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Why Chemical Reactions Happen

What makes a reaction go?

Some reactions appear to be spontaneous, while some do not happen. We use a '→' to show that the reaction almost goes to completion, e.g $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$.

Many chemical reactions result in significant amounts of products and reactants; hence the real question is what determines the position of equilibrium.

Exothermic reactions tend to form stable products. However, we cannot conclude that this is the main driving force between reactions, because some reactions only occur when we supply a lot of heat. **Entropy** is more important.

Entropy and the Second Law of Thermodynamics

The Second Law states that **the entropy of the universe must increase**. However, we can split the universe into a system and the surroundings, and for one of these, local entropy can decrease.

$$\Delta S_{universe} = \Delta S_{system} + \Delta S_{surroundings}$$

Exothermic reactions give heat out to the surroundings so entropy increases. Endothermic processes decrease the entropy of the surroundings. However, these endothermic processes do not violate the Second Law because they are sufficiently exothermic.

However, this is not enough to explain the situation. Water freezing to ice is always exothermic and hence always increases the entropy of the surroundings, yet this freezing is only spontaneous below 0˚C.

The role of temperature

Supplying heat to a hotter object causes a smaller increase in entropy than supplying the same amount of heat to a cooler object. This can be written symbolically:

$$Entropy\ change\ of\ surroundings\ = \frac{heat\ absorbed\ by\ surroundings}{temperature\ of\ the\ surroundings}$$

$$\Delta S_{surr} = \frac{q_{surr}}{T_{surr}}$$

The units of entropy are thus J K^{-1}, though we often quote J K^{-1} mol^{-1}. We can also calculate absolute entropy. We do this by summing very small exchanges of heat energy at different temperatures, then sum this over all of the particles.

$$S^{\ominus} = \sum_{k=1}^{N} \int \frac{dq_k}{T} dT$$

Making ice

For water → ice, the surroundings lose heat and ΔS_{sys} = 22.0 J K^{-1} mol^{-1}.

The heat of fusion of to melt ice is 6010 J mol^{-1}, so when water freezes, 6010 is given out to the surroundings.

At 5°C, $\Delta S_{surr} = \frac{q}{T} = \frac{6010}{278} = 21.6$ J K^{-1} mol^{-1}.

$$\begin{aligned} \Delta S_{univ} &= \Delta S_{sys} + \Delta S_{surr} \\ &= 21.6 - 22.0 \\ &= -0.4 \text{ J } K^{-1}\ mol^{-1} \end{aligned}$$

Since this value is negative, the process cannot happen because of the Second Law.

However, if we repeat this calculation at –5°C, we get a value for ΔS_{univ} of 0.4 J K^{-1} mol^{-1}, which is permitted by the Second Law.

This is why we put ice in a freezer: the cold environment means that the entropy change of the surroundings is large enough to compensate for the entropy decrease of freezing water.

Exothermic and endothermic reactions

Reactions which involve an increase in the entropy of a system are not necessarily exothermic. We can have a situation where the entropy of the surroundings decreases while the entropy of the system increases.

This can also explain reactions such as $MgO(s) \rightarrow Mg(g) + O(g)$.

The reaction is highly endothermic, so it causes the entropy of the surroundings to decrease. At higher temperatures, this decrease will be less.

Gibbs free energy

Because $q_{surr} = -q_{sys} = -\Delta H$, and we assume that $T_{sys} = T_{surr}$, we can write:

$$\Delta S_{univ} = -\frac{\Delta H}{T_{sys}} + \Delta S_{sys}$$

Note that all quantities pertain only to the system. We can rearrange this, then compare with the formula for the Gibbs free energy of the system.

$$-T\Delta S_{univ} = \Delta H - T\Delta S_{sys}$$
$$\Delta G = \Delta H - T\Delta S_{sys}$$

$$\therefore \Delta G = -T\Delta S_{univ} \Longrightarrow \Delta S_{univ} = -\frac{\Delta G}{T}$$

This explains why a negative ΔG is a prerequisite for spontaneity. The definition of ΔG also explains why exothermic reactions (with their negative ΔH values) tend to be favoured.

Ionic interactions

X-ray diffraction experiments have provided evidence that some solid substances exist as lattices of alternating positive and negative ions. We can also measure the **ionic radius**.

Electrostatic interactions

The attraction between oppositely charged objects is the electrostatic or **Coulomb interaction**.

$$force = \frac{q_1 q_2}{4\pi\varepsilon_0 r^2} = k\frac{q_1 q_2}{r^2}$$

The **electrostatic potential energy** between two charges due to the Coulomb interaction:

$$energy = k\frac{q_1 q_2}{r}$$

The **numerical charge** (z_+ or z_-) is a positive integer integer, while the actual charge is the numerical charge times ±*e*. Hence, in an ionic compound:

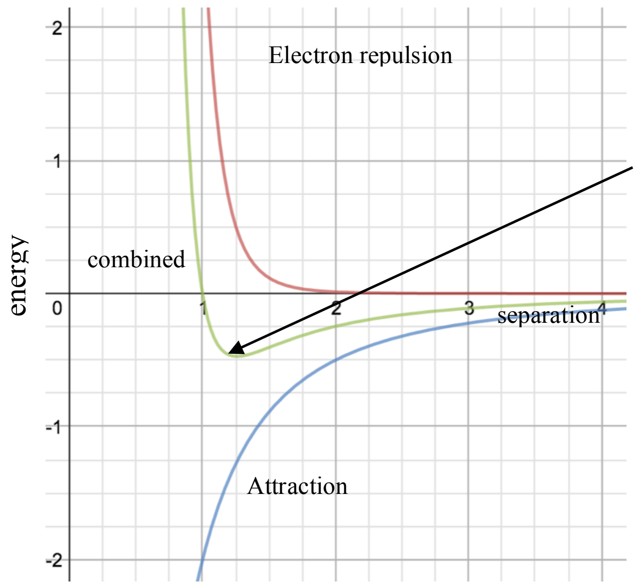
$$energy = -k\frac{z_+ z_- e^2}{r}$$

Because the charges are opposite, the expression is negative so there is a lowering of the energy as r decreases. This means that ion will move towards each other if they are able to.

However, at a certain point, the repulsive interaction between their electrons becomes significant:

$$repulsion\ energy \propto \frac{1}{r^n}$$

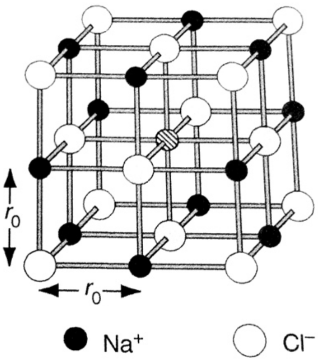
n is some number between 5 and 12, which is around 8 for Na and Cl. The final separation is an equilibrium between the reaction and repulsion.

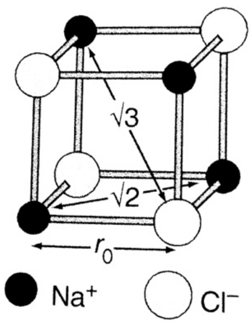


The point of minimum energy is the equilibrium separation.

Forming a lattice

A lattice contains many interacting ions. We will consider the cubic lattice of NaCl from the perspective of one Na^+ ion. The distance between adjacent ions is r_0.





- The Na^+ ion is surrounded by 6 Cl^- at a distance of r_0.

$$E = 6\times\left(-k\frac{z^2e^2}{r_0}\right)$$

- The next closes neighbours are the 12 Na^+ ions at a distance of $\sqrt{2}r_0$.

$$E = 12\times\left(k\frac{z^2e^2}{\sqrt{2}r_0}\right)$$

- Then there are 8 Cl^- ions at a distance of $\sqrt{3}r_0$.

$$E = 8\times\left(-k\frac{z^2e^2}{\sqrt{3}r_0}\right)$$

Hence we note that the signs will alternate, and the distance involves a square root. If we assign a coordinate grid to the lattice, (with our test Na^+ being the origin), we can say:

$$E = -k\frac{z^2e^2e^2}{r_0}\sum_{i=-\infty}^{\infty}\sum_{j=-\infty}^{\infty}\sum_{k=-\infty}^{\infty}\frac{(-1)^{i+j+k}}{\sqrt{i^2+j^2+j^2}}; (i,j,k) \neq (0,0,0)$$

This large sum converges to a value of about 1.75 – the **Madelung constant (M)**. We can also multiply by Avogadro's number to give

$$E = -k\frac{MN_Az^2e^2}{r}$$

However, as showed earlier, we must include a repulsion term for the electrons.

$$E_{total} = -k\frac{MN_Az^2e^2}{r} + \frac{B}{r^n}$$

We can differentiate this and set to zero to find B in terms of the equilibrium separation in terms of r_0. Substituting back into the above gives us the **Born-Landé equation**:

$$Lattice\ enthalpy = -k\frac{MN_Az^2e^2}{r_0}\left(1-\frac{1}{n}\right)$$

The energy of an ion pair can be found with the above formula, with M set to 1.

Kapustinskii noticed that the ratio M/*v*, where *v* is the number of ions in the molecular formula, was more or less constant. This allows for further simplification of the formula.

Applications

For many metals in higher oxidation states, it is possible to form fluoride compounds but not other halides. This is because of the tendency for the halide to decompose, e.g:

$$MX_3(s) \rightarrow MX_2(s) + {}^1\!/_2\, X_2(g)$$

Computing the results reveals that the above decay is exothermic for all halogens except fluorine.

Covalent character

The Born-Landé equation does not produce perfectly accurate results. This is because it assumes completely ionic character. However, looking at X-ray diffraction patterns (which show electron density) reveals that there is some sharing.

For compounds with greater covalent character, the experimental lattice energy is often much larger since our analysis only considered the ionic component.

Electrons in atoms

Electrons exist in quantised states, in main energy levels which contain orbitals.

- The first shell (K) contains the 1s orbital.
- The second shell (L) has the 2s orbital and three 2p orbitals.
- The third shell (M) has the 3s orbital, three 3p orbitals, and five 3d orbitals.

Photoelectron spectroscopy

PES provides a way of probing energy levels. Spectra are produced by irradiating the sample with high frequency, ionising electrons.

energy of photon = energy of ionised electron + ionisation energy of the electron

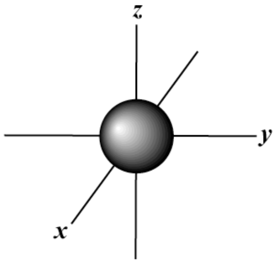
Peaks further to the left represent electrons which are more tightly held.

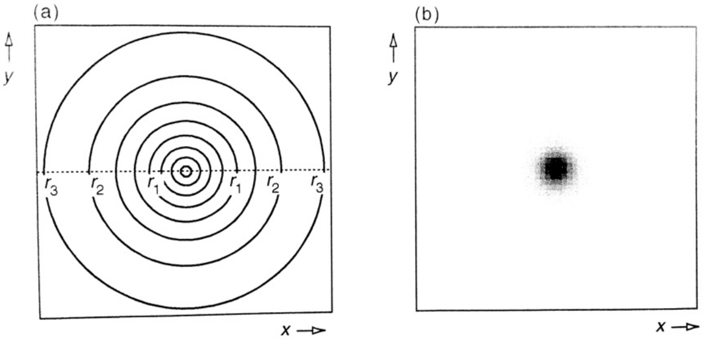
Quantum mechanics

Each energy level has associated with it a **wavefunction**, a function of three variables that returns an amplitude. The **Born interpretation** suggests that at the point *(x,y,z)*, the probability density is proportional to the square of the wavefunction at this point.

We often describe the shape of orbitals in spherical coordinates: $f(r, \theta, \phi)$. For example, the 1s orbital has the following form: $\psi_{1s}(r) = Ae^{-Br}$.

Orbitals can be depicted as contour plots, or as shaded plots, or as surface plots.





Radial distribution functions

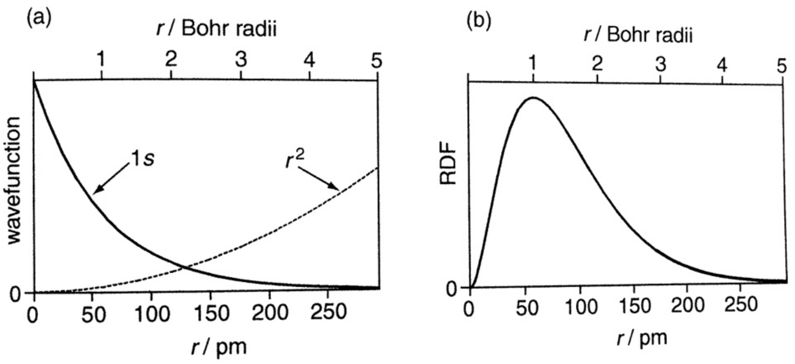
It is useful to know the total probability of finding the electron at a certain distance.

$$\begin{aligned} \mathit{probability\ of\ being\ in\ shell} &= (\mathit{wavefunction})^2 \times \mathit{volume\ of\ shell} \\ &= (\mathit{wavefunction})^2 \times 4\pi r^2 \times \mathit{volume\ of\ shell} \end{aligned}$$

$$\mathit{RDF} = (\mathit{wavefunction})^2 \times 4\pi r^2$$

$$\therefore\ \mathit{probability\ of\ being\ in\ shell} = \mathit{RDF} \times \mathit{thickness\ of\ shell}$$

Although the wavefunction takes a maximum value at *r* = 0, the RDF is zero here because the volume of the shell is zero. While the wavefunction drops off, the RDF has a maximum value – this is the **Bohr radius,** or the lowest energy orbit.



Hydrogen orbitals

Orbitals can be described with quantum numbers.

The principal quantum number, n:

- The principal number is an integer which states the main energy level: $n = 1, 2, 3, \ldots$.
- Its value determines the energy of the orbital, E_n:

$$E_n = \frac{-R_H}{n^2}, where\ R_H\ is\ the\ \boldsymbol{Rydberg\ constant}$$

- As n increases, the energies increase (they become less negative).

The orbital angular momentum quantum number, l:

- Takes values from 0 to $(n - 1)$.
- $l = 0$ is denoted by s, then subsequently values are denoted by p, d, f.

The magnetic quantum number, m_l:

- Takes values from $-l$ to l.
- This corresponds to the number of sub orbitals, e.g there are 5 sub orbitals in the d orbital.

The K shell

- The K shell has $n = 1$. Hence $l = 0$ and there is only one value of m_l.
- The 1s wavefunction is spherical, and the RDF has a single maximum at the Bohr radius.

The L shell ($n = 2$)

- The orbital with $l = 0$ is the 2s orbital, which is just like a bigger 1s orbital.
- However, this 2s orbital has a **radial node** – the wavefunction is zero at a certain distance.
- For $l = 1$, there are three values of m_l: corresponding to $2p_x$, $2p_y$, $2p_z$.
- The $2p_z$ is positive for $z > 0$, and negative for $z < 0$.
- When $z = 0$, the wavefunction is zero for all x and y, so the xy-plane is the **nodal plane.**
- We can describe it in spherical coordinates (separated into an angular part and radial part):

$$\psi_{2p_z}(r) = C[cos\theta][re^{-Dr}]$$

- When θ is 90°, the wavefunction is zero – this results in the **angular node**.
- The other p orbitals have the same radial part, but different angular parts.
- For hydrogen, there are four **degenerate orbitals** in the L shell, that is, there are four distinct orbitals which have the same energy.

The M shell and beyond

- The 3s orbital now has two radial nodes, the 3p has one, the 3d has none.
- The 3d orbitals each have two angular nodes.

Generalising the pattern:

- The number of radial nodes is (n – l – 1).
- The number of angular nodes is l.
- The total number of nodes is (n – 1).

Atoms with more than one electron

Having to include the repulsive effects of an extra electron immediately means that we cannot find analytical solutions. We thus have to make **orbital approximations**, which consider each electron to experience only an effective nuclear charge, Z_{eff}.

For example, although each electron in a helium atom experiences a charge of +2, because of the repulsion with the electrons, the Z_{eff} is 1.4 (value determined numerically with a computer).

Changing Z_{eff} affects the size of the orbitals and their energies. The contraction effect is inversely proportional to Z_{eff}, while the energy is proportional to the square of $(Z_{eff})^2$.

$$E_n = -\frac{Z_{eff}^2 R_H}{n^2}$$

Because of the extra electrons in the atom, higher orbitals are not degenerate. Charge shielding occurs, yet the probability that electrons penetrate this screen is higher for the 2s orbital, so the 2s has lower energy.

The electronegativity is also dependent on the energy of the orbitals. Generally, the atom with lower energy orbitals will end up with the partial negative charge.

Electrons in simple molecules

When two H atoms are brought together, there will be a minimum energy at which the bond forms. The difference in energy between the bottom of the PE curve and the separated atoms is the **dissociation energy**.

Molecular orbitals for H_2

The Schrödinger equations have no analytical solutions for multiple atoms, wo we have to approximate the solution. One way of constructing MO is the **linear combination of orbitals** (LCAO) method:

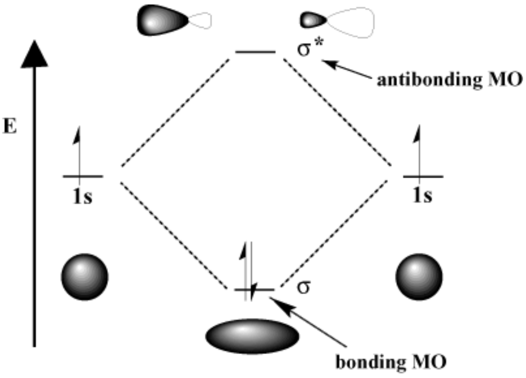
$$MO = c_1 \times (\text{AO on atom 1}) + c_2 \times (\text{AO on atom 2}).$$

When two AOs combine, they form two MOs. For one of them, c_1 and c_2 have the same sign, for the other, c_1 and c_2 have opposite signs.

$$\begin{aligned} bonding\ MO &= 1s_1 + 1s_2 \\ anti-bonding\ MO &= 1s_1 - 1s_2 \end{aligned}$$

When the coefficients are both positive, constructive interference occurs: for this MO, the electron density is thus between the two nuclei. However, when we subtract, there will be a point for which the wavefunction is zero – resulting in a nodal plane.

The bonding MO has lower energy than the AOs from which it formed, since electrons are attracted by both nuclei. The anti-bonding MO has a higher energy than the AOs: the electron density pattern reveals that electrons are repelled from the internuclear region.



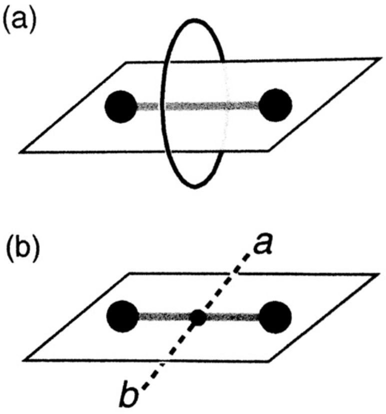
This MO diagram shows the shapes of the orbitals (with shaded being positive), however, it is normally only necessary to show the lines.

We can use this diagram to make a few predictions concerning stability, since we know that occupation of the bonding MO favours bond formation and occupation of the anti-bonding MO disfavours formation.

- H_2^+ should be stable since it has an electron in the bonding MO. However, since there is only one electron, the bond is weaker than that in H_2.
- On the other hand, He_2 is unstable since it's anti-bonding MO is full.
- However, since He_2^+ has two electrons in the bonding MO and only one in the anti-bonding MO, it does exist.

Symmetry labels

MOs are given labels describing their symmetry w.r.t. the molecule. For a diatomic molecule, there are two symmetry labels.



a) If we traverse a circular path perpendicular to the internuclear axis, and the wavefunction does not change sign, we give it the label σ, If the circular path crosses the nodal plane, we use π. We can distinguish the anti-bonding MO with an asterisk.

b) A homonuclear diatomic also possesses a **centre of inversion** symmetry. If, going from a to b, the sign is the same, we use the label *g* ('gerade' means 'even' in German). If the sign changes, we use *u*.

This notation is sometimes combined with the notation for the AOs from which the MOs are derived. For example, the bonding MO in H_2 is labelled $1s\sigma_g$.

General rules for MOs

Only AOs with similar energies interact to a significant extent

- The extent of the interaction is the amount by which the bonding MO goes down in energy (or the anti-bonding MO increases).
- As the energy difference grows, the extent of interaction decreases

The size of AOs is important

- AOs can only interact where their wavefunctions overlap.
- The extent of the overlap is dependent on the size of the orbitals and how close they can get to each other, e.g two 3s orbitals overlap less than two 2s orbitals.

AOs must be of the correct symmetry to interact

- For example, a $2p_x$ orbital will not interact with a 1s orbital on the *z*-axis because the region of constructive overlap cancels with the region of destructive overlap.
- However, a $2p_z$ orbital can overlap with the 1s without any problems.

When *n* AOs interact, the same number of MOs is formed.

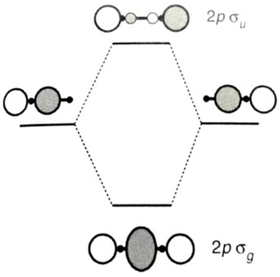
- Two AOs form a bonding and anti-bonding MO.

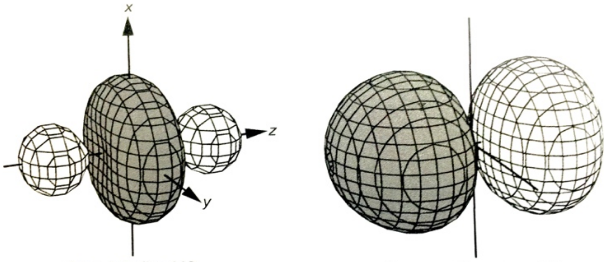
MOs for homonuclear diatomics

We will now consider the MOs for the period 2 elements. The 1s shell is irrelevant, so we only have to examine the 2s and 2p shells. We can disregard the node in the 2s shell because the interaction region is much further away. Also, the 2s2p interactions are negligible.

Two 1s AOs can combine to give a σ bonding orbital (σ_g) or a σ anti-bonding orbital (σ_u). The same thing happens for the 2s orbitals. However, 2p orbitals can interact to give both σ and π.

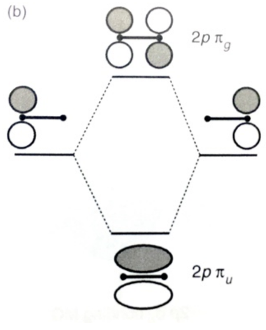
When p orbitals overlap head on, they form σ orbitals. Constructive interference results in a bonding MO with high electron density in between the the two nuclei. Subtracting the AOs results in electron density being pushed away from the internuclear region.



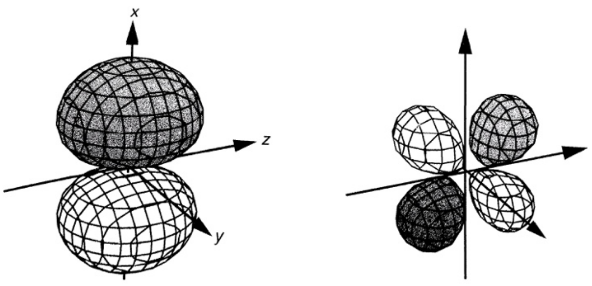


2p σ_g bonding MO

2p σ_u anti-bonding MO



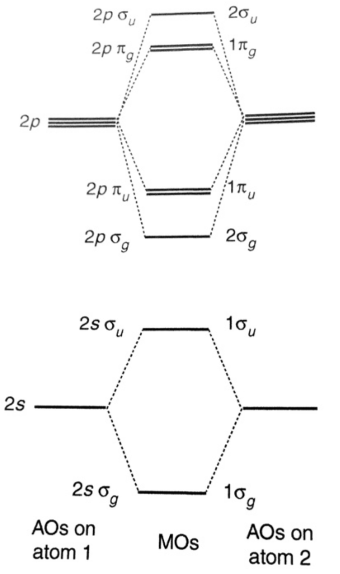
However, when two 2p orbitals interact parallel to each other. There is then electron density above and below the internuclear axis, resulting in π symmetry. The bonding MO has *u* symmetry.



2p π_u bonding MO

2p π_g anti-bonding MO

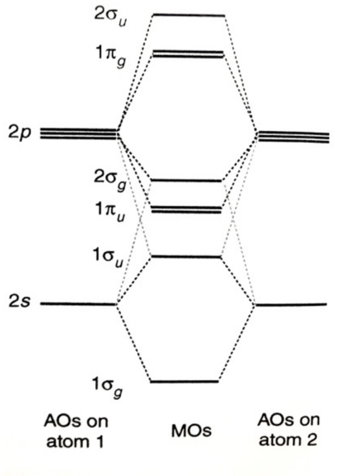
Note that one of the other 2p orbitals can also overlap with one from the other atom, giving a total of two bonding and two anti-bonding π orbitals.



We can thus illustrate the full MO diagram for homonuclear diatomics.

Sigma interactions are generally stronger than pi interactions, which is why the 2p part is above the 2s part.

There are three lines for the 2p part because there are three 2p sub orbitals. Also, the p orbitals can form one sigma and two pi bonds (the pi bonds are a double line because pi bonding can happen in two axes).



This model ignores interactions between 2s and 2p. This is acceptable for oxygen and fluorine, but we cannot ignore the effect for Li to N. The most significant effect of this **s-p mixing** is to make the $2p\sigma_g$ MO less bonding (raise its energy) and the $2s\sigma_u$ less anti-bonding (lower its energy). For Li to N, the shift of the $2p\sigma_g$ is enough to move it above the $2p\pi_u$ MO. We omit the p or s labels (because it's a mix).

We can now imagine slotting in the appropriate number of electrons depending on the diatomic:

- Slotting two electrons into the $1\sigma_g$ orbital gives lithium, and since this is a bonding orbital, there should be a bond.
- In Be_2, since there are two electrons in the antibonding orbital, there should be zero bonding. In fact, there is some bonding, because the antibonding orbital is less antibonding than the bonding orbital is bonding.
- In B_2, each of the $1\pi_u$ takes one electron (Hund's rule) with parallel spins. This results in **paramagnetism.**
- The degree of bonding can be expressed in terms of **bond order**:

$$bond\ order = \frac{1}{2}(number\ of\ bonding\ electrons - number\ of\ antibonding\ electrons)$$

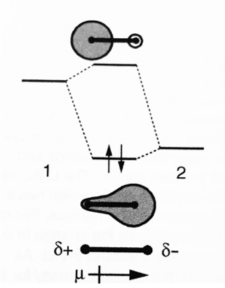
- C_2 and N_2 have bond orders of 2 and 3 respectively.

- O_2 and F_2 do not show significant sp mixing. The 12 electrons in an O_2 molecule leave two unpaired electrons in the $1\pi_g$ antibonding MOs. This means that O_2 is paramagnetic.

Heteronuclear diatomics

When one of the species is different, there will be an energy mismatch.

This means that the molecular orbital is unevenly distributed between the atoms, so the bond is polarised.



Electrons in larger molecules

Finding the MOs for larger molecules tends to get complex, so computers are used. These programmes can, using brute force, also find the positions of atoms resulting in minimum energy (to find bond lengths and equilibrium geometry).

However, given that we often know the shape (e.g through VSEPR), we can attempt to determine MOs, which then help us figure out reactions.

Two-centre, two-electron bonds

A bond between two atoms that involves two electrons (i.e a single bond) is called a 2c-2e bond. This is the basis of the simple dot/cross model.

It is difficult to combine the simple dot and cross picture with hybridisation, because the shapes of the orbitals do not work. Hence we introduce the idea of **hybrid atomic orbitals** (HAOs), such that hybrid orbitals allow for the correct geometry.

HAOs are formed in the same way as MOs, but instead, we combine AOs on the same atom. In methane, the carbon atom forms four sp^3 orbitals, each bonding to a single hydrogen.

Not all of the p orbitals have to hybridise, sp^2 and sp hybridised orbitals leave behind normal p orbitals. As we increase the ratio of s, we increase the bond angle. This explains the extra repulsion from lone pairs: they have an extra s-character.

NH_3 molecules undergo constant ‘inversion’, wherein they go from being trigonal pyramidal to planar then back again, i.e the hybridisation changes from $sp^3 \rightarrow sp^2 \rightarrow sp^3$.

Hybrid AOs in diatomics

Consider N_2. We suppose that the nitrogen is sp hybridised, with the hybrids pointing along the z axis – one forms a sigma bonding MO and the other forms a lone pair. The remaining p orbitals form two π bonds.

This idea is consistent with the MO description: sp mixing causes the $2\sigma_g$ MO less bonding (raise its energy) and the $1\sigma_u$ less anti-bonding. This means they contribute little to bonding, so they are non-bonding pairs.

Simple organic molecules

- In ethane, both carbons are tetrahedral so clearly we use sp^3. Each overlap of orbitals results in a σ bonding and σ antibonding orbital. Putting electrons in this forms a 2c-2e bond.
- Filling all of the σ bonding MOs in ethene accounts for 10 of the 12 electrons. The unhybridised carbon $2p_z$ AOs overlap to form a π bonding and π antibonding orbital, which form out of the plane of the molecule. Twisting the molecule breaks this overlap, which is not energetically viable.
- For aldehydes and esters, the oxygen forms sp^2 hybridised orbitals, with two lone pairs. The oxygen AO is lower in energy, so oxygen contributes more to the bonding MO while carbon contributes more to the antibonding MO.

Hybridisation vs delocalisation

When we hybridise, the resulting orbitals are identical. However, the MO approach does not require this. Experiment agrees with this: the lone pair orbitals aren't identical.

The HAO approach allows us to qualitatively understand molecules without requiring computer modelling. Normally, we use the HAO approach to describe σ bonding, and MO theory to describe the π electrons.

Orbital energies

To understand chemical reactions, we need to identify the **highest occupied MO (HOMO)** and the **lowest unoccupied MO (LUMO)**.

The general order of orbital strength:

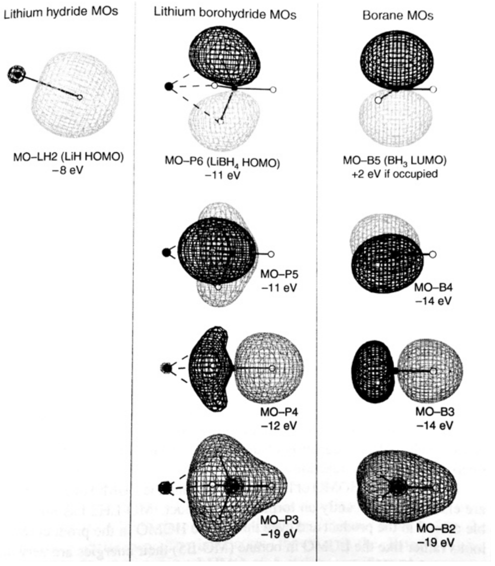
$$\sigma < \pi < \text{lone pairs} < \pi^* < \sigma^*$$

Reactions

One way H_2 could be formed is:

$$H^+ + H^- \rightarrow H_2$$

The reaction does not give an ion pair: the pair of electrons is ultimately shared between the two, resulting in a lower energy. This simple idea can be applied to more complicated interactions.



For example, the formation of lithium borohydride from lithium hydride and borane:

$$LiH + BH_3 \rightarrow LiBH_4$$

P3, P4, and P5 are very similar to the borane MOs B2-B4.

However, the HOMO and LUMO have a big energy difference, and this results in the HOMO of $LiBH_4$ to have a much lower energy (so the reaction is exothermic).

We can use curly arrows to show this:

HOMO-LUMO interactions

Interactions of empty orbitals does not affect the energy. Interactions of filled orbitals only has a small effect, since the decrease in energy for the in-phase combination is cancelled by the increase from the out-of-phase combination. It is the HOMO-LUMO interactions that are often the most favourable.

While there are two possible HOMO-LUMO interactions when two reactants come together, but the more significant one tends to be the one with the smaller energy gap.

To understand the outcome of a reaction:

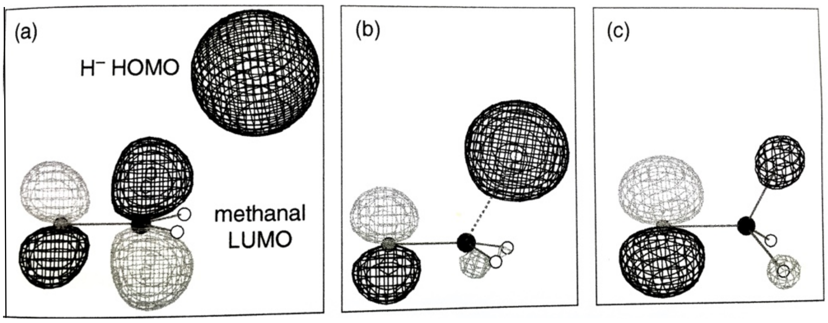
- Identify the molecule with the higher HOMO
- Consider the reaction between his and the LUMO of the second molecule.

Nucleophilic addition to carbonyl

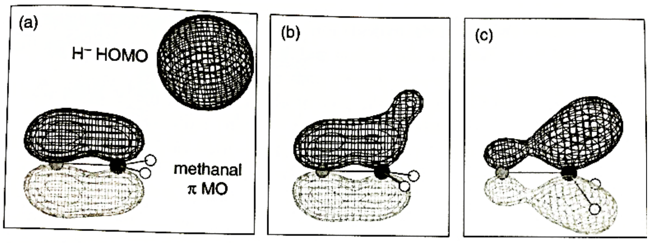
In these reactions, a nucleophile attacks the carbonyl group, forming a new bond and simultaneously breaking the C-O π bond.

We will consider the simplest such interaction: the attack of a hydride ion on methanal.

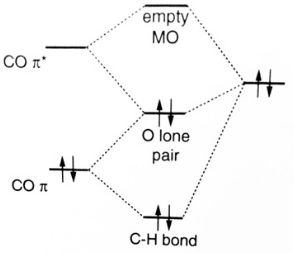
The HOMO is from the hydride ion, while the LUMO is the π* in methanal. The in-phase interaction shows that electron density is transferred from the hydride to the oxygen (the out-of-phase interaction forms an empty MO).



However, the HOMO also interacts with the CO π MO. In this interaction, the C-O interaction weakens, while the C-H strengthens. Eventually, this π bond disappears.



The interaction of these three orbitals gives rise to three MOs since orbital number is conserved. The total energy of the resulting electrons is lowered.



For the hydride to interact with the π and π* orbitals, it must approach with the right geometry. If the hydride were to approach vertically or horizontally, the constructive interference would cancel with the destructive interference.

<u>Borohydride reduction of a ketone</u>

Instead of adding H^- directly, we often use $NaBH_4$. This is a reducing agent, because it adds a hydrogen. The HOMO is the bonding between the boron and hydrogens. The LUMO is the π* in the ketone.

The overall mechanism:

Nucleophilic attack on C=C

A C=C bond is not attacked by nucleophiles in the same way as is C=O.

- Unlike the carbonyl, the C=C bond is not polar.
- The C=C π and π* MOs are higher in energy than those of C=O. This normally results in a poorer energy match between the C=C π* and the HOMO of an attacking nucleophile.
- If a nucleophile did attack, it would leave a charge on the carbon, which is not favourable.

Nucleophilic substitution

Nucleophiles can also donate electrons into σ* MOs, resulting in the breaking of a single bond (i.e a group leaving). This explains why it is substitution rather than addition.

$$HO^- + CH_3Br \rightarrow CH_3OH + Br^-$$

There are two possible σ* MOs: the C-H and the C-Br. There is a better energy match in the AOs between C and H, so the C-Br bond is weaker. This will be true for any other electronegative element, explaining why organic chemistry can ignore C-C/C-H bonds most of the time.

The LUMO is the C-Br σ* MO. The HOMO is the lone pair on the oxygen in the OH^-. The OH^- is attracted to the carbon, because it has a partial positive charge (since the Br is more electronegative). The OH^- attacks from the back, where there is a good orbital match. A new σ bond forms while the C-Br σ bond simultaneously breaks.

During the reaction, the carbon centre undergoes an inversion. This becomes detectable when there are three different groups around the carbon, because of stereoisomerism.

Nucleophilic attack on acyl chlorides

Sometimes, the molecule which the nucleophile attacks has both σ* and π*.

A simple example is the reaction between a hydroxide ion and ethanoyl chloride. The hydroxide can either attack the π* of the C=O to give a tetrahedral species or it could attack the C-Cl σ* to form ethanoic acid. While the final product is, indeed, ethanoic acid, it turns out that the nucleophile attacks the π*.

- The LUMO for ethanoyl chloride is the π*.
- There is a better match between the HOMO of the hydroxide and the C-O π* MO, because the C-Cl σ* is higher in energy than the C-O π*.
- It is easier for the nucleophile to approach the π* with the correct orientation. Approaching the σ* requires approaching from the back, which is hindered by the methyl group and oxygen which are in the same plane as the σ* MO.

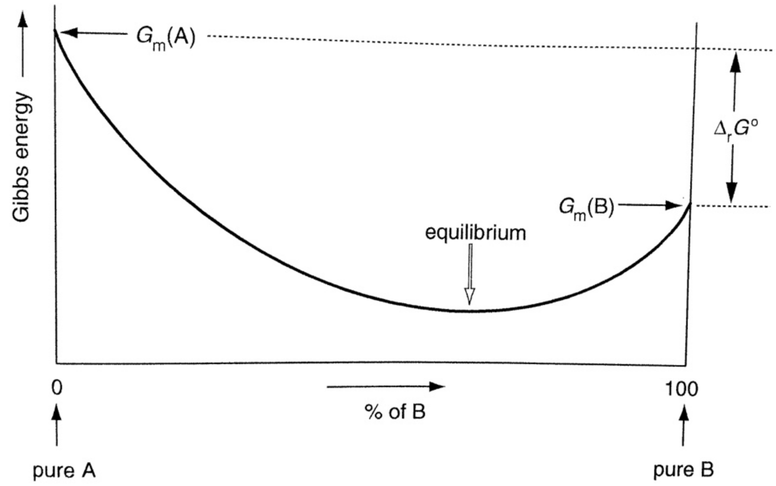
The tetrahedral species is then converted to the final products: the HOMO is on the oxygen, the LUMO is the C-Cl σ*. The overall mechanism:

This mechanism is also used in esterification.

Equilibrium

In any equilibrium reaction, the equilibrium mixture of reactants and products must have lower Gibbs energy than either pure products or pure reactants. Plotting a graph of Gibbs energy against composition, there will be some minimum between pure reactants and pure products, and this minimum will be the position of equilibrium.

Consider the homogeneous equilibrium between A and B, $A \rightleftharpoons B$.



At the far left and far right, the mixture is pure, so the Gibbs energy is equal to the molar Gibbs energy of the substances, G_m(A) and G_m(B). Suppose we start with pure A, then a small amount of A converts to B. There will be a small enthalpy change and a small entropy change, the latter is more significant.

- Since the two substances likely have different molar entropies, converting some A to B will result in an entropy change.
- When we mix some B into A, the disorder of the system increases. This factor is the more dominant of the two, so system entropy always increases.

Because $\Delta G = \Delta H - T\Delta S_{sys}$, this increase in entropy corresponds to a decrease in the Gibbs energy. However, if we start with pure B, we can apply the same arguments to reason that the Gibbs energy must increase in the reverse reaction. Hence, for the curve to join up, there must be a minimum.

The position of equilibrium depends only on the difference between G_m(A) and G_m(B). The greater the difference, the more skewed equilibrium will be.

We can thus relate the equilibrium constant, K, to the amounts of reactants and products.

$$\Delta G^{\theta} = -RT \ln K$$

If ΔG^{θ} is negative, (A has higher molar Gibbs energy than B), lnK will be positive so K > 1.

It turns out that this result is more general, and can be applied to any equilibrium reaction. We can calculate the change in Gibbs energy using $\Delta G^{\theta} = \Delta H^{\theta} - T\Delta S^{\theta}$, where all the values are standard.

Note that if ΔG^{θ} is negative, it does not mean that the reaction won't take place. Rather, some reactant can still be converted to some product, which lowers the Gibbs energy.

Standard state and standard changes

The **standard state** is basically the pure substance, or a solution at unit concentration. However, standard state does not imply a certain temperature – we must also state the temperature.

Standard changes refer to balanced chemical reactions, with all substances being in standard state. Also note that ΔG^{θ} is the standard energy change for the complete reaction (meaning that it is often a hypothetical quantity). ΔH^{θ} and ΔS^{θ} values can be found by using H_f and S^{θ} values.

Changing the position of equilibrium

Temperature

Both the equations relating ΔG^{θ} to the equilibrium constant and the definition of ΔG^{θ} involve temperature explicitly (entropy and enthalpy also depend on temperature, but the relationship is small enough to be negligible).

We can equate the two expressions for ΔG^{θ} :

$$-RT \ln K = \Delta H^{\theta} - T\Delta S^{\theta}$$

$$\ln K = -\frac{\Delta H^{\theta}}{R}\left(\frac{1}{T}\right) + \frac{\Delta S^{\theta}}{R}$$

For an endothermic reaction, ΔH^{θ} is positive so $\Delta H^{\theta}/RT$ is negative. This means that increasing the temperature causes lnK and K to increase, favouring the products. We can make a similar analysis for an exothermic reaction, concluding that increasing temperature favours reactants.

This analysis provides a quantitative basis for Le Chatelier's principle.

Concentration

In order to maintain K constant, removing some reactant/product will force the reaction to produce more of that.

Coupling reactions

If we are interested in a reaction that has a positive ΔG^{θ} (i.e K < 1), we can form more products by coupling the reaction to another reaction with negative ΔG^{θ}. For reactions to be coupled they must be able to influence one another. For example, polymerisation of amino acids into proteins has an unfavourable ΔG^{θ}, but in nature it is coupled with the hydrolysis of ATP into ADP.

$$\text{ATP} + \text{H}_2\text{O} \rightarrow \text{ADP} + \text{phosphate}$$

This has a very favourable ΔG^{θ}, making the polymerisation feasible. This hydrolysis drives many other reactions, which is why ATP is considered the source of energy for life.

Rates of reaction

Just because a reaction has a favourable ΔG^{θ}, this doesn't mean that the reaction will proceed quickly. For a reaction to occur:

- The molecules must collide. This means that the rate is proportional to the product of the concentrations. The faster they move, the more likely they are to collide. The larger the molecules are, the higher chance there is that they will collide.
- The molecules must collide with the correct orientation, i.e there is a **steric requirement**. About 1 in 100,000 random collisions have the correct geometry.
- There must be sufficient energy to overcome the energy barrier (i.e activation energy). About 1 in a billion collisions are sufficiently energetic. Not all reactions have activation energies, for example the lattice enthalpy of formation and free radical termination.

We can understand activation energies by plotting energy level diagrams, with energy on the *y*-axis and the **reaction coordinate** on the *x*-axis. The reaction coordinate is a combination of distances and angles to measure the progress of a reaction. The transition state will have a higher energy.

The Arrhenius equation

The rate of a reaction has a strong dependence on temperature, following the relation:

$$k = Ae^{-E_A/RT}$$

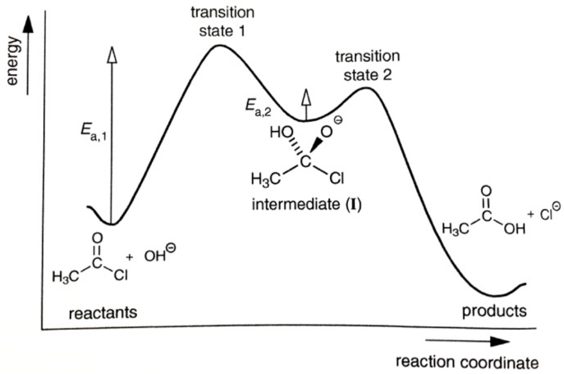
From Kinetic theory, it can be shown that the fraction of collisions with $E \geq E_A$ is given by $e^{-E_A/RT}$. A is therefore the rate constant if the activation energy were to equal zero.

Elementary and complex reactions

Some reactions take place as they are written (e.g S_N2 reactions) – these are **elementary**. However, some reactions are **complex**, they take place in a series of elementary reactions (a **mechanism**), and form intermediates.

The rate expressions for elementary reactions can be written down directly from the chemical equation. However, complex reactions require consideration of intermediates and rate-determining steps. Intermediates are different to transitions states, in that they are ‘real’ molecules and can sometimes be isolated. They are always found at minima of energy profiles.

Using an aforementioned example of nucleophilic attack on acyl chlorides:



The rate of the first elementary step is lower because there is a higher activation energy. The first step is therefore the rate-determining step.

Reversible reactions

The energy profile of the reverse reaction is a mirror-image of that of the forward reaction. We can also relate K_C to the rate constant. Suppose we have the reversible reaction A + B ⇌ P + Q. We can then say that

Rate of forward reaction = $k_f[A][B]$ Rate of reverse reaction = $k_r[P][Q]$

At dynamic equilibrium, these are equal.

$$k_f[A][B] = k_r[P][Q] \Rightarrow \frac{k_f}{k_r} = \frac{[P][Q]}{[A][B]} \Rightarrow K_{eq} = \frac{k_f}{k_r}$$

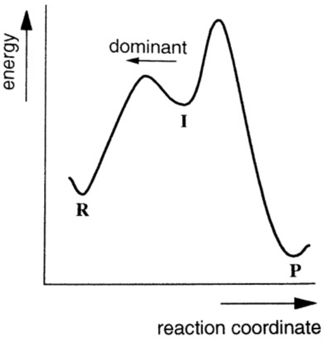
We can then rewrite the expression using the Arrhenius equation, noting that the total energy change of the reaction is equal to the difference of the activation energies of the forward and reverse reactions.

$$K_{eq} = \frac{A_f}{A_r} e^{-\Delta E/RT}$$

This conclusion is very similar to one we made earlier, using thermodynamics.

Pre-equilibrium

When an intermediate forms, we naturally assume that it will be converted to products. However, this is not always the case. In the following energy profile, the activation energy to be converted back to reactants is lower.



The reactants and the intermediate are rapidly interconverting, while a small amount of the intermediates will form products. However, this amount is so small that we can effectively assume that an equilibrium exists between reactants and the intermediate – this is the **pre-equilibrium hypothesis**.

Example: oxidation of methanoic acid with bromine

In aqueous solution, methanoic acid is oxidised to CO_2 by Br_2:

$$HCOOH + Br_2 + 2H_2O \rightarrow CO_2 + 2Br^- + 2H_3O^+$$

Experimentally, the rate expression is:

$$Rate = k\frac{[Br_2][HCOOH]}{[H_3O^+]}$$

As the concentration of H_3O^+ increases, the rate decreases. This gives us a clue that H_3O^+ is the product of one of the steps. One possible mechanism is as follows.

$$HCOOH + H_2O \rightleftharpoons HCOO^- + H_3O^+ \quad \text{(step 1 and –1)}$$

$$HCOO^- + Br_2 \rightarrow \text{products} \quad \text{(step 2)}$$

Step 1 and –1 are the dissociation of acid; the equilibrium constant is the ratio of the forward and backward reactions: $K_A = \frac{k_1}{k_{-1}}$

We also know that

$$[HCOO^-] = \frac{K_A[HCOOH]}{[H_3O^+]}$$

The rate of consumption of bromine is the rate of the step 2.

$$Rate = k_2[HCOO^-][Br_2] = \frac{k_2K_A[HCOOH][Br_2]}{[H_3O^+]}$$

This is equal to the experimental rate law, with:

$$k_{eff} = k_2K_A = k_2\frac{k_1}{k_{-1}}$$

This agrees with the pre-equilibrium hypothesis, which generally holds true when the first step is fast compared to further reactions of the intermediate, e.g when the intermediate is formed by protonating or deprotonating the reactants.

Apparent activation energies

For any reaction, we can determine the activation energy graphically using the Arrhenius equation (this gives us the **apparent activation energy**). For a complex reaction, it is possible to link the apparent activation energy to the activation energies of the elementary steps.

For a reaction where the rate determining step is the first step, the apparent activation energy is the same as the activation energy of the first step. For a complex reaction, after determining the effective rate constant in terms of rate constants of the elementary reactions (and their reverse reactions), we simply substitute in the Arrhenius equation, and combine the logs.

$$k_{eff} = \frac{k_2k_1}{k_{-1}} = \frac{A_2A_1}{A_{-1}}\exp\left(-\frac{E_{a(2)} + E_{a(1)} - E_{a(-1)}}{RT}\right)$$

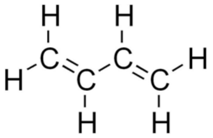
So $E_{a(apparent)} = E_{a(2)} + E_{a(1)} + E_{a(-1).}$

The general result is that the apparent activation energy is the energy difference between the reactants and the **highest energy transiti=on state.**

Bond conjugation

The idea of hybrid atomic orbitals are generally quite useful; the localised picture of bonding generally works fairly well. However, there are some molecules for which the localised approach doesn't work, molecules with extended π systems, like in benzene.

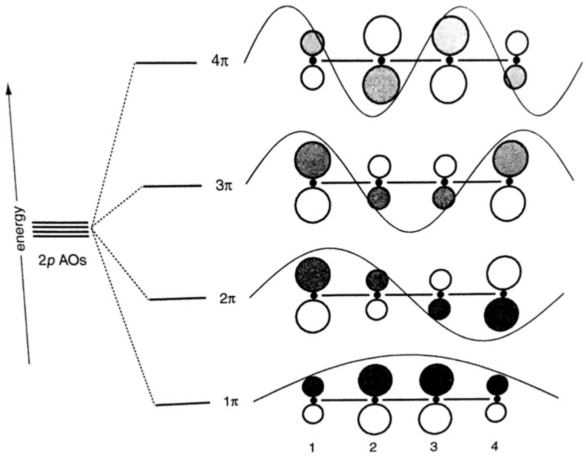
We will examine the electron system in a row of p orbitals, e.g in butadiene



The σ framework uses 18 of the 22 electrons, leaving four electron to go into the π system, which consists of four MOs made from the four 2p orbitals. Recall that an MO can be found using the LCAO method:

$$MO = c_1AO_1 + c_2AO_2 + c_3AO_3 + c_4AO_4$$

Note that some of these coefficients are negative.



The different sized orbitals refer to the different magnitudes of the coefficients.

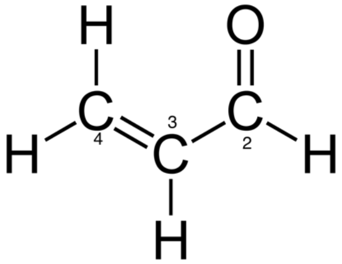
- In the lowest energy MO, 1π, constructive interference occurs because all the 2p orbitals are positive. There is no node.
- In the 2π MO, there is one pair of destructive interference between AO_2 and AO_3. However, there are two pairs of constructive interference, so this is still a bonding MO. There is now one node.
- 3π and 4π have more destructive interference than constructive, so are antibonding MOs. These have two and three nodes respectively.

The coefficients of the MOs each follow a sine curve which can be drawn over the molecule. The actual values of the coefficients for the *j*th atom in the *k*th MO for a species with N overlapping AOs are given by:

$$\sin\left(\frac{jk\pi}{N+1}\right)$$

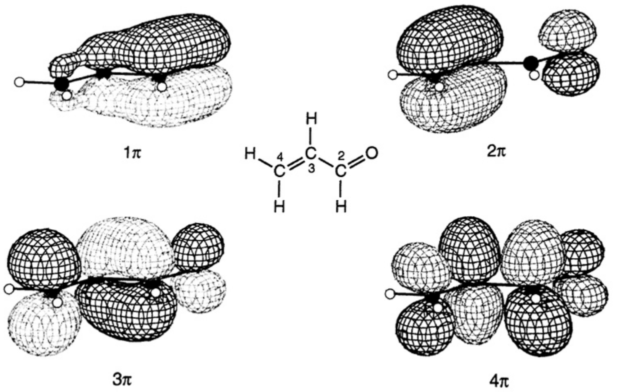
The four remaining electrons in butadiene pair into 1π and 2π. The node between the second and third carbons for the 2π MO is balanced by the antinode from the 1π, so there is still some π bonding. The bond length data supports this, with the $C_2 - C_3$ bond being slightly longer. The bonds are described as **conjugated** – requiring an uninterrupted chain of p orbitals.

Propenal



We can again model the π system as four 2p orbitals in a row. However, they aren't all the same energy, since the oxygen orbital is lower in energy.

- 1π has no nodes between atoms, and constructive interference occurs. However, there is only a small contribution from C_4.
- In the 2π MO, there is constructive interference between C_3 and C_4, but there is little constructive interference between C_2 and O because of small coefficients.
- 3π and 4π have two and three nodes respectively, just like with butadiene.



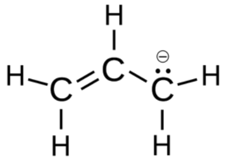
We have discussed that the AO closest in energy to the MO is the largest contributor. Hence, the oxygen makes a greater contribution to the 1π MO, but a smaller contribution to the 4π MO (this can be seen above).

The extra π electrons in propenal once again mean that the bonding is intermediate between a single and double bond. The resonance structure is as follows:

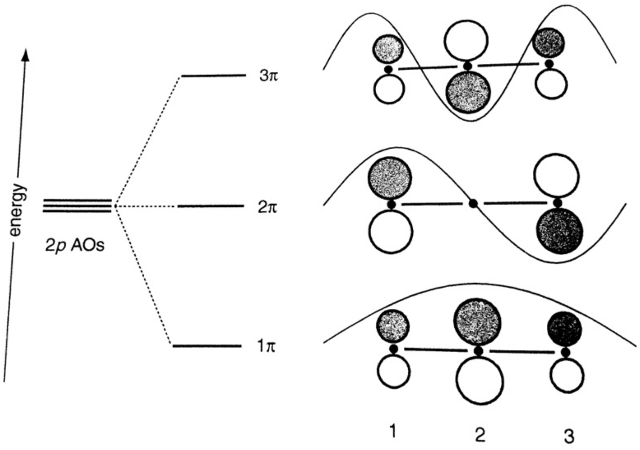
The fact that the C_2–C_3 bond is only slightly shorter than usual suggests that B makes only a minor contribution. Nevertheless, the removal of electron density from C_4 in structure B is necessary to explain nucleophilic attack on C_4 (nucleophiles rarely attack C=C bonds, and in this case would rather directly attack the carbonyl).

Three p orbitals in a row

When three p orbitals overlap, we can construct MOs in the same way as with four MOs. The main difference is that, with three p orbitals, the 2π orbitals will be non-bonding.



One example is the allyl anion, which is basically a propene minus an H^+ (a very strong base is required to form this).

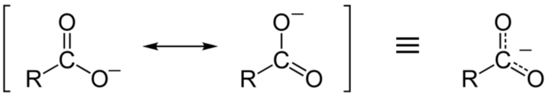


Once again, the coefficients follow a sine relationship.

- Occupation of the 1π orbital contributes to bonding, while occupation of 2π doesn't.
- Since 1π is symmetrical about the centre, we expect the bond lengths to be identical.
- Since there are two electrosn contributing to bonding over three atoms, we expect only partial bonds.
- The electron density is greatest on the end carbons: while for 1π, the central carbon has the most electron density, it turns out that the 2π effects outweigh this.

The localised resonance picture also explains some of these characteristics.

Carboxylate anion



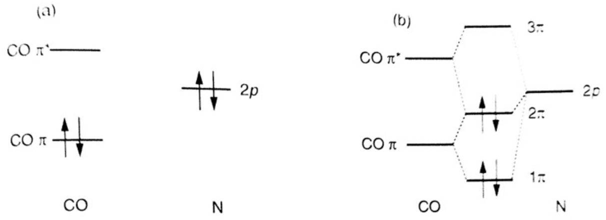
This ion is similar to the allyl, except one of the carbons is replaced by an oxygen. The conclusion is the same, although bear in mind that the oxygen AOs have lower energy

Amides

In an amide linkage, the atoms are in the same plane. As such, we deduce sp^2 hybridisation on the C, N, and O. The π system can be modelled as three p orbitals in a row, but these AOs do not have the same energy.

- The lowest energy MO is bonding across all three atoms.
- 2π evidently contributes little to bonding as carbon makes a negligible contribution.
- 3π is the highest energy orbital, an antibonding MO.

This configuration is what we observe, so it must have the lowest energy. The simple model of imagining that nitrogen has lone pairs which do not interact is wrong, because an interaction would lower the total energy.



The effect of this is that the C-O bond is actually weaker than the normal C=O. Conversely, the CN bond is slightly stronger, and partially restricts rotation.

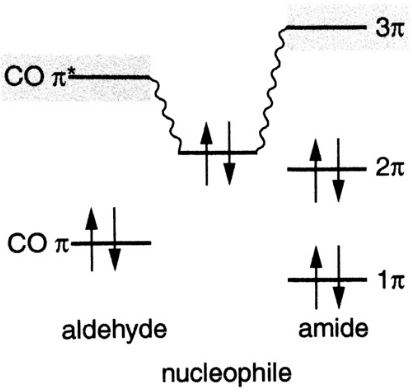
Structure B is a minor contributor, but it is definitely the case that the lone pair interacts.

Chemical consequences

Amides and aldehydes

The attack of a nucleophile on a carbonyl depends heavily on the groups connected to the carbon. For example, amides and aldehydes who very different reactivity towards water aldehydes form tetrahedral hydrates, while amides don’t react.

Any nucleophilic attack on a carbonyl involves the HOMO of the nucleophile (the lone pair on oxygen) interacting with the LUMO of the carbonyl.



- For an aldehyde, the LUMO is the π* from the CO.
- In an amide, the LUMO is the 3π.
- The 3π is higher in energy than the π* (because the 3π is an antibonding MO formed from the π* and something else).
- The lone pair on oxygen is lower in energy than the π* on CO, since the π* is the antibonding MO formed between C and O.
- Hence, it is clear that there will be a better energy match in an aldehyde.

The participation of the nitrogen lone pair in the π system raises the energy of the LUMO, decreasing reactivity – this is the **conjugative effect.**

This explains why amides can work as bases better than alcohols. The energy separation between the HOMO and LUMO is greater for the oxygen in an alcohol, so there is a smaller interaction. However, in an amide, the carbonyl oxygen is more basic than the nitrogen.

We would normally expect the oxygen to have a lower energy, but because of delocalisation, the energy of nitrogen's lone pair decreases as it moves into the 2π MO (see diagram on previous page for lowering of 2π energy). Hence the oxygen lone pairs become the HOMO. In terms of resonance structures, the nitrogen lone pair is involved in the formation of a partial C-N π bond, reducing the availability of the electron pair.

Amides and acyl chlorides

Acyl chlorides react much faster with nucleophiles, while amides do not. There are two factors which the side group on the carbonyl can affect:

1. The **inductive effect**, which is when the side-group draws electron density density and increase the partial positive charge on the carbonyl cation (shortens C=O bond length).
2. The conjugative effect, which decreases the reactivity of the carbonyl by raising the energy of the π* LUMO (lengthens C=O bond).

Nitrogen and chlorine have a similar electronegativity, so a similar inductive effect. However, the orbital involved in delocalisation is 3p for chlorine and 2p for nitrogen. The large size of the chlorine orbital means that it interacts less strongly with the other 2p orbitals, so the conjugative effect is smaller. This means that the LUMO for chlorine has lower energy, so the acyl chloride is more reactive than the amide.

Amides and esters

Hydrolysis of an amide requires more extreme conditions than for an ester. The difference between the two is a nitrogen versus an oxygen (the single bonded one). The oxygen 2p orbital has lower energy, meaning that it is further away from the CO π^* orbital. The resulting 3π MO is hence lower in energy for the oxygen, and since this 3π orbital is responsible for reactivity, the ester is more reactive.

Since oxygen is more electronegative, the inductive effect also means that an ester should be more reactive.

Stabilisation of negative charge by conjugation

While both alcohols and carboxylic acids contain a C-O-H, dissociation in alcohols is negligible. In both cases an O-H bond is being broken, but the difference is that the the carboxyl ion has delocalisation.

This makes it more stable, because the negative charge (i.e lone pair) on the oxygen becomes involved in the carbonyl π system.

An **enolate anion** is formed by treating an aldehyde or ketone with a strong base, removing a proton from a carbon adjacent to the carbonyl. An **allyl anion** needs a very strong base to form. The enolate is much easier to form, because the negative charge can be delocalised onto the electronegative oxygen. There are no electronegative species in the allyl, making it difficult to form.

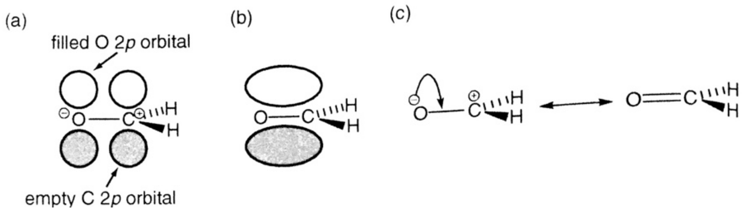
If the negative charge is conjugated with a more extensive π system, the charge can be delocalised more and is more stable. This means that species like phenol are weak acids, despite the aforementioned almost non-existent acidity of alcohols. The benzene ring stabilises the charge.

Substitution and elimination reactions

- Tertiary halogenoalkanes do not under go S_N2 because of **steric hindrance**, the three methyl groups prevent the backside attack required for S_N2.
- S_N1 involves the formation of a tertiary carbocation, which is very reactive and reacts quickly with the nucleophile in the second step to form the product.
- Primary halogenoalkanes do not undergo S_N1 because the carbenium ions that form are very unstable, compared to the $(CH_3)_3C^+$ that is formed by a tertiary halogenoalkanes.

We have already examined the stabilisation of negative charge through delocalisation; it turns out positive charges can be similarly stabilised. The main difference is that positive charges do not move: the movement of electrons is responsible for the appearance of a localised positive charge on a molecule.

CH_3^+ is a planar species, whose LUMO is a vacant 2p orbital going above and below the carbon. This LUMO is highly susceptible to nucleophilic attack, so is too reactive to be stable. However, imagine replacing one of the hydrogens with an oxygen. The orbitals overlap to form a delocalised π system (which is more stable). However, this also means that the positive charge is no longer localised over the carbon.



The stabilising effect can also come from an OH; the lone pairs on the O contribute to a C-O π bond.

The acylium ion

Acyl chlorides can form salts, e.g $[CH_3CO]^+BF_4^-$. It turns out that this cation is linear, with a short C-O bond length. Once a chloride is lost from the acyl chloride, the remaining acylium flattens out. The oxygen then stabilises the charge.

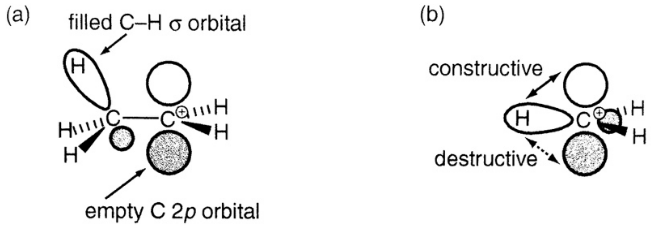
Even when it is not an ion, the C=O bond length in an acyl chloride is shorter than usual. The chlorine is electronegative enough to draw electrons from the carbon, making it slightly positive. This then attracts electron density from the oxygen lone pair, strengthening the bond.

Stabilisation by adjacent π systems

The next highest energy after lone pairs are π bonding orbitals. As has previously been examined, the extra carbon 2p interacts with C-C π bond to form a 1π, with lower energy. The delocalisation on an allyl cation allows the positive charge to be stabilised.

Stabilisation by σ bonds – σ conjugation

In order for the σ bond to help stabilise the cation, it must be in the correct orientation. The interaction between this σ bond and the 2p orbital on the C^+ lowers energy, stabilising the cation. This is not possible in CH_3^+ because there is no net overlap.



The stabilising effect of σ conjugation is relatively small, because the energy of the σ orbital is much lower than the 2p, so there isn't much interaction. However, the more sigma conjugation there is, the more stabilised the cation will be. This explains the stability of the trimethylcarbenium ion, $(CH_3)_3C^+$: extensive σ conjugation spreads electron density over the whole molecule.

Elimination reactions

Adding hydroxide to a halogenoalkanes, under the right conditions, produces an alkene rather than an alcohol. We will assume that the first step of the elimination is the formation of the trimethycarbenium, $(CH_3)_3C^+$. Delocalisation means that electron density is removed from C-H σ MOs, weakening them. However, the C-C bond is strengthened. This means that the H can escape more easily.

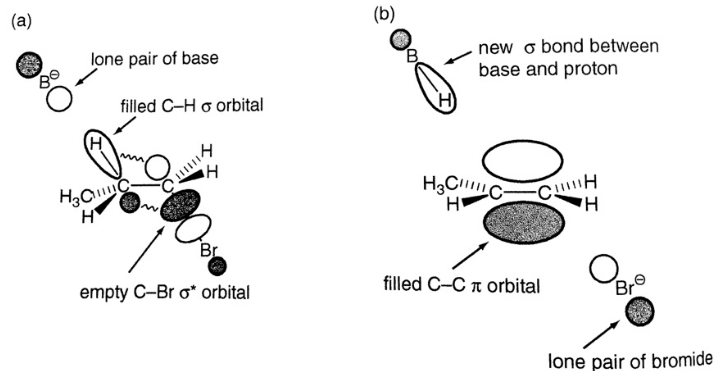
When a strong base is present, rather than attacking the p orbital on the carbon (in a substitution reaction), the hydroxide reacts with the acidic H, in the second step of the E1 mechanism.

Using water will force a substitution reaction, since water is not strong enough a base to remove the acidic hydrogen. It instead attacks the empty carbon orbitals.

However, it is important to note that the action of hydroxide is not fundamentally different for a substitution vs an elimination. In both cases there is a nucleophilic attack, but in the elimination case, this attack is on the H-X bond rather than the C-X bond.

E2 elimination

The previously-used E1 mechanism does not work for primary halogenoalkanes, since there is insufficient σ conjugation for the carbenium ion to be stable. We have previously seen that a base can remove the H^+ from a trimethylcarbenium because the interaction of the C-H σ MOs with a vacant p orbital. While primary halogenoalkanes do not have this vacant p orbital, they do have an empty C-BR σ* that has similar (but weaker) effects.



As some of the electron density from the C-H bond goes into the C-Br σ*, the C-Br bond weakens (because σ* is an antibonding MO). This leads to fission of the C-Br bond, so Br is eliminated.

- A base is needed to remove the hydrogen; the electron density effects are not sufficient.
- The π bond forms when the electrons from the C-H σ bond move to the C-Br σ* orbital.
- All the steps in this mechanism occur at the same time, so in fact one could visualise what is happening with a transition state.

Since the mechanism involves the base and the halogenoalkanes at the same time, the kinetics are second order. In actual fact, the E2 is far more common than the E1. E1 requires the group to leave then a base to remove the H. However, if the base is too strong, an E2 reaction will be favoured.

Addition reactions

Addition reactions are the opposite of elimination reactions. For example, we can add an H^+ to an alkene to form a carbenium ion.

The LUMO in this interaction is clearly the 1s from the H^+, which has no electrons. Since there are no lone pairs in the alkene, the HOMO is the C-C π bonding MO. However, the 1s can interact with this MO from two different directions, affecting the product.

The first is more stable because of the stabilising effect of the alkyl groups. Generally, the proton will attach to the carbon atom that results in the most stable cation.

Once this carbenium ion has been formed, it will react with a nucleophile. For example, when adding HBr to the alkene, the alkene is first protonated before the carbenium is attacked by the bromine to give the halogenoalkanes.

The effects of the solvent

The solvent can sometimes affect the reaction. For example, a tertiary nucleophile first forms the trimethylcarbenium ion, but after this, it can either undergo S_N1 or E1, depending on the solvent.

For tertiary halogenoalkanes, the fastest reaction occurs in a water solvent (because ions are more soluble in this). However, the S_N2 mechanism is much faster when no solvent is present.

Different types of solvent

Solvents might be classified in terms of polarity, but this is difficult to quantify. Instead, we measure the **relative permittivity**. Ideally, the electrostatic force between ions is given by:

$$force = \frac{z_+ z_- e^2}{4\pi\varepsilon_0 r^2}$$

In a solvent, this force is reduced by a factor called the relative permittivity, ε_r.

$$force = \frac{z_+ z_- e^2}{4\pi\varepsilon_0 \varepsilon_r r^2}$$

This means that a larger permittivity decreases the force between ions. Solvents with higher relative permittivity (higher than 15) are described as **polar**. Substances with a high permittivity will cause ions to be split up, hence the origin of 'like dissolves like'. Water has a permittivity of about 80.

Hydrogen bonds

When hydrogen is bonded to an electronegative species, that species withdraws electron density. Any other molecule with a high energy lone pair can interact with this acidic hydrogen, forming a partial bond.

- The hydrogen bond is much weaker than full bonds
- The strength increases with electronegativity, i.e HF > H_2O > NH_3 > H_2S
- The hydrogen bond to an ion is much stronger than to a neutral molecule, e.g the H bond between HF and F^- is significantly stronger than that between two HF molecules.

Solvents with weakly acidic hydrogens capable of forming hydrogen bonds are called **protic** solvents. Protic solvents are always polar. However, not all polar solvents are protic, these are called **polar aprotic** solvents.

The solvation of ions

Energy is always given out when ions are transferred from a vacuum to solvent. We can quantify this by finding the difference in energy required to charge a particle in vacuum and in solution (equivalent to the work done when transferring from vacuum to solution).

$$w = \int_0^Q \frac{Q}{4\pi\varepsilon_0\varepsilon_r r}dQ - \int_0^Q \frac{Q}{4\pi\varepsilon_0 r}dQ$$

Where $Q = ze$. We also know that ΔG is equal to the maximum non-expansion work done (this can be shown with thermodynamics). Evaluating the above gives the Born equation:

$$\Delta G^{\theta}_{solv} = -\frac{z^2e^2N_A}{8\pi\varepsilon_0 r}\left(1-\frac{1}{\varepsilon_r}\right)$$

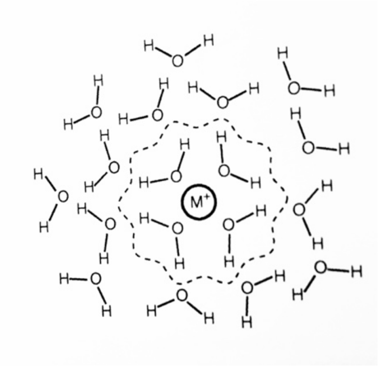
The greatest decrease in Gibbs energy is when the bracket term is close to 1, i.e ε_r is large and the substance is polar. The main difference between polar protic and polar aprotic solvents is their interaction with anions. Protic solvents form hydrogen bonds, while aprotic polar solvents can only solvate ions using dipole-dipole interactions.

Solvating different ions

How well an ion is solvated also depends on the ion itself. From above, the Gibbs energy of solvation becomes more negative as the ion gets smaller and the charge increases. It turns out both entropy and enthalpy change of solvation are negatively proportional to z^2/r .

As ions coordinate with the water, one would expect entropy to decrease as the system is more ordered. However, for +1 cations, this decrease in entropy is less significant than the disruption of the water network, so entropy increases as ions go from gas into the solution.

When an ion is introduced to water, it coordinates directly with a number of water molecules, the **primary hydration shell**. However, this may in turn affect more molecules to form a **secondary hydration shell**. The total number of water molecules associated with various ions is called the **hydration number.** This depends on charge density.



Solubilities of salts in water

To determine the solubilities of salts, we need to consider the Gibbs energy change when the salt dissolves. If $\Delta G_{solution}$ is negative, the salt is soluble. The equilibrium reaction in question is

$$MX(s) \rightleftharpoons M^+(aq) + X^-(aq)$$

Bearing in mind that the equilibrium expression does not include the solids, we can examine the equilibrium constant for the above expression, called the **solubility product** K_{sp}.

$$K_{sp} = [M^+][X^-]$$

$$\Delta G^\theta_{sol} = -RT \ln K_{sp}$$

We can construct an energy cycle:

It is often helpful to break down ΔG into ΔH and ΔS. Conversely, instead of working out $\Delta G_{solution}$, we can work out $\Delta G_{precipitation}$.

Acid strengths and the role of the solvent

Acids dissociate in water in an equilibrium reaction; the equilibirium constant is called K_a:

$$HA(aq) \rightarrow H^+(aq) + A^-(aq)$$

In all cases, we can determine the strength of the acid by considering ΔH and ΔS for the above process. In turn, these values are best determined using energy cycles:

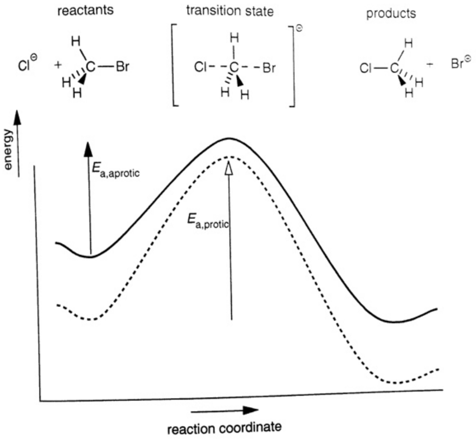
For example, HCl has a favourable ΔH but a negative ΔS, though the ΔH term dominates. However, HF has a higher lattice enthalpy and also a larger entropy decrease during solvation, both because fluoride is very electronegative.

Ethanoic acid is a weak acid, but as we replace the methyl hydrogens with chlorine, the strength increases dramatically. These effects are centred on the entropy changes. The more chlorine ions, the less negative the entropy change. The electron-withdrawing chlorine atoms mean that charge is spread better across the molecule; the charge is less concentrated so solvent water molecules are less tightly held, so there is a smaller increase in order.

Effect of solvents on the reaction rate

When *t*-butyl chloride reacts with a nucleophile, it can either undergo S_N1 or E1. In both cases, the first step is the formation of the trimethylcarbenium and chloride ions. Forming both transition states and intermediates will require less energy in a polar medium, since the transition states and intermediates are ionic.

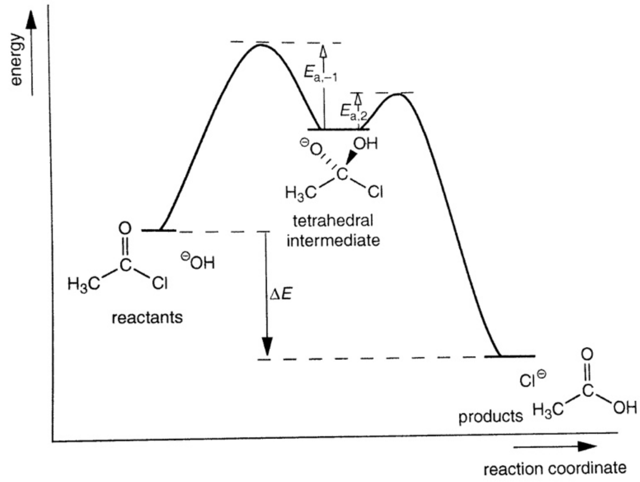
However, S_N2 is faster in non-polar solvents (and fastest with no solvent). Consider the reaction between CH_3Br and Cl^-. In a protic solvent, this anion is strongly solvated by H bonding, whereas in an aprotic solvent like acetone, it is less strongly solvated. S_N2 requires the correct geometry, but the formation of hydrogen bonds inhibits the anions from attacking the back of the halogenoalkanes.



Having a protic solution lowers the energy of both the reactants and the transition state, but there is a larger decrease in the energy of the reactants, so overall the E_A is higher.

Leaving groups

Some reactions happen (e.g OH attacks ethanoyl chloride to form ethanoic acid), while similar processes do not (e.g chloride attacks ethanoic acid to form ethanoyl chloride). This can be explained in terms of the energy profiles of the reactions.



When the intermediate is formed, it can either lose an OH or a Cl, (the C-O and C-C bonds are too strong). However, it is easier to lose the Cl^-, i.e Cl^- is a better **leaving group**.

This explains why chloride does not attack ethanoic acid: the tetrahedral intermediate formed would just regenerate the ethanoic acid.

Leaving group ability

The leaving group ability depends on the bond strength between the leaving group and the atom with which it was bonded, and the ‘stability’ of the leaving group once it has been formed.

- N_2 (e.g in R-N_2^+) is a very good leaving group, because the N_2 is very stable. Also, the C-N bond is weak because N_2 draws electron density.
- In the halides, we note that the strength of leaving group ability increases down the group. This is because the bond strengths decrease down the group (increasing leaving ability), while stability only decreases minimally.
- Species with a higher effective nuclear charge are more stable, and are hence better leaving groups. This explains why $F^- > OH^- > NH_2^- > CH_3^-$.
- Species which can delocalise charge better are more stable so are better leaving groups.

The strength of a leaving group can be measured by the strength of the conjugate acid, i.e we look at the dissociation ability of H-(leaving group). While this works most of the time, there are some discrepancies. For example, while H_2O is a better leaving group than Cl^-, HCl is a stronger acid than H_3O^+. Such discrepancies are explained by the fact that leaving group ability actually depends on bond strength to carbon, not hydrogen.

Competing reactions

Reactions of carbonyls with hydroxide

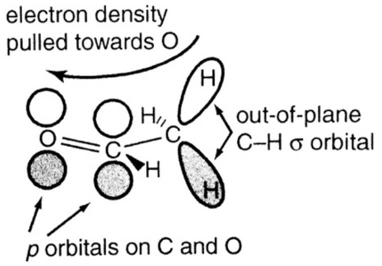
When OH^- reacts with aldehyde, two reactions can happen. Both these reactions do happen, resulting in a mixture.

1. The OH^- attacks the C=O π^* MO to form a tetrahedral intermediate (then reforms an aldehyde)

Because of the electronegative oxygen, the carbon is slightly positive charged (so is more readily attacked). After the intermediate is formed, the strongest leaving group is the OH^-.

2. The hydroxide deprotonates the aldehyde to form an enolate.

The C-H bonds on the carbon adjacent to the carbonyl are weakened because of interaction between their orbitals and the π system. There is some weak σ conjugation between the C-H σ bonds on the methyl group and the π system. This weakens the C-H bond, resulting in a more acidic H. However, this process is still not very rapid.



Carbonyl compounds have pK_a values of around 20, compared to about 50 in alkanes.

Isotopic labelling

We know that both of the previous reactions happen through experimental evidence using isotopically labelled reagents. This method involves introducing another isotope (^{18}O or deuterium) then following it.

Labelled carbonyl compounds can be made like so:

Reactions of enolates

The enolate ion has significant negative charge on both the oxygen and the methyl carbon. This means that the anion acts as a nucleophile, either being attacked at the carbon or oxygen – the enolate is an **ambident** nucleophile. The reactivity is not the same at these positions because the charge is not equally distributed.

We can add an **alkylating agent** to the enolate – this is a substance that undergoes nucleophilic substitution, adding an alkyl group to the nucleophile. For example we can add iodomethane to the enolate. There are two possible reactions:

Experimentally, only the second occurs. The LUMO is the the C-I σ* orbital. The HOMO on the enolate is the 2π, which has a greater contribution from the carbon than the oxygen. This means that the reaction takes place through the end carbon rather than the oxygen.

Changing the alkylating agent

It is possible to ‘encourage’ the oxygen to become alkylated, by choosing an alkylating agent that has a greater emphasis on charge rather than orbital energy (since the oxygen is more negative). For example, if we replace CH_3I by CH_3Cl, the C-Cl bond is far more polar, so electrostatic interactions will be more important.

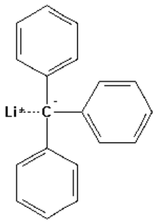
Changing the solvent

If we use a polar protic solvent, the carob product is highly favoured because the solvent forms hydrogen bonds to the oxygen, hindering approach.

Unsymmetrical enolates

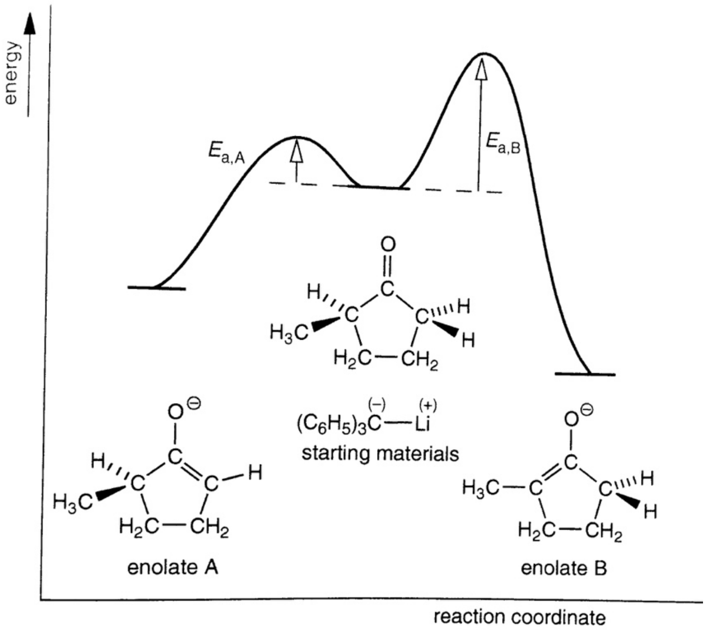
There is only one enolate that can be formed from ethanal, but for asymmetrical ketones two different enolates can be formed depending on which hydrogen is removed.

We need a strong base to take off the hydrogen, but we do not want it to attack as a nucleophile. Hence we choose a bulky base, e.g triphenylmethyllithium.



Using such a base means that it is easier to remove a hydrogen from a carbon which does not have a methyl group (less steric hindrance).

Because of steric hindrance, it is easier to form enolate A. However, enolate B is lower in energy because of σ conjugation with the π system.



When an excess of base is added to the ketone at room temperature, much more of enolate A is formed. This is because products cannot be regenerated (since there is nothing acidic from which the enolates can reform ketones), so the reaction with lower activation energy is favoured.

However, if we have a slight excess of ketone, enolate B is more dominant. This is because the enolate can react with a molecule of the ketone to form the other type of enolate.

Enolate (A or B) + ketone ⇌ ketone + enolate (B or A)

An equilibrium exists, so the lower energy product (enolate B) is favoured.

Substitution vs elimination

When a nucleophile/base reacts with a halogenoalkanes, we can either form substitution or elimination products. Reacting 2-bromohexane and methoxide gives one substitution product and three elimination products (which are just isomers of each other).

If we increase the partial positive charge on the carbon, nucleophiles will be more attracted to it, favouring substitution. This is what happens if we replace the bromine atom with something more electronegative. However, whilst the amount of the substitution product increases, the rate decreases significantly because these more electronegative species are poorer leaving groups.

We can further increase the substitution product by using a solvent that is a weaker base, so elimination (which involves deprotonation) is unfavourable.

Energy profiles

If there are several possible pathways, the one with the lowest activation energy will have the largest rate constant. When the temperature is lowered, the rate constants for all pathways will decrease, but the decrease will be greatest for pathways with the highest activation energies.

When the temperature is high enough for all reactions to be readily reversible, the lower energy products will be favoured.

Conclusion: why do chemical reactions happen?

- Why does the chemical reaction increase the entropy of the universe?
    - Consider both the entropy change and the enthalpy change
- Why is a reaction exothermic?
    - The products will have to be, in some way, lower in energy
- Why is one arrangement of atoms lower in energy than the other?
    - Bonding interactions (both ionic and covalent) can result in a lowering of energy
- How can we make the product we want?
    - We need to consider whether examines are favourable (in terms of kinetics, energetics and equilibrium), and consider any competing reactions.