**Kinetics: The VERY Short Story**

*Covered in Chapter 14 of Brown et al.*

**Kinetics**: The speed or rate at which chemical reactions occur (M/s)

* Factors affecting kinetics: physical state; concentration; temperature; catalyst

**Reaction rates:** expressed as ∆[X] / ∆t (M/s)

* Reactants disappear, so have negative (-) reaction rates
* Products appear, so have positive (+) reaction rates
* The rate of a reaction can change with time (zB slow as reactants are used up)
* Average rates are calculated over defined periods of time
* Instantaneous rates are calculate using a line tangent to a single point in time
* Stoichiometry affects reaction rate:

For this equation: aA + bB 🡪 cC + dD

**Rate =** - 1 ∆[A] **=** - 1 ∆[B] **=** + 1 ∆[C] = + 1 ∆[D]

 a ∆t b ∆t c ∆t d ∆t

**Rate laws:** equations showing that reaction rate is proportional to concentrations of

 reactants

For the reaction: aA + bB 🡪 cC + dD

**Rate = k [A]m[B]n**

* k is the rate constant & can have a variety of units
* m & n are the orders of reaction for the reactants
	+ Overall order of reaction is the sum of all individual orders for reactants
* Solve for m & n by:
	+ Finding a set of experiments in which the concentration of one reactant, but not the other, changes;
	+ Compare the two experiments and calculate the change in concentration of that reactant and the change in rate;
	+ Solve for orders by using this equation: ∆[X]m = ∆rate; and
* Solve for k by plugging in data from one experiment into the rate law.

**Graphical determination of orders of reaction:** The order of reaction for a reactant can be determined by plotting the concentration of that reactant vs. time.

* First order if a linear curve is produced by plotting the ln [reactant] vs. time
* Second order if a linear curve is produced by plotting 1/[reactant] vs. time

**Half-life:** is the time required for the concentration of a reactant to decrease to half of its

 original value. Different formulas are used to calculate the half-lives of first- and
 second-order reactions.

* First order t1/2 = 0.693/k
* Second order t1/2 = 1/k[A]0

**Temperature:** Generally, reaction rate doubles for every 10°C increase in temperature.

There are three ways that temperature increases rate:

* Collision frequency is increased because molecules move more quickly at higher temperatures;
* Orientation must be correct or “on-target” for reactions to go forward; temperature increase the rate at which orientation changes;
* Activation energy – higher temperatures allow reactions to get up and over activation energies.

**Transition states:** are the highest energy states occurring during a chemical reaction; they are states between reaction and product and are very unstable and fleeting.

**Reaction mechanisms** describe how chemical reactions occur; steps of reaction.

* Elementary reactions
* Multistep reactions
	+ Multistep reactions can involve reaction intermediates, chemical species that are formed and used in the reaction, but are neither reactants nor products.

**Rate-limiting steps** are the slowest step of multistep chemical reactions. Rate-limiting steps govern, or determine, the overall speed of a chemical reaction.

**Catalysts** are substances that increase the speed of chemical reactions but are not changed by them.

* Homogenous catalysts have the same physical state, or phase, as the reactants.
* Heterogenous catalysts have a different physical state than the reactants.
* All catalysts speed reactions by decreasing activation energy.
* Enzymes are the catalysts of biological reactions and are usually protein molecules.