

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 possess at least a minimum 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 reaction, 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

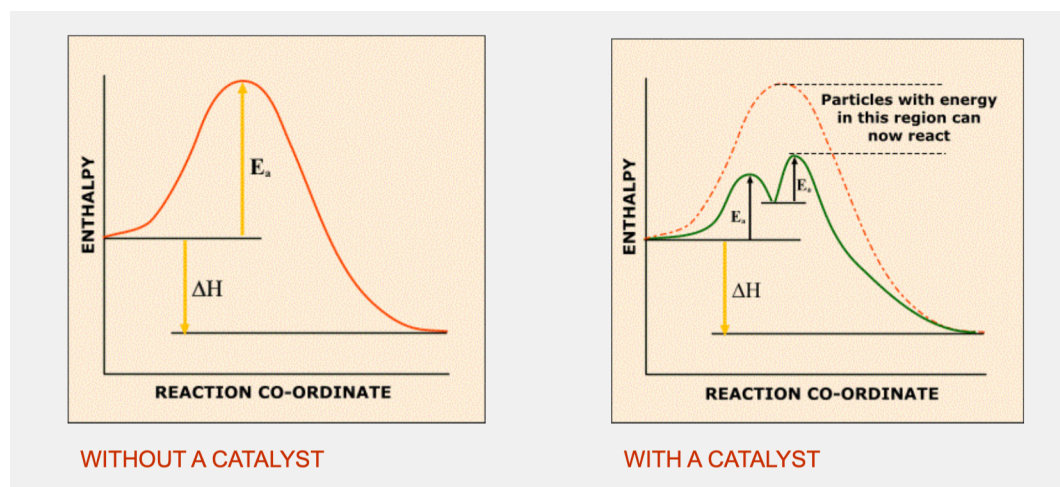
- more kinetic energy to your particles
- this increases the speed of your particles
- this increases frequency of collisions
- the extra kinetic energy also allows the particles to get over the activation energy threshold
 - Boltzmann distribution

5.2.3 | shine light

- same as temperature, just a different source of energy

5.2.4 | add a catalyst

- when added, catalysts provide an alternative reaction path with a lower activation energy (change the mechanism of the reaction)
- decreasing the activation energy means that more particles will have sufficient energy to overcome the energy barrier and react
- because they reduce the activation energy in the forward direction, they also reduce the activation energy in the reverse direction, see diagram below
- the catalyst remains chemically unchanged at the end of the reaction.



5.2.5 | increase the pressure of gases

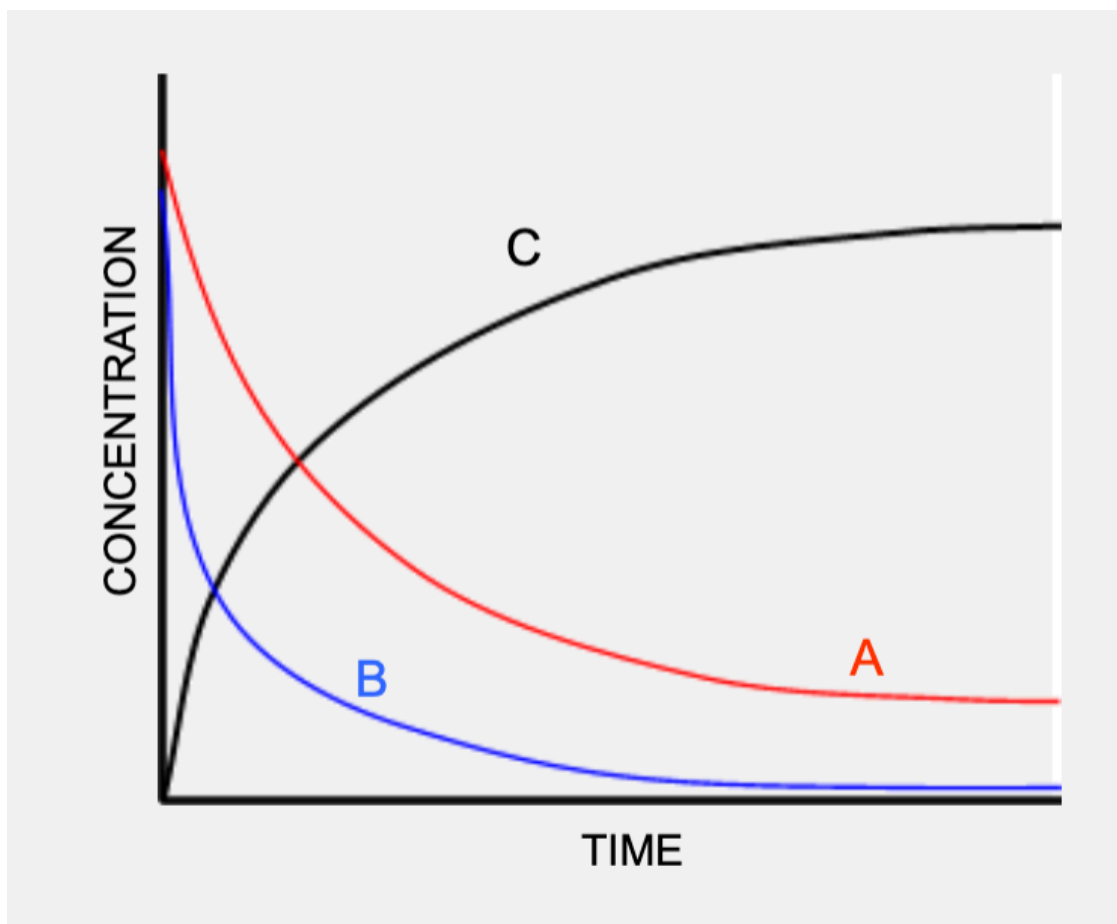
- when you increase the pressure of the vessel you are essentially decreasing the volume of the vessel
- this means your molecules are going to interact with each other more often
- this will increase the frequency of collisions

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 particles 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 of the very beginning 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 changes based on the reaction
 - y is an exponent that changes based on the reaction
- unit of the rate of an equation: $\text{mol/dm}^3/\text{s}$

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-exponential factor (frequency collision factor)
 - e is the constant
 - E_a is activation energy
 - R is a constant, $8.31 \text{ J/(K}\cdot\text{mol)}$
 - T is temperature (in kelvin)

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