

Comprehensive and Easy Course Notes for CHEM1200

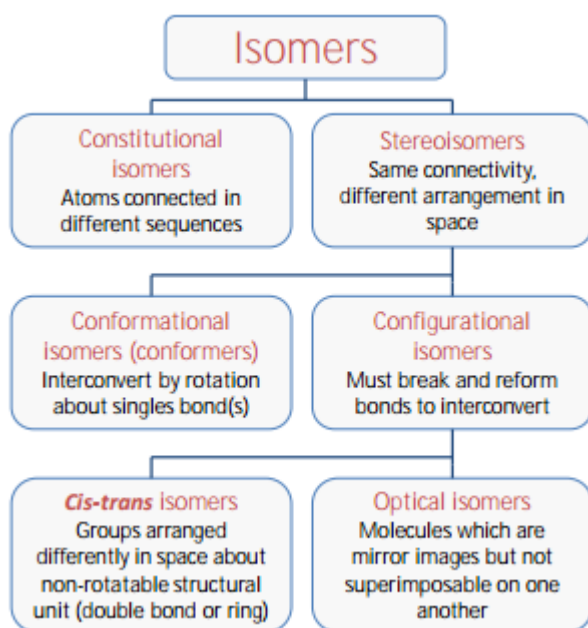
Exams and Assessment

MODULE 1: STEREOCHEMISTRY

What is stereochemistry?

- Concerned with the actual shapes of molecules in 3D space
- The spatial arrangement of atoms within molecules can influence their reactivity and properties
- Stereochemistry gives rise to new types of stereoisomers which concern the spatial arrangement of the atoms

The types of isomers



Constitutional Isomers

- The atoms are connected in different sequences to produce completely different compounds

Stereoisomers

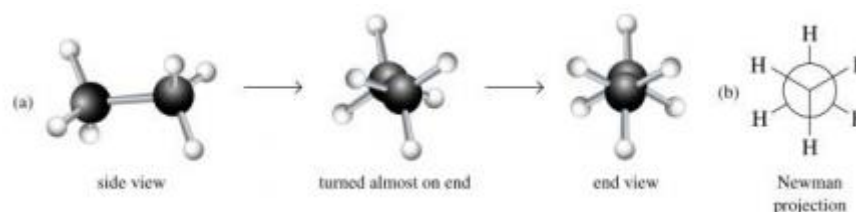
- Have the same connectivity of atoms but they are arranged differently in space – the type of these isomers depends on whether they must break bonds or not to interconvert between each form

Conformational Isomers (conformers)

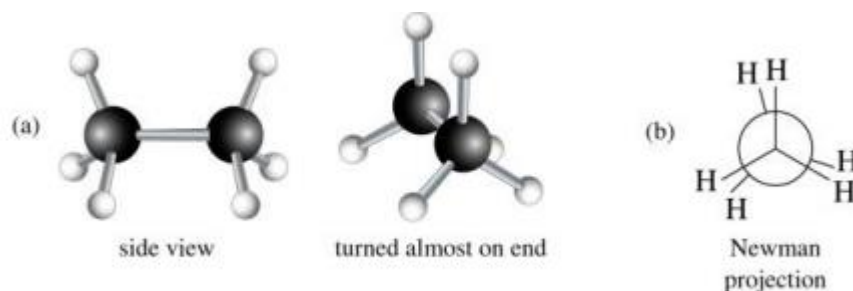
- Interconvert by rotation about a single bond as single bonds are rotatable due to the end on overlap of sp^3 orbitals.
 - Rotate to relieve high energy (unstable) conformers preferring to adopt low energy conformers.
- Shape can determine direction of attack by reagents
- Interaction of organic compounds with enzymes governed by shapes

- There are two types of conformation:

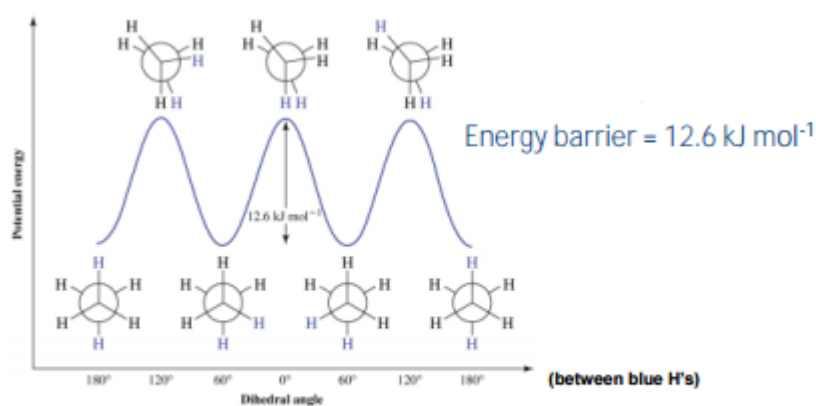
- Staggered: which is the most stable as the C-H bonds experience maximum separation and therefore minimise repulsive forces between these bonds on the neighbouring carbons



- Eclipsed: highest energy, but least stable as the constituents bond to neighbouring carbons are overlapping in position therefore repulsive forces are the highest



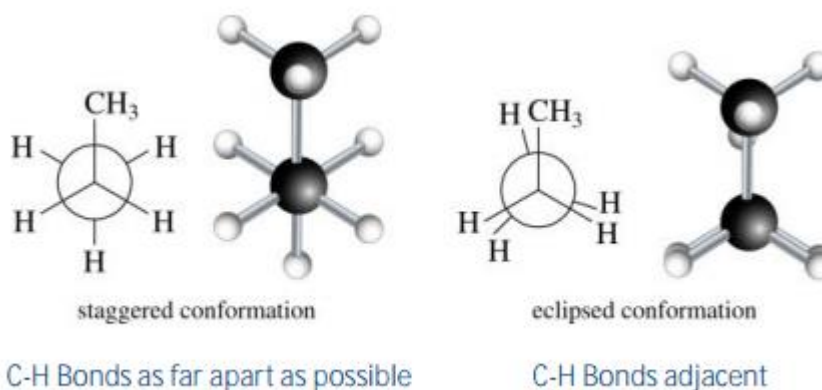
Staggered is at a much lower energy than eclipsed because of the minimal repulsive forces and thus most alkanes are found in this form.



- At room temp, rotation is **easy and rapid** →
- Mixture of all possible conformations, but > 99% staggered

Conformers get slightly more complex as the number of carbons on the alkane increase as you have more carbon pairs to consider. For example on a propane you can consider the conformation of carbon one and carbon two or carbon two and carbon three.

Note in the example below that one of the bonds on the front carbon is a methyl group (which in this case is the third carbon in the carbon chain).

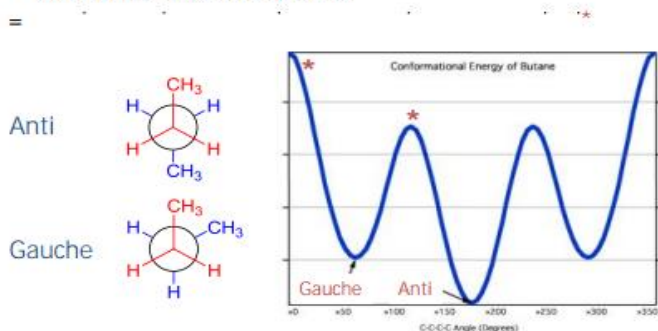


Then when you get to a carbon chain of four, butane, there are two possibilities for both the staggered and eclipsed forms of this alkane due to the two methyl groups.

1. For eclipsed
 - a. You can either have the form where the two methyl groups on the neighbouring carbons overlap and this is the highest energy form due to the high level of repulsion between the two bulky groups.
 - b. The other form is where the perceived angle between the neighbouring methyl groups is 120 degrees and neighbouring methyls line up with neighbouring hydrogens, this is still high in energy by slightly lower than the first.
2. For staggered
 - a. The two methyl groups could be close together (with a perceived angle of 60 degrees between them). THIS IS KNOWN AS THE GAUCHE FORM.

- b. The two methyl groups are as far away as possible (180 degrees). THIS IS KNOWN AS THE ANTI FORM. This is lower in energy than the gauche form because of minimised repulsive forces.

- At room temperature all possible conformers present
 - about 80% anti; 20% gauche

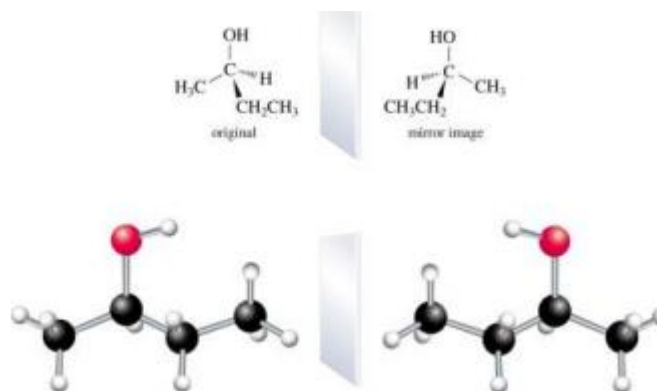


Configurational Isomers – must break and reform bonds to interconvert

There are two types:

Optical isomers – molecules which are mirror images but not superimposable on one another.

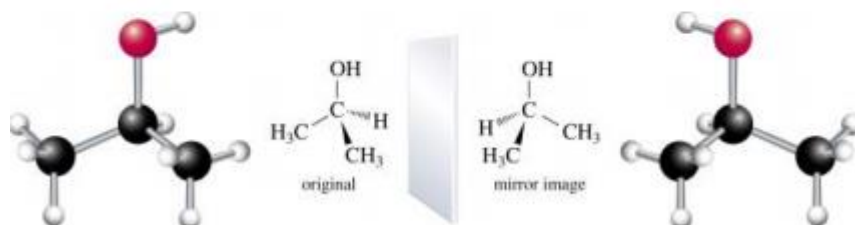
- This involves the concept of chirality (the object is not superimposable on its mirror image). Each one of the pair of non-superimposable molecules is called an **enantiomer**.
- In the simplest case we can consider a molecule which is bonded to four different groups



- A molecule that contains a carbon centre with two identical groups is not chiral. It is superimposable on its mirror image and therefore is the same molecule.

e.g. propan-2-ol

- Not chiral
- has internal plane of symmetry
- is superimposable on its mirror image
- does not have enantiomers!

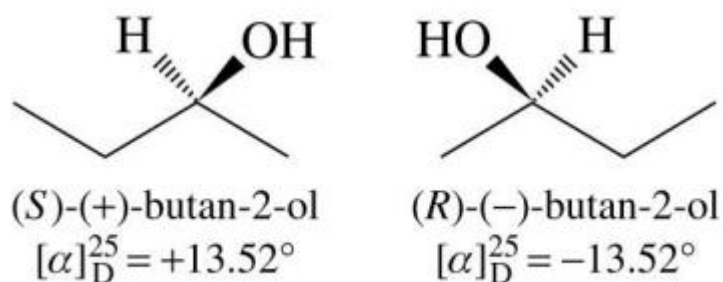


How do you differentiate between enantiomers?

Enantiomers have identical physical properties except for the direction in which they rotate the plane of plane-polarised light.

- In plane polarised light oscillation occurs in one direction only (a polariser is used to filter out unwanted oscillations)
- For optical isomers, polarised light rotates in opposite directions at the same degree
 - Clockwise rotation DEXTROTATORY (+) (d)
 - Anticlockwise rotation LAEVOROTATORY (-) (l)

Each molecule of a pair of enantiomers rotates plane-polarised light by the same amount, but in opposite directions.



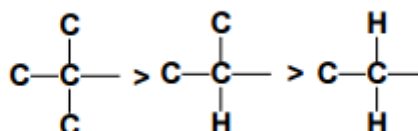
- What is a racemic mixture?
 - A 50:50 mixture of enantiomers is known as a racemic mixture
 - Zero optical rotation is observed because the rotations cancel each other out.
 - We use the prefix (+-) to indicate a racemic mixture

Why are enantiomers important?

- Can have very different biological properties
- Interact differently with other chiral molecules
 - o Enzymes, drug receptors, taste and smell receptors, are all chiral proteins
 - o Hence, different enantiomers are expected to behave differently in biological systems

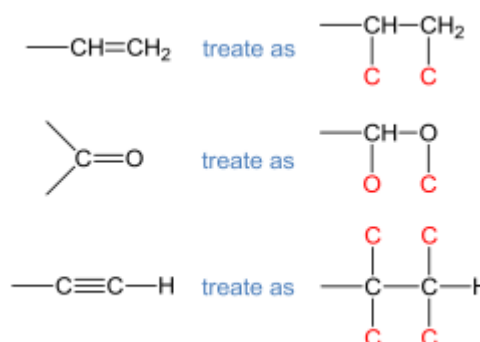
Assigning absolute configuration of stereogenic centres

- A set of arbitrary rule sot specify the spatial arrangement of the four groups attached to a chiral carbon atom
1. Assign priority to each of the 4 groups (based on the atomic number of the attached atoms where first priority is the highest atomic number)
 2. View the molecule from side opposite lowest priority group (so put the lowest priority group (usually a hydrogen) as far away from the viewer as possible)
 3. If priority sequence is:
 - a. **Clockwise: R configuration**
 - b. **Anticlockwise: S configuration**
 4. If two atoms attached to the stereogenic carbon are the same
 - a. Compare the next closest sets of atoms until a priority can be assigned (the more Hs attached to the bonded carbon (generally) the lower the priority group)



If this is the carbon attached to the central carbon.

5. Multiple bonds should be treated as single bonds with added dummy atoms

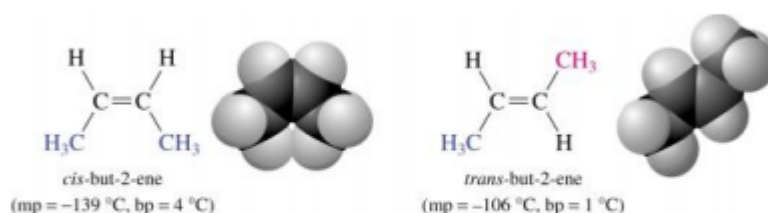


Cis-trans isomers

- Groups arranged differently in space about non-rotatable structural unit (double bond or ring)
- In a pi bond, sp² hybridised carbons form a sigma bond and then their remaining p orbital shares electron density about the top and bottom of the sigma bond. This means that the carbons sharing this double bond cannot twist – this allows for two configurations of the same molecule.
- **FOR DISUBSTITUTED DOUBLE BONDS (WITH TWO METHYL GROUPS ON EITHER SIDE OF DB)**

Cis-trans isomerism

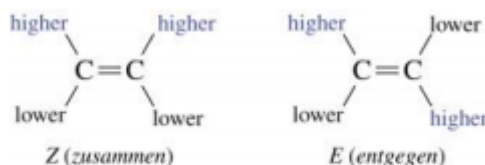
So, the two forms of but-2-ene cannot interconvert without breaking the π -bond:



- **CIS** - methyls on **same side** of double bond (close together)
- **TRANS** - **opposite sides** (far apart)
- these are called “configurations”

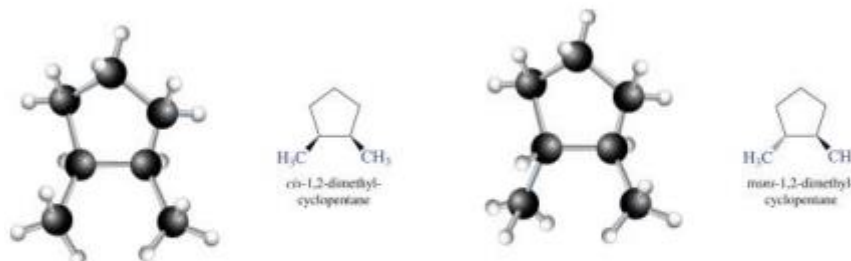
Which is more stable - *cis* or *trans*?

- Cis is higher in energy than trans making it less stable due to steric interactions between larger substituents (as they are on the same side of the double bond and therefore interact more)
 - STERIC INTERACTION – spatial interactions
- NOTE: for cis-trans isomerism in alkenes, each end of the double bond must have two different groups attached (if one end have two of the same groups, evening flipping the molecule gives the same molecule).
- **FOR TRI OR TETRASUBSTITUED DOUBLE BONDS:**
- Assign priority to groups on either end of the double bond (separately)
- If the two highest priority groups are on
 - The same side: it is ‘Z’ isomer
 - Opposite sides: it is ‘E’ isomer



- same side of ring – cis
- Opposite side of ring - trans

Cis-trans isomerism in rings



- Cannot interconvert structures without breaking & reforming bonds:
 - Ring prevents rotation of the C-C bonds

REACTIONS AND REACTIVITY

All reactions involve bond breaking and bond making.

What are the options for breaking bonds? Where do the electrons go?

- Heterolytic fission (heterolysis)
 - for a covalent bond between atoms of different electronegativity
 - Bonding electrons are unequally divided producing cations and anions through bond breaking.
 - This occurs as the electrons in the bond are more strongly associated with the more electro negative element meaning both will leave with this more electronegative element which has a partial negative charge in the bond.
- Homolytic fission
 - For a covalent bond between atoms of similar electronegativity
 - Bonding electrons are equally divided as the constituents have the same electronegativity.
 - This produces free radicals which contain an unpaired electrons and are very reactive.

What is a reactive intermediate?

A carbon-based cation (carbocation) has three bonds around the carbon centre meaning it has one empty orbital.

A carbon based free radical also has less than four bonds meaning it has an sp^3 orbital containing a single unpaired electron.

Both of these are quite unstable and energy therefore must be supplied to allow them to form and they are easier to form if some feature in the molecule stabilises them (they have a large group attached).

What is an addition reaction?

These are often reactive intermediates which are unstable structures that form fleetingly during the course of the reaction. It is important to note however that these are still physical intermediate products.

How do we represent electron movement?

Curly (mechanistic) arrows

Provides a way of showing electron movement when writing equations

Used to show electron movement:
"Electron pushing"



Single-headed arrow

- Indicates movement of **one electron**

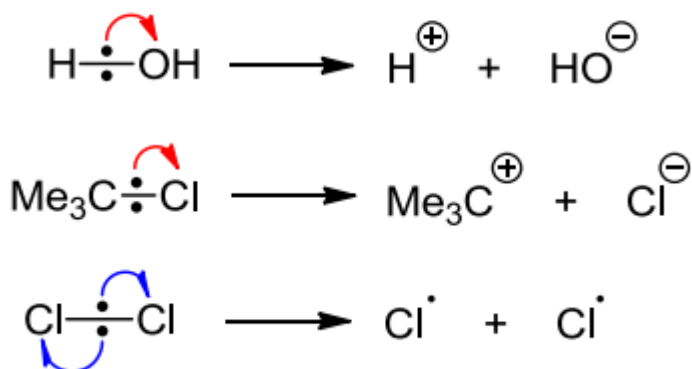


Double-headed arrow

- Indicates movement of a **pair of electrons**

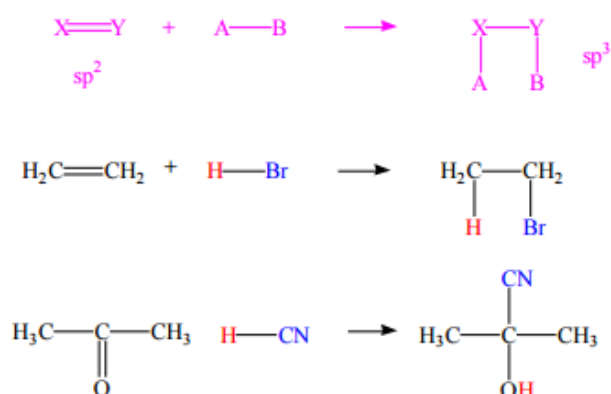
Curly (mechanistic) arrows

Provides a way of showing electron movement when writing equations



What is an addition reaction?

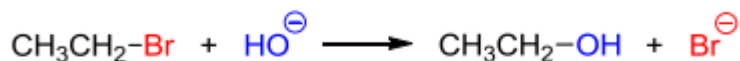
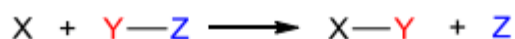
Involves the breaking of a pi bond and the addition of two groups across the double/triple bond. The Carbons transition from sp² hybridised to sp³ (for a double to a single bond).



What is a substitution reaction?

Involves a

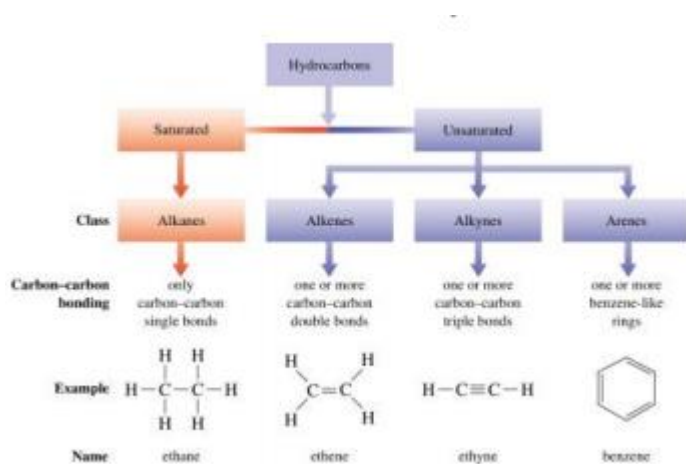
- Nucleophile – has more electrons than it needs and is looking to share electrons
OR
- Electrophile – has less electrons than it needs and is looking to gain electrons



- HO^- is a **nucleophile** → nucleophilic substitution
- NO_2^+ is an **electrophile** → electrophilic aromatic substitution
 - Note: **conservation of charges**

Alkene and Alkyne Chemistry

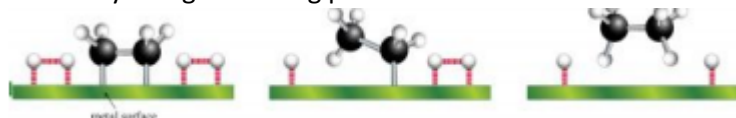
The main reaction of alkenes and alkynes is addition to give saturated products (derivatives of alkanes).



What the reactions of alkenes?

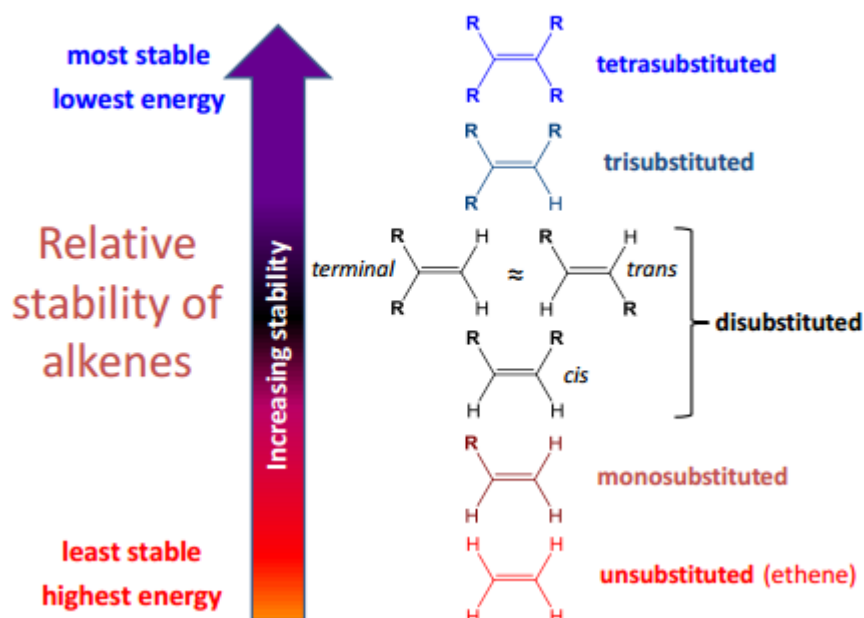
1. Hydrogenation

- Reaction with molecular hydrogen (H_2)
- Without a catalyst no reaction occurs, needs a pd, pt or ni metal catalyst to facilitate the reaction by being a meeting platform for the two reactants



- Due to the nature of the catalyst, the addition of the two Hs will always occur on the same side. Therefore any addition methyl groups end up cis on the double bond for rings.

What are the relative stability of alkenes?



The presence of hydrogen groups do not stabilise the nucleus. The pi bonded p orbitals are not supported by the Hs. Therefore the more large things they are attached to, the more stable the molecule.

What is an example of hydrogenation in the real world?

Fatty acids are the simplest type of lipid which consist of a long carbon chain with a carboxylic acid at one end. Unsaturation leads to a lower melting point due to the kinks in the straight chain and thus less efficient molecular stacking.

Unsaturated oils at room temperature can be converted to spreadable fats by a process known as hardening by which they catalyse hydrogenation of some or all of the oil's alkene groups.

2. Electrophilic addition to alkenes (hydrohalogenation)

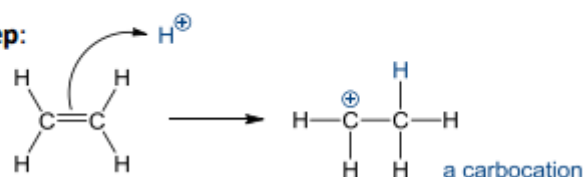
- An electrophile will react with the pi electrons (two electrons in the pi bond).
- An electrophile (here H^+) adds to one end of the double bond and a nucleophile (here Cl^-) adds to the other end

Step 1: A pair of electrons is supplied to H^+ . This forms a carbocation intermediate.

Electrophilic addition to alkenes: stepwise mechanism

e.g. Addition of HCl to ethene (simplest alkene)

1. First reaction step:



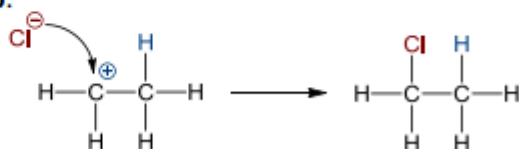
Step 2: To form a bond with the positively charged C (electrophile), the Cl^- (nucleophile) provides a pair of electrons for bond formation.

Electrophilic addition to alkenes: stepwise mechanism

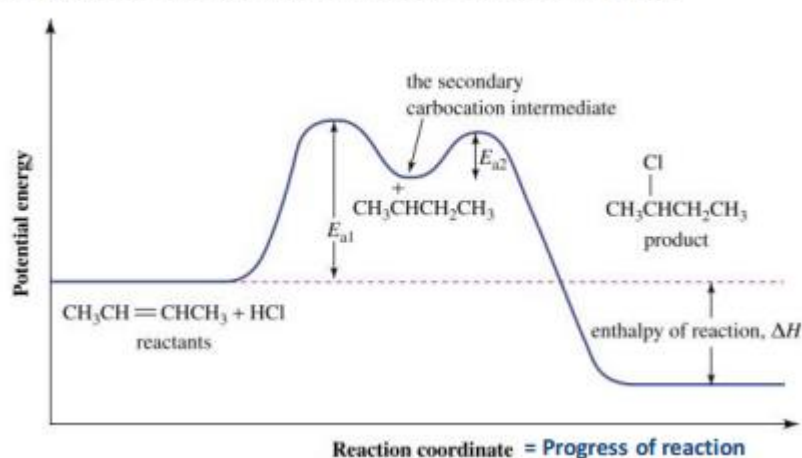
e.g. Addition of HCl to ethene (simplest alkene)

- A typical mechanism involving ionic intermediates:

2. Second reaction step:



Energy level diagram for the addition of HCl to but-2-ene

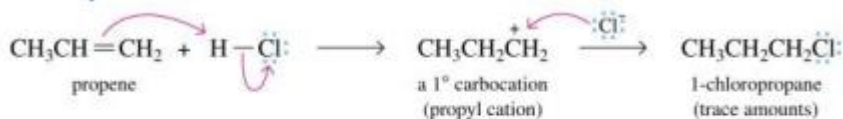


What is regioselectivity and markovnikov's rule?

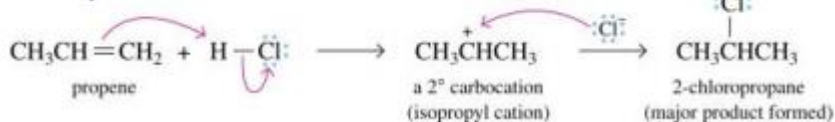
When a reaction can yield two or more regioisomers but produces mainly one it is regioselective.

For more complex alkenes there are two possible products depending on which end of the double bond is protonated (has the H attached).

Pathway A



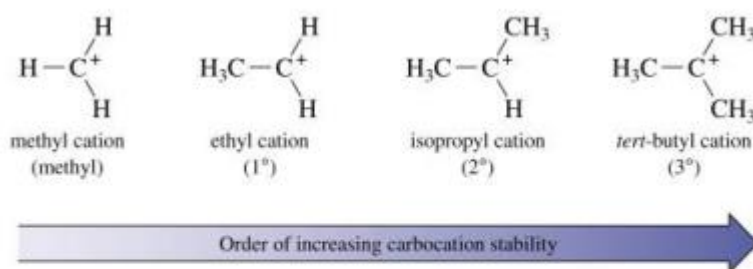
Pathway B



We use markovnikov's rule to determine which carbon the hydrogen bonds to.

Rule: when a hydrogen halide (HX) is added to an unsymmetrical alkene, the hydrogen of the reagent always adds to the less substituted alkene carbon (the one with most hydrogen atoms attached to it).

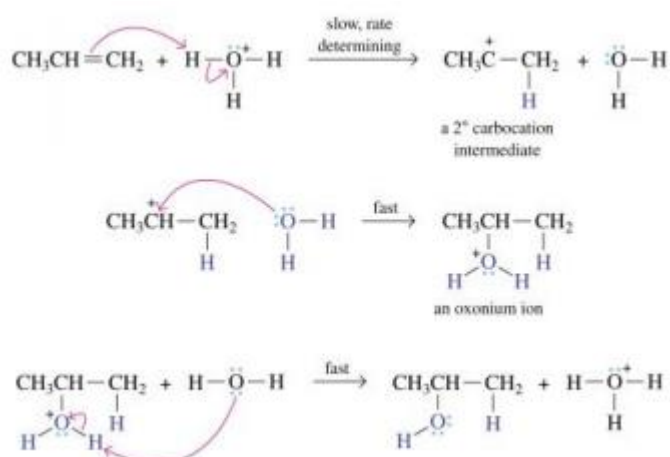
This produces the most stable carbocation intermediate (as it will have the most about of stabilising carbons attached to it).



- Alkyl groups bonded to the positively charged carbon atom donate electron density towards it
 - This helps stabilise the positive charge and is known as **the inductive effect**
- The more alkyl groups attached, the bigger the effect

3. Hydration of alkenes

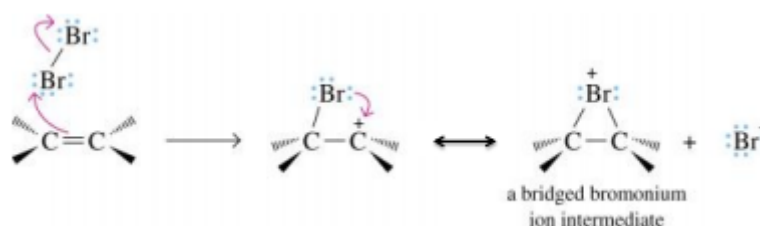
- Acts similar to the addition of halides, is acid-catalysed, highly regioselective



The acid facilitates the reaction but then you get the acid at the end therefore it is a catalyst.

4. Addition of Cl_2 or Br_2 to alkenes

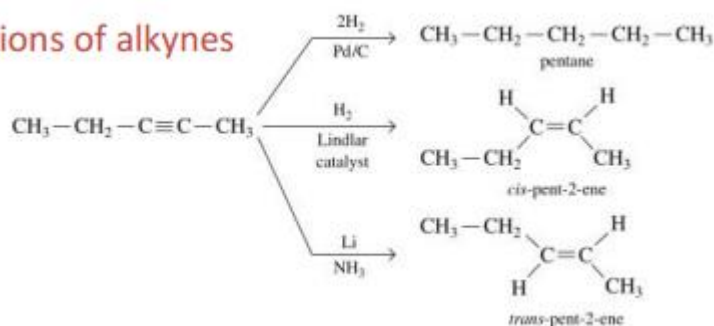
- Bromine and chlorine are non-polar but they still react and act as nucleophiles and electrophiles
- They do this due to the way they act in the presence of other electron density
 - They are polarised as electron density is repelled by close approach to electron rich alkene
 - One end is electrophilic and will react with the alkenes



- Electrophilic Br receives two electrons from carbon resulting in heterolytic cleavage of the Br-Br bond.
- Three membered ring in which the bromine atom bears the positive charge is the bridge bromonium ion intermediate

What are the reactions of alkynes?

Reactions of alkynes



1. Alkynes can be hydrogenated to alkanes
2. Hydrogenation with a **deactivated catalyst (lindlar catalyst)** stops after one mole of H_2 is added giving the **cis-alkene**
3. Reduction of alkynes with **lithium metal** gives the **trans-alkene**

CHEMICAL KINETICS

RATES OF REACTION

When is a reaction spontaneous?

$$\Delta G = \Delta H - T\Delta S$$

If $\Delta G < 0$ process is spontaneous

If $\Delta G < 0$ reverse process is spontaneous

If $\Delta G = 0$ system is at equilibrium

What is the rate of reaction?

Determined by measuring the change in concentration or product with time measured in Ms^{-1} . It is the gradient of a concentration time graph.

Rate = - change in concentration of reactant / time = change in concentration of product / time

What are factors that affect the rate of reaction?

1. Chemical nature of the reactants

Chemical reactivity is similar for similar groups. The lower the activation energy for a

2. Ability of the reactants to come in contact (collide) with each other

Collision theory

- Reactants must collide at specific orientation and have sufficient kinetic energy at collision. The more collisions the more likely to get a successful orientation.

- The larger the surface area of contact between reactants, the faster the reaction rate.

3. Temperature

The higher the temperature the faster the reaction rate. See Arrhenius equation.

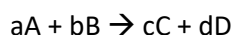
4. Concentrations of the reactants

The greater the concentration, the greater the probability for the reactants to collide successfully and therefore a faster reaction.

5. Rate-accelerating agents (catalysts)

RATES LAWS

What are relative rates of change?



The relative rate is given by:

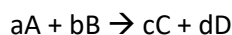
$$rate = -\frac{1}{a} \frac{d[A]}{dt} = -\frac{1}{b} \frac{d[B]}{dt} = \frac{1}{c} \frac{d[C]}{dt} = \frac{1}{d} \frac{d[D]}{dt}$$

This is because relatively the rates of each will be effected by the coefficients.

You can base the rates on any given reaction by making the coefficient of that reactant 1 and changing the other coefficients relative to this reactant.

What is the rate law?

To predict how the rate of a reaction will change as a function of time, we need to experimentally determine the mathematical relationship between the rate of the reaction and the concentrations of reactants. Rate laws are normally simply functions of the concentration of the reactants.



the rate law is:

$$rate = k[A]^n[B]^m$$

- n is the order in terms of reactant [A] and m is the order in terms of reactant [B].
- the sum n+m is the overall order of the reaction
- the order of a reaction is related to the mechanism of and not determined by the stoichiometry
- k is the rate constant

k, n and m are found experimentally

How do you determine reaction orders?

This is determined by how the rate depends on the concentration of a reactant.

The method of initial rates:

The initial rate of reaction is the rate just after the reaction has started (just after $t=0$).

If a reaction is:

- first order in a particular reactant, then doubling its concentration will double the initial rate
- second order in a particular reactant, then doubling its concentration will increase the rate by factor of 4
- zero order in a particular reactant, the rate will be independent of the reactant concentrations.

What are the units of the rate constant, k ?

If the overall order is;

0 – k has units $\text{mol L}^{-1} \text{s}^{-1}$

1 – k has units s^{-1}

2 – k has units $\text{mol}^{-1} \text{Ls}^{-1}$

What are the integrated rate laws?

A rate law that expresses how a rate depends on concentration is called a differential rate law or just the rate law.

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

A rate law that expresses how the concentration depends on time is called an integrated rate law.

Consider: $A \rightarrow \text{products}$

For a reaction that is first order in reactant A:

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

Move all the concentrations together

$$\frac{d[A]}{[A]} = -kdt$$

Integrate both sides

If we integrate this expression from $t = 0$ (when $[A] = [A]_0$) to time t (when conc. of A is $[A]_t$)

$$\int_{[A]=[A]_0}^{[A]=[A]_t} \frac{1}{[A]} d[A] = -\int_{t=0}^{t=t} k dt$$

$$\ln[A]_t - \ln[A]_0 = -(kt - kt_0)$$

since $t_0=0$:

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

$$\ln[A]_t = -kt + \ln[A]_0$$

Maths hint

$$\int 1/x dx = \ln x$$

$$\int k dx = kx$$

- We now have an expression for how $[A]$ varies with time.

FIRST ORDER INTEGRATED RATE LAW

$$\ln[A]_t = -kt + \ln[A]_0$$

This leads to the exponential decay, the larger the rate constant the more rapid the decay.

For a reaction that is second order for reactant A:

$$rate = -\frac{d[A]}{dt} = k[A]^2$$

$$\frac{d[A]}{[A]^2} = -kt$$

Integrate both sides:

SECOND ORDER INTEGRATED RATE LAW

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

For a reaction that is zero order (rate is independent of concentrations):

$$rate = -d[A] = kdt$$

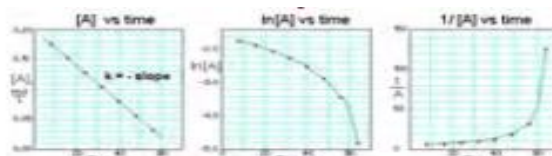
Integrate both sides:

ZERO ORDER INTEGRATED RATE LAW

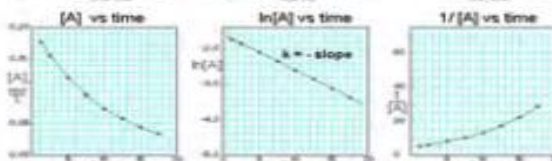
$$[A]_t = -kt + [A]_0$$

Order of Reaction			
	First	Second	Zero
Rate law	Rate = $k[A]$	Rate = $k[A]^2$	Rate = k
Integrated rate law	$\ln[A] = -kt + \ln[A]_0$	$1/[A] = kt + 1/[A]_0$	$[A] = -kt + [A]_0$
Plot needed to give a straight line	$\ln[A]$ versus t	$1/[A]$ versus t	$[A]$ versus t
Relationship of rate constant to the slope of straight line	Slope = $-k$	Slope = k	Slope = $-k$

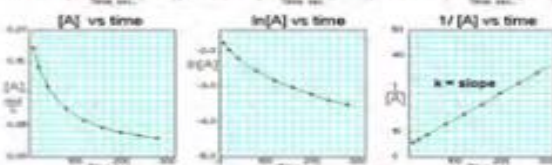
Zero order reaction



1st order reaction



2nd order reaction



How do you find the half-life for each of the different orders of reaction?

The half life is the time it takes for the concentration of a reactant, A, to drop to half of its original value,

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

FIRST ORDER REACTION:

$$\ln[A]_t = -kt + \ln[A]_0$$

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

$$t_{\frac{1}{2}} = \frac{\ln 2}{k}$$

For a constant rate – the half-life is constant.

SECOND ORDER REACTION:

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

The half-life increases as the concentration decreases.

THIRD ORDER REACTION:

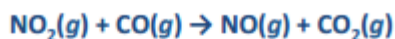
$$t_{\frac{1}{2}} = \frac{[A]_0}{2k}$$

The half-life decreases with an increase in concentration.

What is the mechanism of a reaction?

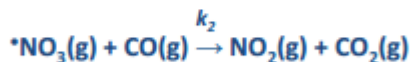
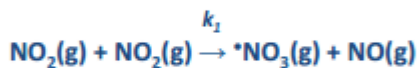
A balanced chemical equation does not tell you how the reaction occurs, i.e. the reaction mechanism. The reaction mechanism consists of a series of simple reactions, known as elementary steps leading from reactants to the products.

Example



$$\text{Rate} = k [\text{NO}_2]^2$$

The mechanism of the above reaction is thought to involve two elementary reactions.



- *NO_3 is an intermediate (an unstable radical)
- *NO_3 is produced and used up during the reaction.
- *NO_3 **does not appear in the overall equation**.
- k_1 and k_2 are the rate constants of the **elementary** steps.

What is the rate law related to?

The rate law is related to the mechanism of the reaction! We can predict the rate law for the elementary step without experiment.

Elementary Step	Molecularity	Rate Law
$A \rightarrow \text{products}$	Unimolecular	$\text{Rate} = k[A]$
$A + A \rightarrow \text{products}$ ($2A \rightarrow \text{products}$)	Bimolecular	$\text{Rate} = k[A]^2$
$A + B \rightarrow \text{products}$	Bimolecular	$\text{Rate} = k[A][B]$

The rate law for an elementary reaction is determined by its **molecularity** (the number of molecules involved).

What conditions must a reaction mechanism satisfy?

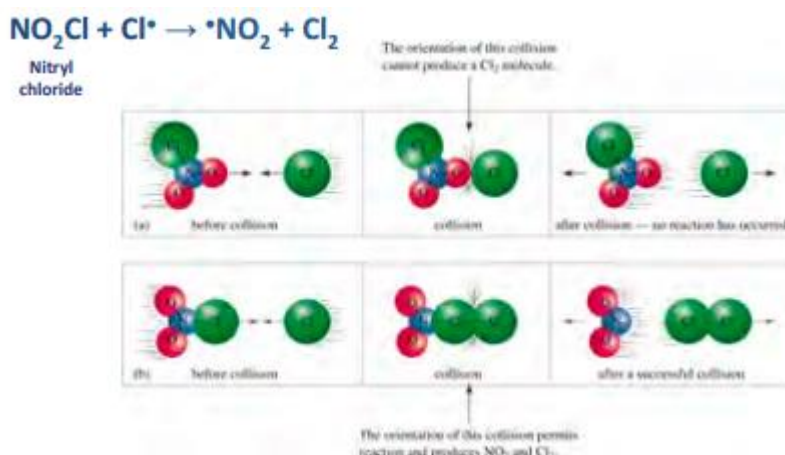
1. The sum of the elementary steps must give the overall balanced equation.
2. The mechanism must agree with the experimentally determined rate law.

What are the requirements for two molecules to react?

1. The molecules must meet

- The rate of reaction will be proportional to the number of collisions per unit time.
- The number of collisions between two molecules will depend on the product of their concentrations.
- E.g. $A + B \rightarrow C$ $\text{rate} = k[A][B]$ it depends on both concentrations

2. The molecules must have the right orientation



3. The molecules must have sufficient combined energy

- The molecules must collide with sufficient kinetic energy to overcome the activation energy (E_a).

What is the relationship between internal energy and enthalpy?

- The general definition for a change in enthalpy:

$$\Delta H = \Delta U + \Delta(pV)$$

- For reaction involving only condensed phases, expansion is minor

$$(\text{i.e. } \Delta H \approx \Delta U)$$

- For a reaction involving an ideal gas:

$$\Delta H = \Delta U + \Delta(nRT)$$

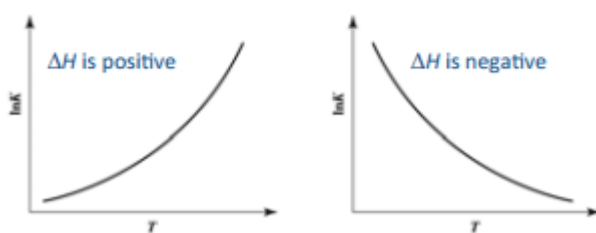
at constant temperature:

$$\Delta H = \Delta U + RT\Delta n$$

(Δn = moles product gas – moles reactant gas)

What is the relationship of reaction rate with temperature (Arrhenius equation)?

van't Hoff Equation	Relationship between rate and equilibrium constants
$\ln K = -\frac{\Delta H}{RT} + \frac{\Delta S}{R}$	$K = \frac{k_f}{k_r}$
Dependence of equilibrium constant K on temperature	Relationship between rate constants of a reversible reaction and the equilibrium constant for a simple one step reversible reaction*



Arrhenius equation represents how the rate of reaction changes with temperature.

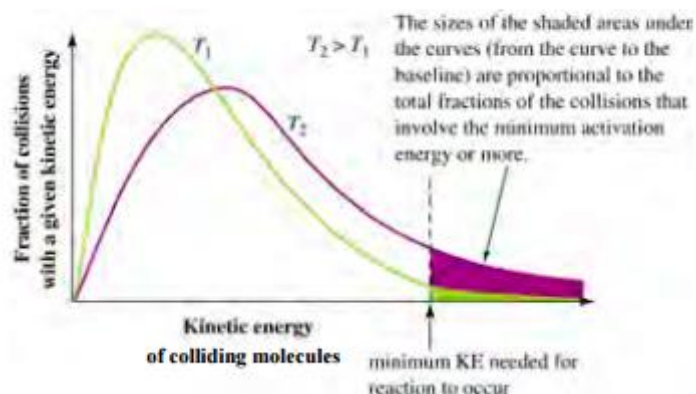
$$k = Ae^{\frac{E_a}{RT}}$$

A = frequency factor. Probability that a collision which could lead to a reaction occurs (temperature independent)

k = rate constant

- E_a/RT = the probability that a collision has enough energy to react (dependent on temperature)

As the temperature increases, the fraction of collisions with sufficient kinetic energy to react increases exponentially.



What is the effect of temperature on the energy at a collision?

At a higher temperature, the kinetic energy of colliding molecules is skewed upward so more molecules have an energy above the required activation energy for reaction.

How can we use the Arrhenius equation to determine the value of k at two different temperatures?

$$\ln \frac{k_2}{k_1} = -\frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

What is a transition state?

Reactions go from reactants to products via a transition state, or activated complex.

What is a catalyst?

Catalysts are substances which speed up chemical reactions without themselves being consumed.

They:

- Increase the rate of a reaction
- Lower the activation energy required for the reaction to proceed
- Do not affect the position of the equilibrium

What are the two types of catalyst?

HOMOGENEOUS CATALYSTS – exist in same phase as the reactants

HETEROGENEOUS CATALYSTS – exist in a different phase to reactants (e.g. a solid catalyst for a gas phase reaction).

MODULE 2: ORGANIC CHEM