# 1 | What are kinetics:

- how fast a reaction proceeds

# 2 | Why care?

- Real world application: if you want to make a chemical (say some sort of medicine), then you will want to make it faster so you can keep with the demand of the product.
- Kinetics is the study of the speed of reactions, so you can apply it to speeding the production of your medicine

# 3 | Basis of kinetics:

- collision theory

# 4 | Collision theory:

- particles must collide before a reaction can take place
- not all collisions lead to a reaction
- reactants must posses at least a minimun amount of energy - **Activation energy**
- **Steric effect** : particles must approach each other in a certain way

# 5 | Qualitative side to kinetics:

## 5.1 | according to collision theory, for a faster recation, you want:

### 5.1.1 | more frequent collisions

- more speed
- more particles

### 5.1.2 | more successful collisions

- more energy
- lower activation energy

## 5.2 | what this looks like

### 5.2.1 | increase the surface area

- this way more particles are exposed
- powdered solids react faster than larger lumps

### 5.2.2 | increase the temperature

- more kinetic energy to your particles
- this increases the speed of your particles
- this increases frequency of collisions
- the extrac kinetic energy also allows the particles to get over the activation energy thershold
  - Boltzman distribution

### 5.2.3 | shine light

- same as temperature, just a different source of energy

### 5.2.4 | add a catalyst

### 5.2.5 | increase the pressure of gases

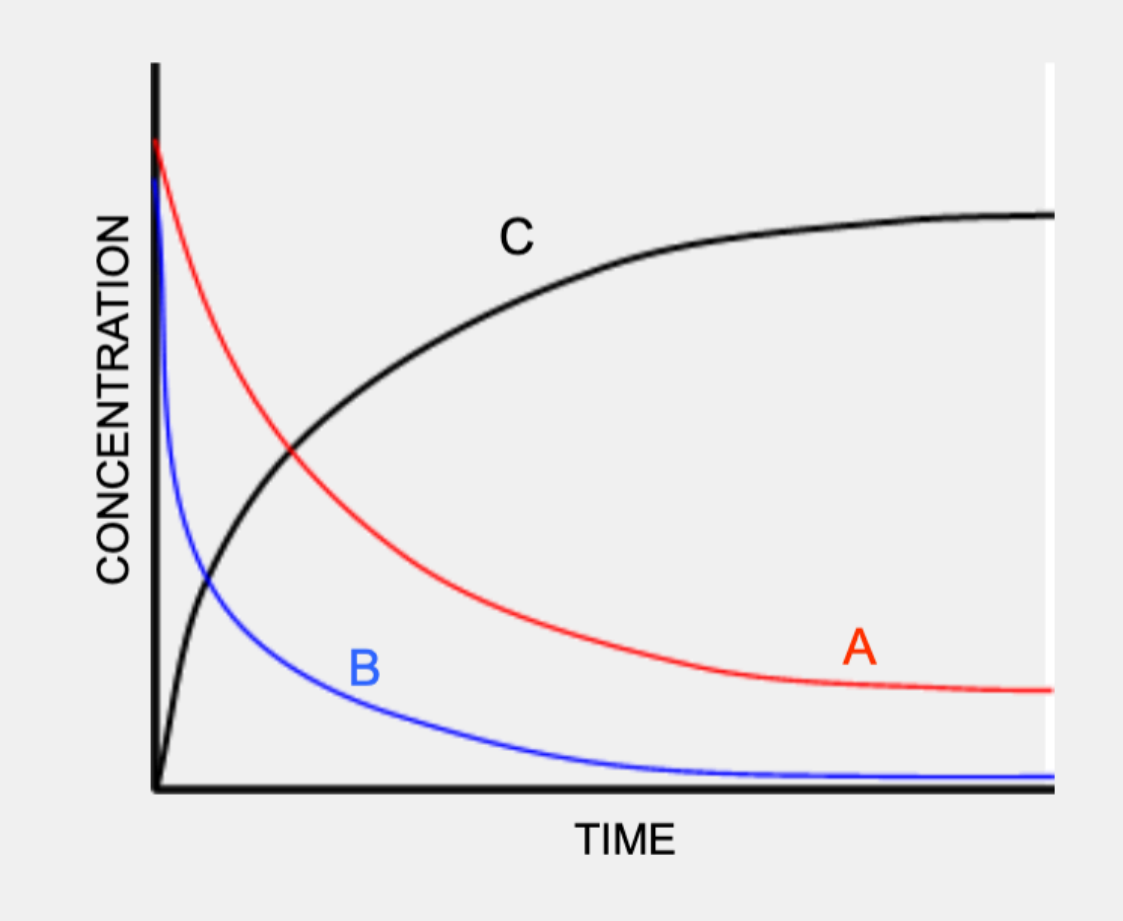
- when you increase the pressure of the vessel you are essentially decreasing the volume of the vessel
- this means your molecules are going to interact with each other more often
- this will increase the frequency of collisions

### 5.2.6 | increase the concentration of reactants

- you can also increase the amount of the reactants, this leaves a higher chance of collision which will lead to more collisions and thus a faster reaction

# 6 | Quantitative side to kinetics:

## 6.1 | **rate of a reaction:**




- the reaction starts off fast and then asymptotically gets slower
- this is because as time goes on there will be less and less reactant particals to collide with and keep the reaction going.
- the rate of the reaction would be the derivative of the C curve
- the **initial rate** of a reaction is the slope af the very begining of the curve

## 6.2 | **The initial reaction rate equation:**

- $r = k[A]^x[B]^y$, where:
  - k is a constant, depends on the reaction
  - [A] is the concentration of A
  - [B] is the concentration of B
  - x is an exponent that change based on the reaction
  - y is an exponent that changes based on the reaction
- unit of the rate of an equation: mol/dm$^3$/s

### 6.3 | **Order of a reaction:**

- $x + y$, where x and y are what was from the rate equation

## 7 | **Given the graph of concentration (of one chemical) vs time:**

- you can look at the shape to get the order of the concentration
- if you have a line (linear) that points downward then you have an order of 0
- if you have a curve, that curves downward you have an order of 1
- if you have a half of a parabola that points upward, then you have an order of 2

## 8 | **Arrhenius Equation:**

- $k = Ae^{-\frac{E_a}{RT}}$
  - Where k is the rate constant
  - A is pre-expoential factor (frequency collision factor)
  - e is the constant
  - $E_a$ is activation energy
  - R is a constant, 8.31 J/(K*mol)
  - T is temperature (in kelvin)