Rate Law

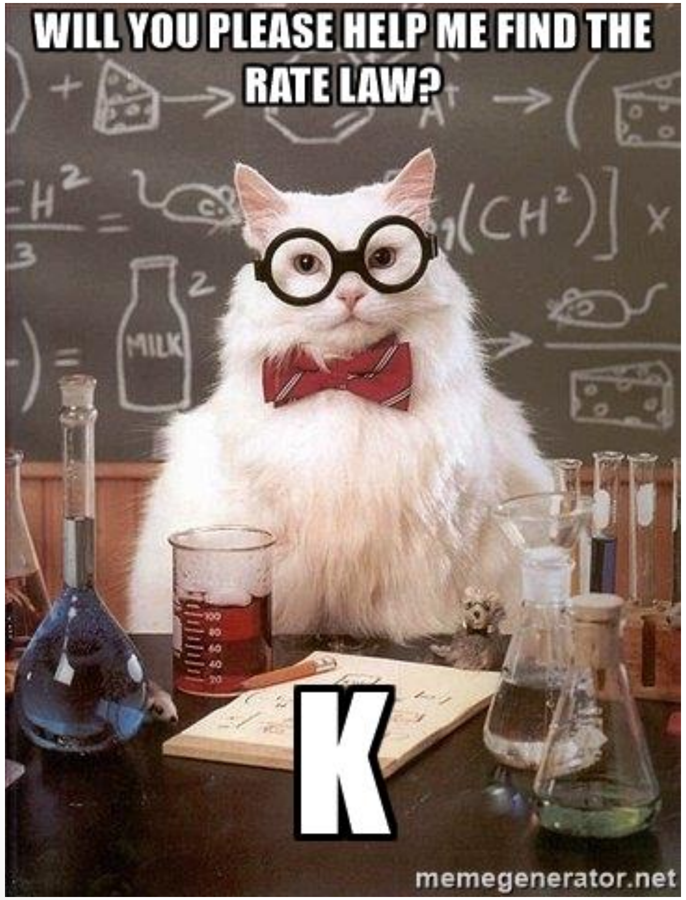
https://www.pinterest.com/pin/442267625876101351/

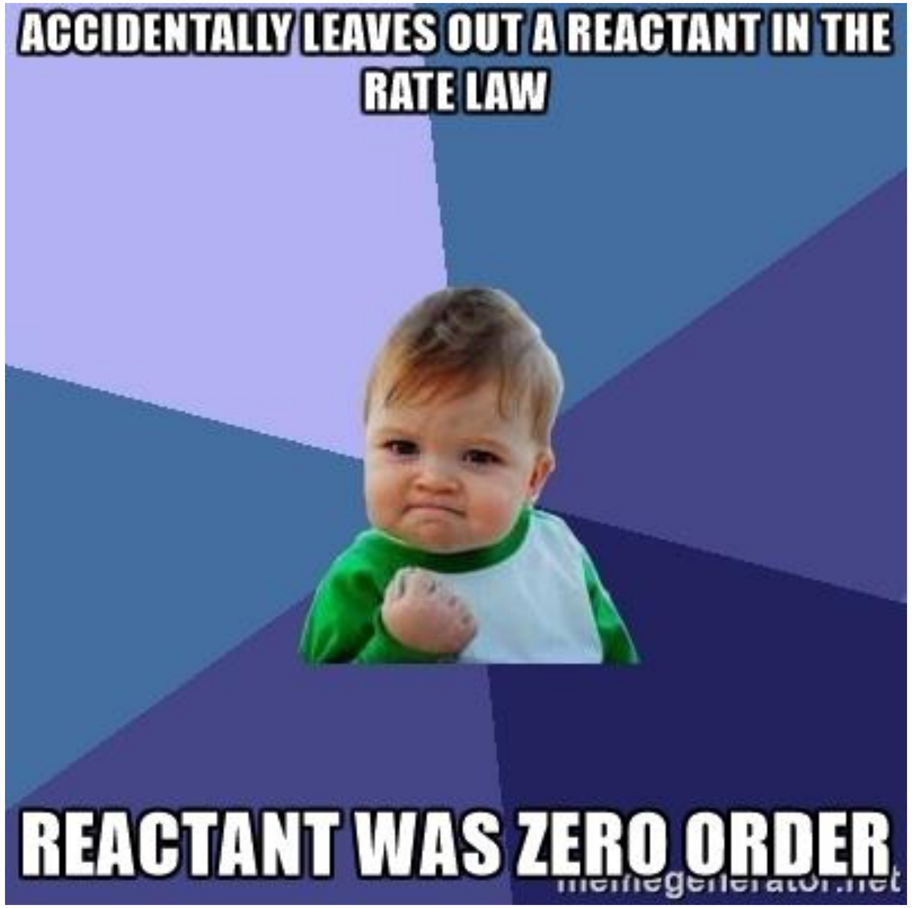
Outcomes:

- Determine the rate law of a chemical reaction from experimental data.

Rate Law:

- An **EXPRESSION** which relates the **RATE** of a reaction to the **CONCENTRATION** of the reactants.
- Allows us to **CALCULATE** the **RATE** of a reaction with given **CONCENTRATIONS** of **REACTANTS**.
- Shows the **QUANTITATIVE** effect of **CONCENTRATION** on reaction rate.





Rate Law:

Example:

For the reaction: ***A* → *Products***

$$Avg.Rate = \frac{\Delta[A]}{\Delta t}$$

The rate of consumption of ***A*** is **<u>DIRECTLY</u> <u>PROPORTIONAL</u>** to its **<u>CONCENTRATION</u>**.

- i.e. the **<u>FASTER</u>** ***A*** is **<u>CONSUMED</u>**, the **<u>LOWER</u>** its **<u>CONCENTRATION</u>**.

This is represented by the equation:

$$Rate = k[A]^x$$

Where:

k is the **<u>RATE</u> <u>CONSTANT</u>**

[A] is the **<u>CONCENTRATION</u>** of ***A***

x is a power, called the **<u>ORDER OF THE REACTION</u>**

Rate Law:

The Rate Constant (k):

- Can only be determined **EXPERIMENTALLY**.

- Is **SPECIFIC** for each **REACTION** at a specific **TEMPERATURE**, since its value depends on the **SIZE**, **SPEED** and **TYPE** of molecules.

- Changing **TEMPERATURE** changes **SPEED** of reactant molecules, which changes **RATE CONSTANT**.

- **TEMPERATURE** is the **ONLY FACTOR** that **AFFECTS** the rate **CONSTANT**

Order of Reaction:

$A \rightarrow B$

$RATE = k[A]^x$

Reaction Order:

Indicates how the **CONCENTRATION** of reactants **AFFECTS** the **RATE** of the reaction. Can only be determined **EXPERIMENTALLY**.

$Rate = k[A]^1$ $\quad$ $Rate = k(1)^1 = 1k$ $\quad$ $Rate = k(2)^1 = 2k$

First Order Reactions (X = 1)

- The reaction rate is **DIRECTLY PROPORTIONAL** to changes in reactant **CONCENTRATION**.
- If reactant concentration were:
  - **DOUBLED**, the rate would **DOUBLE**
  - **TRIPLED**, the rate would **TRIPLE**, etc

$RATE = k(3)^1 = 3k$

Order of Reaction:

$RATE = K[A]^2$

$RATE = K(1)^2 = 1K$ $\xrightarrow{\times 2}$ $RATE = K(2)^2 = 4K$ $\xrightarrow{\times 4}$

Second Order Reactions (<u>X = 2</u>)

- **<u>DOUBLING</u>** the concentration would **<u>QUADRUPLE</u>** the rate.
  - i.e. $\boldsymbol{2^x = 2^2 = 4}$
- **<u>TRIPLING</u>** the concentration would **<u>INCREASE</u>** the rate by a **<u>FACTOR</u>** of **<u>9</u>**:
  - i.e. $\boldsymbol{3^x = 3^2 = 9}$

$Rate = K(3)^2 = 9K$

$RATE = K[A]^0 = K(1 \times 10^{23})^0 = K$

Zero Order Reaction (<u>X = 0</u>)

- Here, a change in **<u>CONCENTRATION</u>** will **<u>NOT AFFECT</u>** the **<u>RATE</u>** of the reaction.
  - i.e. $\boldsymbol{2^x = 2^0 = 1}$

Order of Reaction:

	Order of Reaction (exponent)			
Conc. Change	**0**	**1**	**2**	**3**
x1	$1^0 = 1$	$1^1 = 1$	$1^2 = 1$	$1^3 = 1$
x2 (doubling)	$2^0 = 1$	$2^1 = 2$	$2^2 = 4$	$2^3 = 8$
x3 (tripling)	$3^0 = 1$	$3^1 = 3$	$3^2 = 9$	$3^3 = 27$

Rate law & Order of Reaction:

Example:

For a reaction with more than one reactant:

A + B → Products

The rate law would be:

$$Rate = k[A]^x[B]^y$$

The rate depends on both ***[A] and [B]***. Each reactant can affect the rate **DIFFERENTLY**.

The ***total order of the reaction*** is the **SUM** of the **EXPONENTS** in the **RATE LAW**.

Determining Rate Law:

Determining Rate Law – Initial Rate Method:

- Can only be done **EXPERIMENTALLY**.
- Even though rate does depend on stoichiometry to an extent, the rate law cannot be determined from **COEFFICIENTS**.
- We can **MEASURE** the effect on changes in concentration of one reactant on the rate by keeping the other reactant concentration **CONSTANT**.

Determining Rate Law:

Example:

What is the rate law for the following reaction, given the following experimental data?

$$H_2O_2 + 2HI \rightarrow 2H_2O + I_2$$

Trial Number	$[H_2O_2]$ (mol/L)	[HI] (mol/L)	Rate (mol/L·s)
1	0.10	0.10	0.0076
2	0.10	0.20	0.0152
3	0.20	0.10	0.0152

×1 ×2 ×2 ×2 ×2 $2^1 = 2$

$$\boxed{\text{RATE} = K[H_2O_2]^1[HI]^1}$$

Trials 1 & 3, [HI] = const, $[H_2O_2] \times 2$, Rate ×2

$2^1 = 2$ ∴ 1st order in $[H_2O_2]$

Trials 1 & 2, $[H_2O_2]$ = const, [HI] ×2, Rate ×2

$2^1 = 2$ ∴ 1st order in [HI]

Determining Rate Law:

Example:

Find the rate law for the reaction: $\boldsymbol{2NO_{2(g)} + F_{2(g)} \rightarrow 2NO_2F_{(g)}}$

[NO2] (mol/L)	[F2] (mol/L)	Rate (mol/L·s)
0.0010	0.0010	0.4
0.0010	0.0030	1.2
0.0050	0.0030	6.0
0.0050	0.0050	10

$5^x = 5$

$x = 1$

$3^x = 3$

$x = 1$

$$\text{Rate} = k[NO_2]^1[F_2]^1$$

Determining Rate Law:

Try this one…

When the reaction $\boldsymbol{CH_3Cl_{(g)} + H_2O_{(g)} \rightarrow CH_3OH_{(g)} + HCl_{(g)}}$ was studied, the following data was obtained:

3rd order

Trial	$[CH_3Cl]$ (mol/L)	$[H_2O]$ (mol/L)	Initial Rate (mol/L·s)
1	0.100	0.100	0.182
2	0.200	0.100	1.45
3	0.200	0.400	5.81

$\times 2$ $\times 1$ $\times 8$

$\times 1$ $\times 4$ $\times 4$

$2^x = 8$

$x = 3$

$4^x = 4$

$x = 1$

Determine the rate law for this reaction…

$$\text{RATE} = k[CH_3Cl]^3[H_2O]^1$$

Finding the Value of K:

Once we have found the rate law for a reaction, we can find the value of the rate constant ***k***. Simply **SUBSTITUTE** the data from **ANY TRIAL** into the rate law and solve for ***k***.

Example:

For the reaction: $3A_{(g)} + B_{(g)} + 2C_{(g)} \rightarrow 2D_{(g)} + 3E_{(g)}$ the following data was obtained:

Trial	[A](mol/L)	[B] (mol/L)	[C] (mol/L)	Rate (mol/Ls)
1	0.10	0.10	0.10	0.20
2	0.20	0.10	0.10	0.40
3	0.20	0.20	0.10	1.60
4	0.20	0.10	0.20	0.40
5	0.50	0.40	0.25	??
6	??	0.60	0.50	6.00

×2 ×2 ×2 ×4 ×2 ×1

$2^1 = 2$

$2^2 = 4$

$2^1 = 0$

a) Determine the Rate Law:

$Rate = K[A]^1[B]^2[C]^0$

$Rate = K[A]^1[B]^2$

Finding the Value of K:

$$\left(\frac{mol}{L}\right)\left(\frac{mol}{L}\right)^2 = \frac{mol}{L} \times \frac{mol}{L} \times \frac{mol}{L}$$

Example (con't):

Trial	[A](mol/L)	[B] (mol/L)	[C] (mol/L)	Rate (mol/Ls)
1	0.10	0.10	0.10	0.20
2	0.20	0.10	0.10	0.40
3	0.20	0.20	0.10	1.60
4	0.20	0.10	0.20	0.40
5	0.50	0.40	0.25	??
6	??	0.60	0.50	6.00

b) Calculate the value of the rate constant.

$$RATE = K[A]^1[B]^2$$

$$\frac{0.2\frac{mol}{L\cdot S}}{(0.1\frac{mol}{L})(0.1\frac{mol}{L})^2} = \frac{K\cancel{(0.1\frac{mol}{L})}\cancel{(0.1\frac{mol}{L})^2}}{\cancel{(0.1\frac{mol}{L})}\cancel{(0.1\frac{mol}{L})^2}}$$

$$K = \frac{0.2\frac{mol}{L\cdot S}}{\left[(0.1\frac{mol}{L})(0.1\frac{mol}{L})^2\right]}$$

$$K = 200$$

$$\frac{\left(\frac{mol}{L\cdot S}\right)}{\left(\frac{mol}{L}\right)\left(\frac{mol}{L}\right)^2} = \frac{\frac{mol}{L\cdot S}}{\frac{mol^3}{L^3}} = \frac{\cancel{mol}}{\cancel{L}\cdot S} \times \frac{L^{\cancel{3}\,2}}{mol^{\cancel{3}\,2}} = \frac{L^2}{mol^2\cdot S}$$

c) Calculate the rate for trial #5.

$$RATE = 200[A][B]^2$$
$$= 200(0.5)(0.4)^2$$
$$= 16\ \frac{mol}{L\cdot S}$$

d) Calculate the concentration of A in trial #6.

$$RATE = 200[A][B]^2$$
$$6 = 200[A](0.6)^2$$
$$[A] = \frac{6}{(200)(0.6)^2}$$
$$[A] = 0.083\ \frac{mol}{L}$$