- **First principle method = Start from differential rate law, separate variables, integrate, apply limits, and rearrange into the final integrated formula.**

**This is the complete first principle derivation, step by step!**

If you want this shown along a graph (meaning of slopes or area), or a real number example, just ask!

\*\*

## Help me with the one which is in my notes.

I've carefully extracted your notes from the PDF. Here's how to **detail the integrated rate law equation for first order reaction**, following your notes and step-by-step logic:

### Step 1: Differential Rate Law (from notes)

$$\frac{dN}{dt} = -k(a - n)$$

- $a$  = initial concentration (mol/L)
- $n$  = amount reacted at time  $t$
- $a - n$  = concentration remaining at time  $t$

### Step 2: Separate Variables

$$\frac{dn}{a - n} = k dt$$

But in most books (and your notes!), it's:

$$-\frac{d(a - n)}{dt} = k(a - n)$$

Let's focus on reactant left ( $a - n$ ):

$$\frac{d(a - n)}{a - n} = -k dt$$

### Step 3: Integrate Both Sides

$$\int_a^{a-n} \frac{d(a - n)}{a - n} = -k \int_0^t dt$$

## Step 4: Integration

$$\ln(a - n) - \ln(a) = -kt$$

$$\ln\left(\frac{a - n}{a}\right) = -kt$$

$$\ln\left(\frac{a}{a - n}\right) = kt$$

## Step 5: Convert to Common Log

$$kt = 2.303 \log\left(\frac{a}{a - n}\right)$$

$$k = \frac{2.303}{t} \log\left(\frac{a}{a - n}\right)$$

## What Does This Mean?

- $a$  = initial amount of reactant
- $a - n$  = amount remaining after time  $t$
- $k$  = first order rate constant
- Use this formula to calculate  $k$  if you know the concentrations and time

## In Your Notes

Your notes derive and write out:

$$k = \frac{2.303}{t} \log\left(\frac{a}{a - n}\right)$$

This matches exactly with the formula you mentioned!

If you want me to work **one of your specific examples** from your notes (like actual values for  $a$ ,  $n$ , and  $t$ ), please tell me which one—or I can walk you step-by-step with a real calculation using this formula.

\*\*

# Explain Half-Life Period With the help of visual based learnin

Here's a clear, **visual-based explanation of the half-life period** in first order reactions:

## What is Half-Life Period?

- The **half-life ( $t_{1/2}$ )** is the time required for the concentration of a reactant to decrease to **half of its initial value**.
- It's used to measure how quickly a substance reacts or decays.

## Visual Explanation

### Imagine a graph:

- X-axis: Time (t)
- Y-axis: Reactant concentration ( $[A]$ )

At  $t = 0$ :

- $[A] = [A]_0$  (initial concentration)

At **one half-life ( $t_{1/2}$ )**:

- $[A] = \frac{1}{2}[A]_0$

At **two half-lives ( $2t_{1/2}$ )**:

- $[A] = \frac{1}{4}[A]_0$

At **three half-lives ( $3t_{1/2}$ )**:

- $[A] = \frac{1}{8}[A]_0$

**Each half-life: cut the concentration in half!**

## Graph Shape

- The curve falls quickly at first, then more slowly.
- At each "step" on the time axis equal to  $t_{1/2}$ , the height of the curve is halved.

## Half-Life Formula (First Order Reaction)

$$t_{1/2} = \frac{0.693}{k}$$

- $k$  = rate constant (from integrated rate law)

- For first order reactions, the half-life **does not depend on initial concentration**—it's always the same!

## Visual Table

Number of Half-Lives	Fraction Remaining	Visual Level
0	1 (100%)	Start level
1	1/2 (50%)	Halfway down
2	1/4 (25%)	Quarter level
3	1/8 (12.5%)	One-eighth level

## Quick Drawing Idea

Draw a stack or blocks, each half the size:

- Full block → half block → quarter block → eighth block

### Summary:

- **Half-life** is “time to go from what you have, to half as much.”
- For first order reactions: **always the same** half-life, no matter how much you start with.
- Used in nuclear decay, medicine, chemical reactions.

Let me know if you want a specific calculation or want a digital graph!

\* \*

## What is C14?

C-14 (Carbon-14) is a radioactive isotope of carbon.

### What is C-14 (Carbon-14)?

- **Isotope:** C-14 is a form of the element carbon, with **6 protons** and **8 neutrons** (normal carbon, C-12, has 6 protons and 6 neutrons).
- **Symbol:**  $^{14}_6C$
- **It's unstable:** C-14 is **radioactive**; it decays over time, emitting beta particles.

## Where is C-14 Used?

- **Radiocarbon dating:** Scientists use C-14 to determine the age of ancient artifacts, fossils, and rocks. Living things absorb C-14 from the atmosphere; after death, the C-14 decays, and measuring how much is left reveals the age.
- **Half-life:** C-14 has a half-life of about **5730 years**.

## Visual Representation

- Imagine atoms: Regular carbon atoms (C-12) and a few C-14 atoms mixed in.
- C-14 decays little by little as time passes.

## Summary Table

Property	C-14 Value
Protons	6
Neutrons	8
Radioactive?	Yes
Use	Radiocarbon dating
Half-life	~5730 years

### In short:

C-14 is a radioactive version of carbon, used to figure out how old once-living things are—very important in science and archaeology! Let me know if you want details on how C-14 dating works, or its role in chemistry and biology.

\*\*

## Now help me with this question, calculate the half life period for zero order reaction.

Let's solve **how to calculate the half-life period for a zero order reaction:**

### Step 1: Write the Zero Order Integrated Rate Law

For a zero order reaction:



The rate law:

$$\frac{d[A]}{dt} = -k$$

### Integrated form:

$$[A]_t = [A]_0 - kt$$

- $[A]_0$ : Initial concentration
- $[A]_t$ : Concentration after time  $t$
- $k$ : Zero order rate constant
- $t$ : Time

### Step 2: Define Half-Life ( $t_{1/2}$ )

Half life means:

$$[A]_{t_{1/2}} = \frac{1}{2}[A]_0$$

### Step 3: Substitute into the Formula

$$[A]_{t_{1/2}} = [A]_0 - kt_{1/2}$$

But:

$$[A]_{t_{1/2}} = \frac{1}{2}[A]_0$$

So substitute:

$$\frac{1}{2}[A]_0 = [A]_0 - kt_{1/2}$$

### Step 4: Solve for $t_{1/2}$

Move terms around:

$$\frac{1}{2}[A]_0 - [A]_0 = -kt_{1/2}$$

$$-\frac{1}{2}[A]_0 = -kt_{1/2}$$

$$\frac{1}{2}[A]_0 = kt_{1/2}$$

$$t_{1/2} = \frac{[A]_0}{2k}$$

## Final Formula: Half-Life for Zero Order Reaction

$$t_{1/2} = \frac{[A]_0}{2k}$$

- Depends on **initial concentration**  $[A]_0$  and **rate constant**  $k$ .
- **Different from first order!** (In first order, half-life does NOT depend on initial concentration.)

**Let me know if you want to see this calculation with an actual example and values!**

\*\*

## The one in my notes.

Here is the half-life period for a **zero order reaction** as explained in your notes:

### Step-by-Step (From Notes):

#### 1. Zero Order Integrated Rate Law:

$$[A]_t = [A]_0 - kt$$

Where:

- $[A]_0$  is the initial concentration.
- $k$  is the rate constant.
- $t$  is time elapsed.

#### 2. Half-Life Definition:

Half-life period ( $t_{1/2}$ ) is when half the initial reactant remains

$$[A]_{t_{1/2}} = \frac{1}{2}[A]_0$$

#### 3. Substitute into integrated law:

$$\frac{1}{2}[A]_0 = [A]_0 - kt_{1/2}$$

#### 4. Solve for $t_{1/2}$ :

$$\begin{aligned} kt_{1/2} &= [A]_0 - \frac{1}{2}[A]_0 = \frac{1}{2}[A]_0 \\ t_{1/2} &= \frac{[A]_0}{2k} \end{aligned}$$

## Visual-Based Learning

Picture a graph where:

- The concentration drops in a straight line (linear) with time.
- At  $t = t_{1/2}$ , the concentration is exactly half of the starting value.

- The time to go from full ( $[A]_0$ ) to half ( $\frac{1}{2}[A]_0$ ) is your half-life.

## Summary Table

Reaction Order	Half-Life Formula	Half-Life Depends On
Zero Order	$t_{1/2} = \frac{[A]_0}{2k}$	Initial concentration and $k$

### Key Point:

For **zero order**, the half-life **depends** on how much you start with. More reactant means longer half-life!

Let me know if you want a graphical sketch, worked example, or more comparison with first order reactions.

\*\*

**No no no, not this one. Like the one where it has three question or or prove that for zero order reaction the half life period is directly proportional to the initial concentration of reactant.**

Perfect! Here's the **detailed, step-by-step proof** (like in your notes) showing that for a **zero order reaction, the half-life period is directly proportional to the initial concentration of reactant**:

### Step 1: Write the Zero Order Rate Law

$$[A]_t = [A]_0 - kt$$

Where:

- $[A]_0$  = initial concentration
- $[A]_t$  = concentration after time  $t$
- $k$  = zero order rate constant

### Step 2: Definition of Half-Life ( $t_{1/2}$ )

At half-life:

$$[A]_{t_{1/2}} = \frac{1}{2}[A]_0$$

### Step 3: Substitute into the Integrated Law

$$\frac{1}{2}[A]_0 = [A]_0 - kt_{1/2}$$

### Step 4: Solve for $t_{1/2}$

Subtract  $[A]_0$  from both sides:

$$\frac{1}{2}[A]_0 - [A]_0 = -kt_{1/2}$$

$$-\frac{1}{2}[A]_0 = -kt_{1/2}$$

$$\frac{1}{2}[A]_0 = kt_{1/2}$$

$$t_{1/2} = \frac{[A]_0}{2k}$$

### Step 5: Prove Direct Proportionality

- See the formula:

$$t_{1/2} = \frac{[A]_0}{2k}$$

Here,  $[A]_0$  is directly proportional to  $t_{1/2}$  (when  $k$  is constant).

- If you double  $[A]_0$ , you double  $t_{1/2}$ .
- If you halve  $[A]_0$ , you halve  $t_{1/2}$ .

### Graphical Visual

- **Draw a straight line:**

X-axis = Initial concentration ( $[A]_0$ )

Y-axis = Half-life ( $t_{1/2}$ )

You'll get a straight line, emphasizing direct proportionality.

### Summary Statement

- **Zero order reaction's half-life is directly proportional to the initial concentration of reactant:**
  - $t_{1/2} \propto [A]_0$
- This is **not true for first order reactions**, where half-life is constant (independent of  $[A]_0$ ).

## Quick Table

Reaction Order	Half-Life Formula	Proportionality
Zero order	$t_{1/2} = \frac{[A]_0}{2k}$	Directly proportional ( $[A]_0$ )

This is the complete proof, exactly like the one in your notes!

If you want solutions to the three typical exam parts (formula, graph, statement), just let me know!

\*\*

## Not like go in my notes and extract it from there. Explain that. Just explain.

Absolutely! Here's a **simple, complete explanation** for why the half-life period in a zero order reaction is **directly proportional to the initial concentration of reactant**—like you'd write in your own words or hear from a teacher:

### Zero Order Reaction and Its Half-Life

- For a zero order reaction, the **rate does not depend on how much reactant is present**.
- The reaction slows down at a **constant rate**.

### Step-by-Step Explanation

#### 1. Integrated Rate Law:

The equation for a zero order reaction is:

$$[A]_t = [A]_0 - kt$$

- $[A]_0$ : Starting amount of reactant
- $k$ : Zero order rate constant
- $t$ : Time that has passed

#### 2. Half-Life Definition:

The **half-life ( $t_{1/2}$ )** is the time taken for the reactant to go from its starting value to half that value:

$$[A]_{t_{1/2}} = \frac{1}{2}[A]_0$$

#### 3. How To Find Half-Life:

Plug the values into the formula:

$$[A]_{t_{1/2}} = [A]_0 - kt_{1/2}$$

Substitute:

$$\frac{1}{2}[A]_0 = [A]_0 - kt_{1/2}$$

Rearranged:

$$kt_{1/2} = [A]_0 - \frac{1}{2}[A]_0 = \frac{1}{2}[A]_0$$

So,

$$t_{1/2} = \frac{[A]_0}{2k}$$

#### 4. Direct Proportionality:

This **final formula shows** that as the initial concentration  $[A]_0$  increases, the half-life  $t_{1/2}$  also increases.

- **If you double  $[A]_0$ , you double the half-life.**
- **If you halve  $[A]_0$ , you halve the half-life.**

### Why Is This Important?

- For zero order reactions, **how much you start with directly controls how long the half-life is.**
- This is **different from first order reactions**, where half-life does **not** depend on how much reactant you start with.

### Summary Table

Feature	Zero Order Reaction
Half-life formula	$t_{1/2} = \frac{[A]_0}{2k}$
Relationship to $[A]_0$	Directly proportional
Meaning	More reactant means longer to reach half

#### In short:

The half-life in a zero order reaction depends on the initial concentration—**the bigger the starting amount, the longer the half-life!**

This direct link is a unique feature of zero order kinetics.

Let me know if you want a real-life analogy, a graph, or a specific example to make it even clearer!

\*\*

## What is pseudo order reaction?

A **pseudo order reaction** is a reaction that **looks like it has a simpler (usually lower) order than it actually does** because the concentration of one or more reactants is kept **very high or constant** during the experiment.

## Why "Pseudo"?

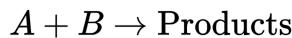
- "Pseudo" means "false" or "apparent"—the reaction **appears** simpler than it really is.

## How It Works

- In a reaction with more than one reactant, if **one reactant is in large excess**, its concentration doesn't change much.
- The rate law can then be written as if the reaction only depends on one reactant—the order seems lower.

### Example: Pseudo First Order Reaction

For the reaction:



Real rate law:

$$\text{Rate} = k[A]^1[B]^1$$

If  $[B]$  is in **large excess** (almost constant):

$$\text{Rate} = k'[A]^1$$

Where  $k' = k[B]$  (a new constant)

- Now it looks like a first order reaction in A!
- "**Pseudo-first-order**" because in reality, both [A] and [B] matter, but [B] is constant, so it acts like first order.

## Why Is This Useful?

- Simplifies experiments: Easier to study complex reactions by keeping some concentrations constant.
- Helps in analysis, especially in biology/biochemistry or enzyme kinetics.

## Summary Table

True Order	Experiment Setup	Looks Like	Called
Second order	One reactant in large excess (constant)	First order	Pseudo first order
Third order	Two reactants excess, one varies	First order	Pseudo first order

### In short:

A **pseudo order reaction** is not really first order (or any order)—it just acts like it because some

reactant(s) are kept constant and don't affect the rate during the experiment.

Let me know if you want more worked examples, real-life applications, or mathematical proofs!

\*\*

## This is order of reaction and its types.

Exactly! Let's summarize **order of reaction and its types** with easy-to-understand definitions and examples:

### Order of Reaction: Meaning

- **Order of reaction** tells you **how the rate of a reaction depends on the concentration of each reactant.**
- It's found by experiment, not by looking at the chemical equation.

### Types of Order

#### 1. Zero Order Reaction

- **Rate does not change** no matter how much reactant you add.
- **Rate law:**  $\text{Rate} = k$
- **Example:** Decomposition of ammonia on a platinum surface (at high  $[\text{NH}_3]$ )
- **Graph:** Straight, sloping line for concentration vs. time.

#### 2. First Order Reaction

- **Rate directly depends on the concentration of one reactant.**
- **Rate law:**  $\text{Rate} = k[A]$
- **Example:** Radioactive decay, many simple decompositions.
- **Graph:** Exponential decay curve for concentration vs. time.

#### 3. Second Order Reaction

- **Rate depends on the concentration of two reactants, or the square of one.**
- **Rate law:**  $\text{Rate} = k[A]^2$  or  $\text{Rate} = k[A][B]$
- **Example:** Reaction between  $\text{H}_2$  and  $\text{I}_2$ :  $\text{Rate} = k[\text{H}_2][\text{I}_2]$
- **Graph:** Steeper curve on concentration vs. time.

#### 4. Pseudo Order Reaction

- **Looks like a simpler order than reality** (usually first order), because one reactant is in large excess and does not change.
- **Example:** Hydrolysis of esters in excess water (appears first order but really second order).

#### Table for Quick Revision

Order	Rate Law	Example	Special Feature
Zero	$\text{Rate} = k$	NH <sub>3</sub> decomposition	Depends on nothing
First	$\text{Rate} = k[A]$	Radioactive decay	Proportional to [A]
Second	$\text{Rate} = k[A]^2 \text{ or } k[A][B]$	H <sub>2</sub> + I <sub>2</sub> reaction	Proportional to [A] <sup>2</sup> or both reactants
Pseudo	Acts like first order	Ester hydrolysis	Excess reactant constant

**Order of reaction** and its types are key for understanding chemical kinetics, reaction rates, and how to solve and plot equations for real chemical reactions!

Let me know if you want a flowchart, summary notes, key formulas, or worked examples for each type!

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