

Self-Assembly of Organic Molecules

Last week we discussed basic forms of self-assembly such as crystal growth

But these cannot create the complexity required for full blown nanotechnology

However, nature provides solution: Organic Chemistry + Molecular Biology = "life"

(Subjects of this and the lecture to follow)

Classes on organic chemistry spend endless time on memorizing lists of chemical names or on hundreds of alternate synthesis procedures

I didn't want to be organic chemist myself,

but I did have a long and successful collaboration with organic chemists

Experience that led me to an understanding of how to exploit organic chemistry via collaboration - an understanding I now want to share with you

I also have two other goals:

1) To bring to chemistry what WE have now learned about quantum mechanics

Chemists allude to the dark mysteries of s & p orbitals, pi bonds . . .

But they're really just standing-waves, like those in springs and water

Which, with your new experience, are now readily explained

2) To show how simple organic chemistry explains things affecting your daily life

Perhaps the chemists got around to this in CHE 300 or Org CHE 200

But they needn't have waited so long!!

With very simple chemistry / organic chemistry I can explain:

free radicals, anti-oxidants, trans-fats . . . and quite a bit more

So, to begin: Demystifying atomic orbitals

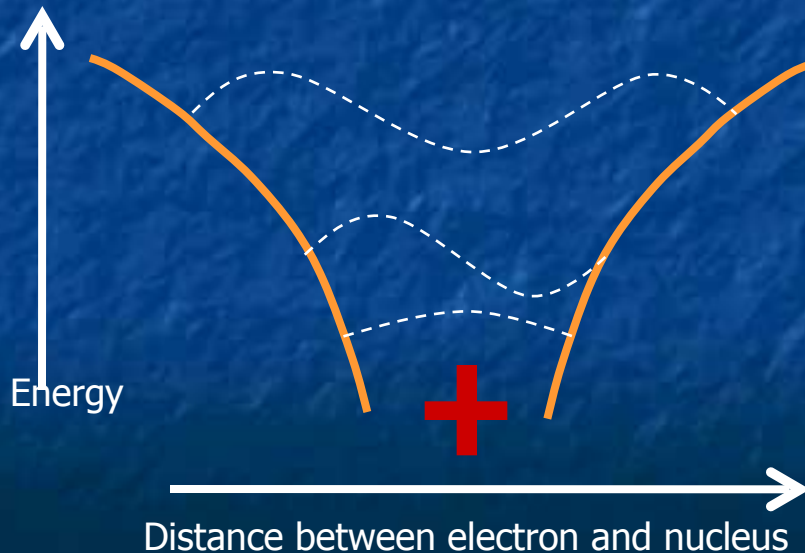
Waves trapped in boxes => Standing waves = Atoms

It's that simple and it's what most of Quantum Mechanics is about

Only becomes hairy if want PRECISE solutions => spherical coordinates . . .

But approximate solutions (treating atoms ~ cubical boxes) are almost intuitive

Origin of atomic "boxes" = attraction between electrons and nuclear protons



$$\text{Potential Energy} = - q_{\text{electron}} q_{\text{nucleus}} / R$$

Box in sense that electron waves are contained

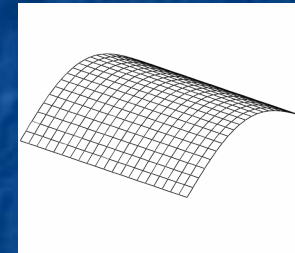
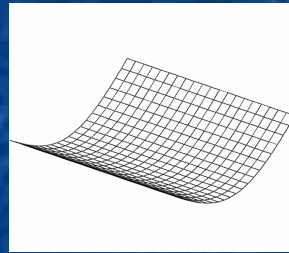
But box that gets bigger for outer electrons

So let's figure out lowest energy electron standing wave:

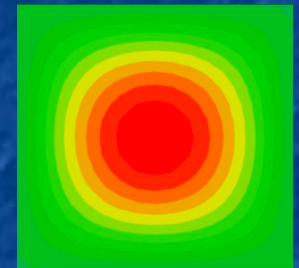
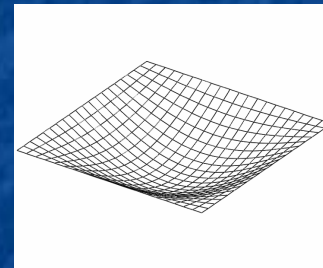
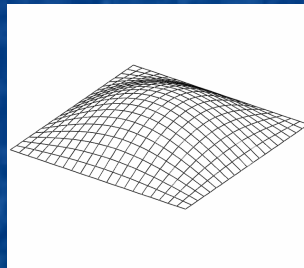
Plane wave bouncing of opposite walls of a square box => standing wave



$\lambda/2$ standing wave in 1D box:



$\lambda/2$ standing wave in 2D box (i.e. $\lambda/2$ X wave + $\lambda/2$ Y wave = "Superposition"):



In 3D box: Add similar $\lambda/2$ wave in Z direction => **spherical cloud**:

What your Chemistry teacher called an "s orbital:"

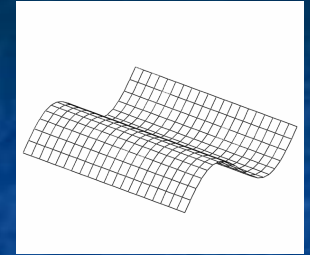
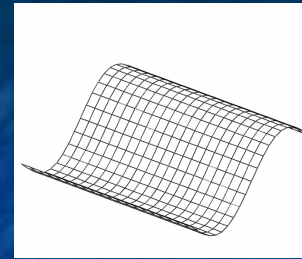


For more energetic electron standing waves: Add oscillations!

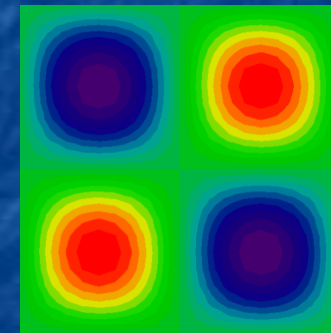
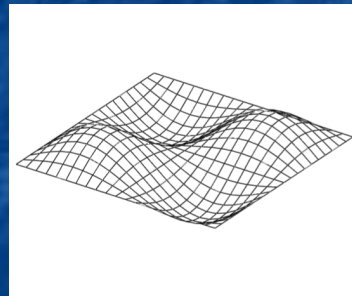
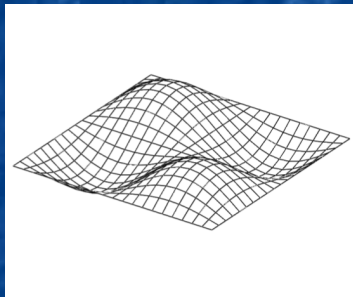
More oscillations:



In 1D =

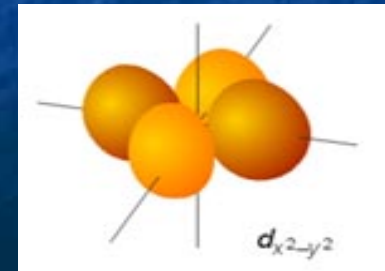


λ standing wave in 2D box (i.e. λ X wave + λ Y wave):



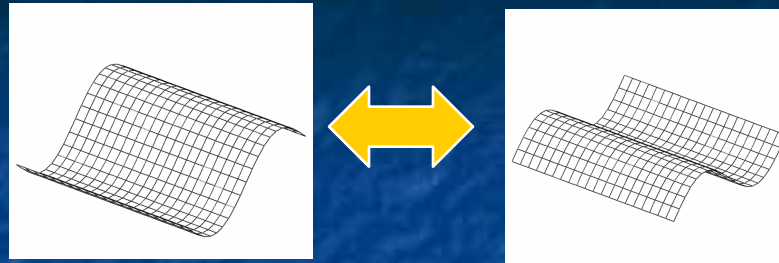
In 3D box: If keep 2D wave (above) but use only $\lambda/2$ wave in Z direction =>

What your Chemistry teacher called an "d orbital:"

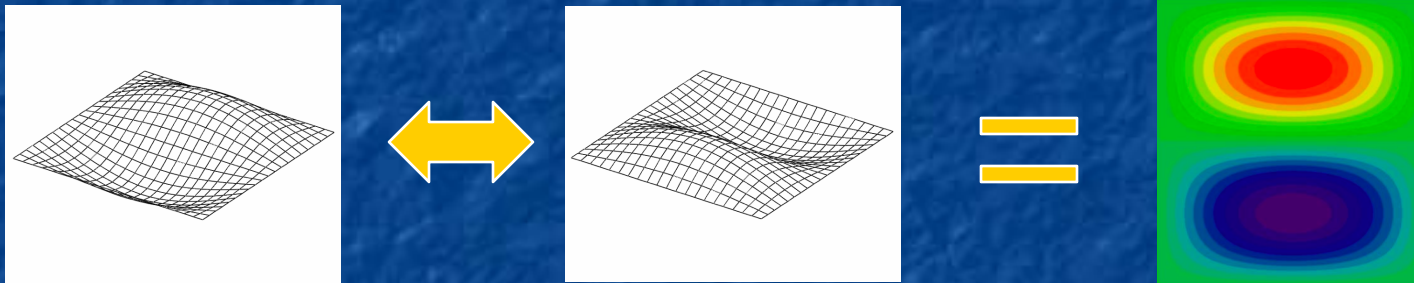


But we didn't **HAVE** to add oscillations in **BOTH** x and y directions!

In 1D: Full λ wave



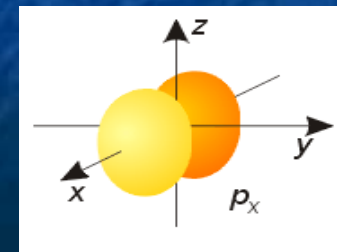
In 2D box: ONLY add $\lambda/2$ oscillation in perpendicular direction!



Why? Because oscillations \Rightarrow Higher Frequency \Rightarrow Higher Energy

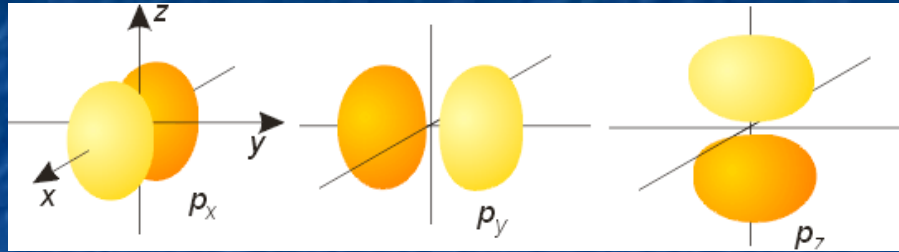
Adding only $\frac{1}{2}$ oscillation in other direction = smallest step up in energy!

Yielding your Chemistry teacher's "p orbital:"



But what about the other possible directions?

That is, why wouldn't the full set of second energy level waves be:



Google images / Winona.edu

Each DOES involve simple standing waves

Only difference is direction of higher frequency wave: X vs. Y vs. Z

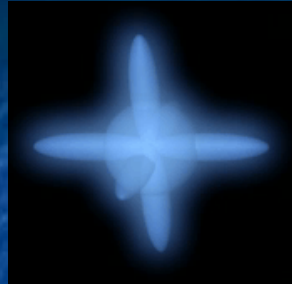
Each arrangement DOES keep charge away from earlier s orbital

Each arrangement DOES keep charge away from other p orbitals

So these are indeed three viable alternatives

Derived solely from standing waves / charge repulsion!!

Yielding lowest energy electron waves (s p_x p_y p_z):



(Now represented as properly fuzzy clouds)

But what is difference in energy between this set of four trapped electron waves?

Can get the answer directly from our knowledge of waves:

Increased frequency (decreased wavelength) => Increased energy

Three p waves have identical wavelength => identical energy (no surprise)

Earlier "s" cloud only had 1/2 the oscillation, but was down in narrower part of box

So its wavelength (distance for full cycle of oscillation) is not much different

Hence, its frequency was similar => Not **that** much lower in energy

So we've accounted for first 4 electrons? No: 8

The above "orthogonal" waves helped electron waves avoid one another

But electrons have a property beyond their charge, and this is called "spin"

Name from Maxwell's Equation that says spinning blob of charge = electro-magnet

Pairs of magnets ATTRACT: Positioning their poles N-S to S-N

Attraction counters charge repulsion => Electron pairs share same wave pattern

What happens in atoms with MORE than above eight electrons?

Possibility 1) Add more complex "d" waves (standing waves with more oscillations)

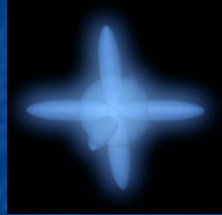
Possibility 2) Use stretched out versions of s and p waves that exploit enlarging box

Nature ultimately uses both possibilities

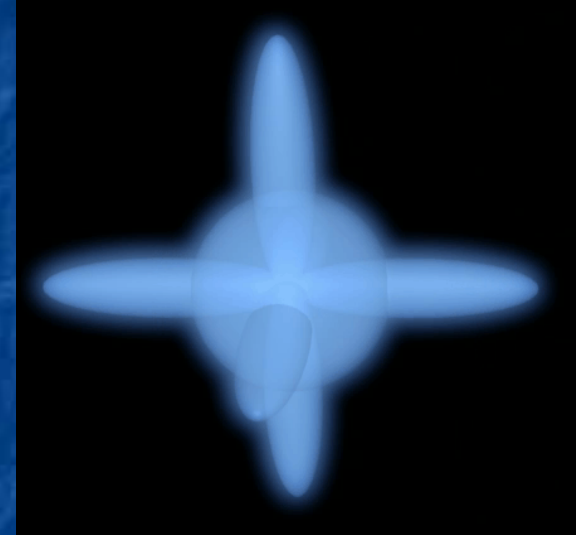
But for organic chemistry (~ first few rows of periodic table) => Possibility 2

Nature just adds enlarged versions of s and p waves:

One set of orbitals in narrow part of box:



Next set of orbitals in wider part of box:



New expanded set of 8 electron waves largely avoid earlier (lower energy versions)

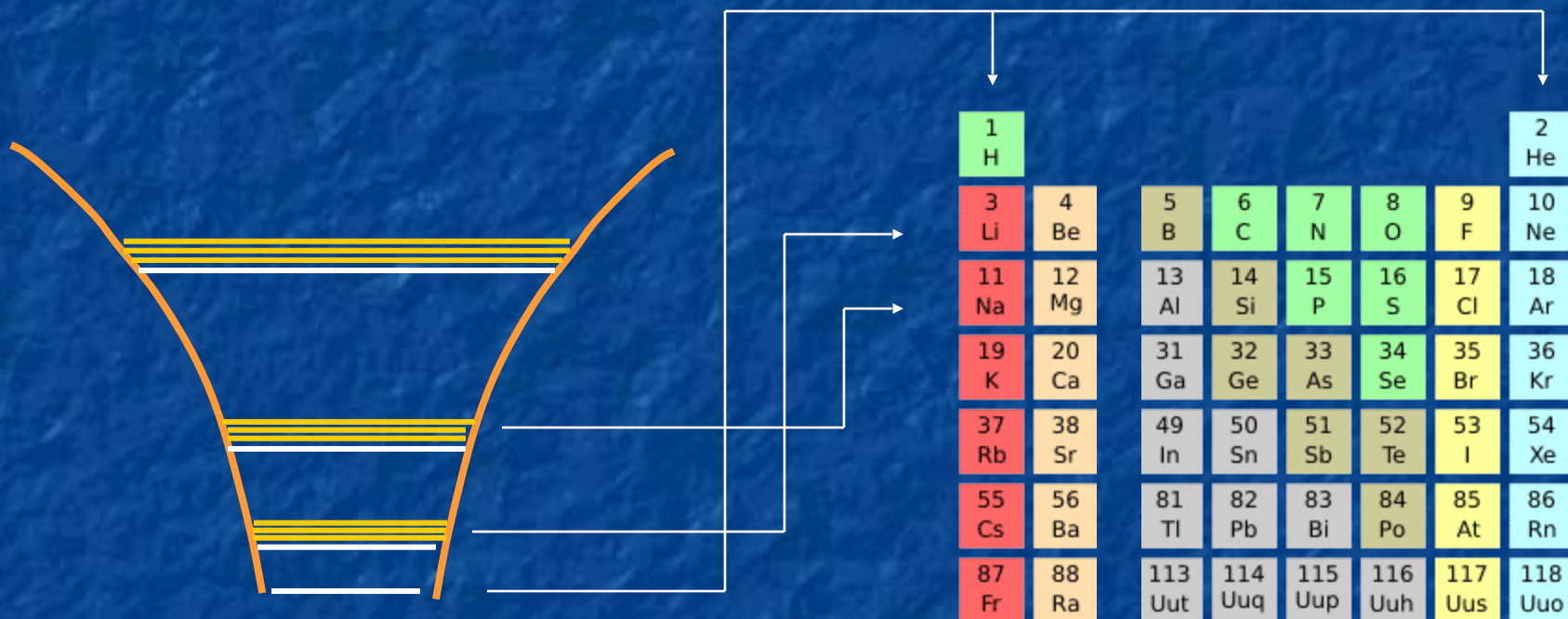
But because outer electron structure is similar, new set acts similarly =>

PERIODIC TABLE with repeated rows of 8 atoms. **But what about first row?**

Tiny p orbitals are too close to other electron clouds => too energetic => deleted

Leading to my "standing wave" version of the Periodic Table

Modified "Periodic Table" that leaves out atoms in middle part:



After the first (lone) energy level (→ H and He) we get

tight groups of four energy levels (capable of holding 8 electrons)

Now, before continuing, we must answer a BIG QUESTION:

Why does "organic" chemistry = carbon chemistry

Are we just being parochial? Failing to see other "paths not taken" (on earth)?

Reasons why Carbon might indeed be special:

1) Carbon has four bonding (valence) electrons = perfect number of bonds for 3D assembly

Bonds = electrons → which naturally repel on another

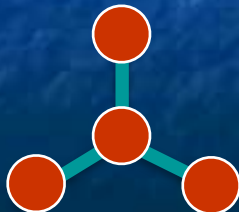
1 bond → line (1D):



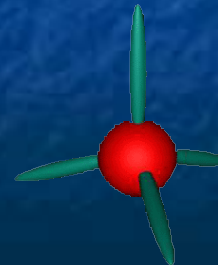
2 bonds → chain (1D):



3 bonds w/ max separation → Plane (2D):



4 bonds w/ max separation → Tetrahedron (3D):



But ALL Column IV atoms have four bonding electrons!

So, as suggested by Sci-Fi, why couldn't "organic" life be based on Silicon?

Covalent Bonds = Pairs of electrons, coupled by magnetic "spin"

Strength of magnetic pairing increases with distance over which the pair overlaps

Carbon is at the TOP of periodic table's Column IV → Smallest nucleus



Carbon Nucleus

Nucleus + electrons

C - C pair

Small core → Snuggle closely → Long overlap → SHORT (1.54 Å) STRONG (3.6 eV) bond

As compared to silicon:

Much larger core (14/6 times more protons, neutrons, electrons):



Silicon Nucleus

Nucleus + electrons

Si - Si pair

Much bigger separation: 2.35 Å (versus carbon's 1.54 Å)

And WEAKER bond: 2.3 eV (versus carbon's 3.6 eV)

And compare non-self bonding energies:

Bond energies in units of eV (1 kcal/mole = 0.04336 eV):

	H	C	N	O	F	Cl	Br	I	Si
H	4.5	4.3	4	4.8	5.9	4.5	3.8	3.1	3.3
C	4.3	3.6	3.2	3.7	5.0	3.5	2.9	2.3	3.1
N	4	3.2	1.7	2.3	2.8	2.0			
O	4.8	3.7	2.3	2.0	2.0	2.3	2.1	2.4	4.7
F	5.9	5.0	2.8	2.0	1.6				5.9
Cl	4.5	3.5	2.0	2.3		2.5			3.9
Br	3.8	2.9		2.1			2.0		3.2
I	3.1	2.3		2.4				1.6	2.4
Si	3.3	3.1		4.7	5.9	3.9	3.2	2.4	2.3

Row 3) C-C bonds are stronger than most C-X bonds (big exceptions are bonds to H & F)

Row 10) Si-Si bonds are MUCH weaker than ALL Si-X bonds in table!!

Silicon very much WANTS to bond with something OTHER than itself

Carbon PREFERS to bond with itself => C chains, rings . . . = "Organic" chemistry

So now let's explore carbon bonding more deeply:

In Periodic Table's 1st row, all atoms **could** support two electron standing waves

Beyond first row, all atoms **could** support full set of 8 electron standing waves

But they won't add electrons above number of charge-balancing nuclear protons

In order to add **BOTH** electrons and protons => They must bond to another atom

As implemented in first row of periodic table by hydrogen:



Or, using the surprisingly powerful **Lewis Representation**:

Core of atom (nucleus + inner non-bonding electrons) = Atom's letter symbol

Bonding electrons = dots (max of 2 for H/He, max of 8 for other atoms)

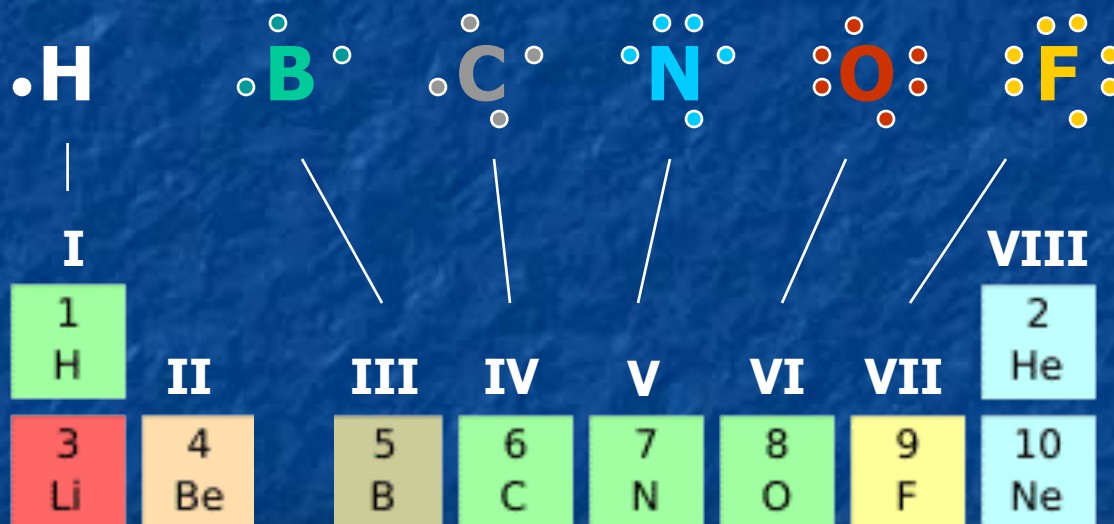
Reaction of hydrogen bonding then becomes



Extending Lewis representation to other atoms:

- Number of bonding (valence) electrons = Column number = number of dots
- Atoms in right section have max of eight closely spaced energy levels
- Electrons like to pair

Eight possible electrons + pairing → 4 likely bond directions

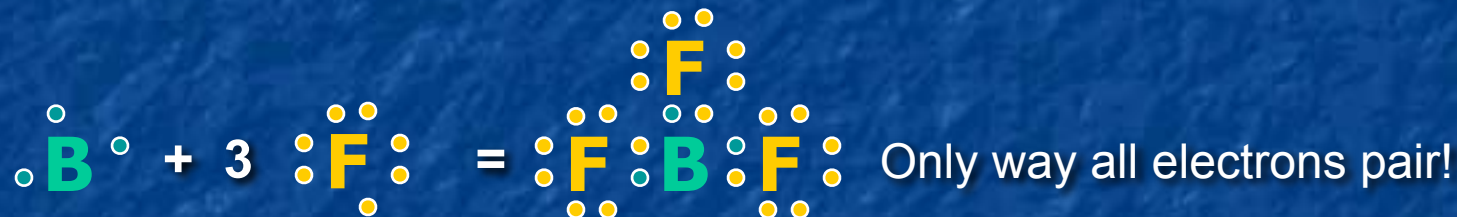


Note semi-standard color code used for atoms above

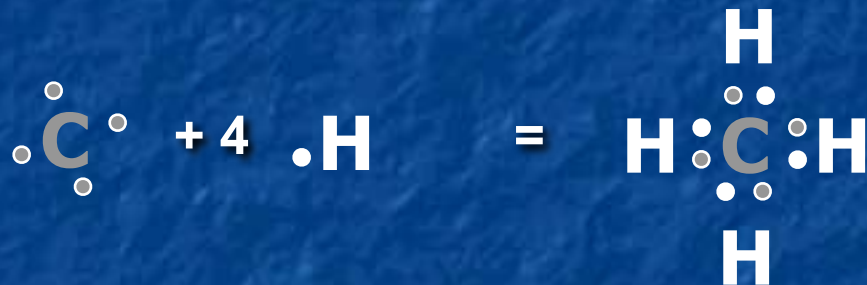
Start to do some chemistry with this "Lewis" notation:

For instance, how would Boron and Fluorine form a compound?

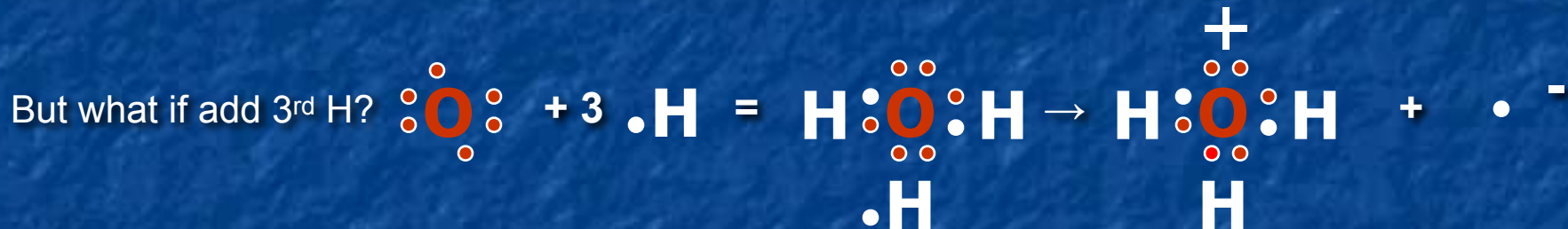
After first spreading Boron's electrons out in different directions



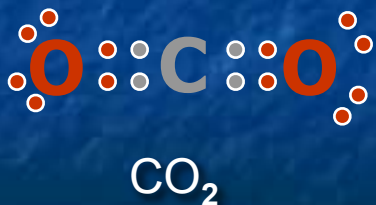
What about carbon bonding with hydrogen?



Again, just spread an atom's electrons out, then pair them into bonds



Or begin to mix things up a bit:

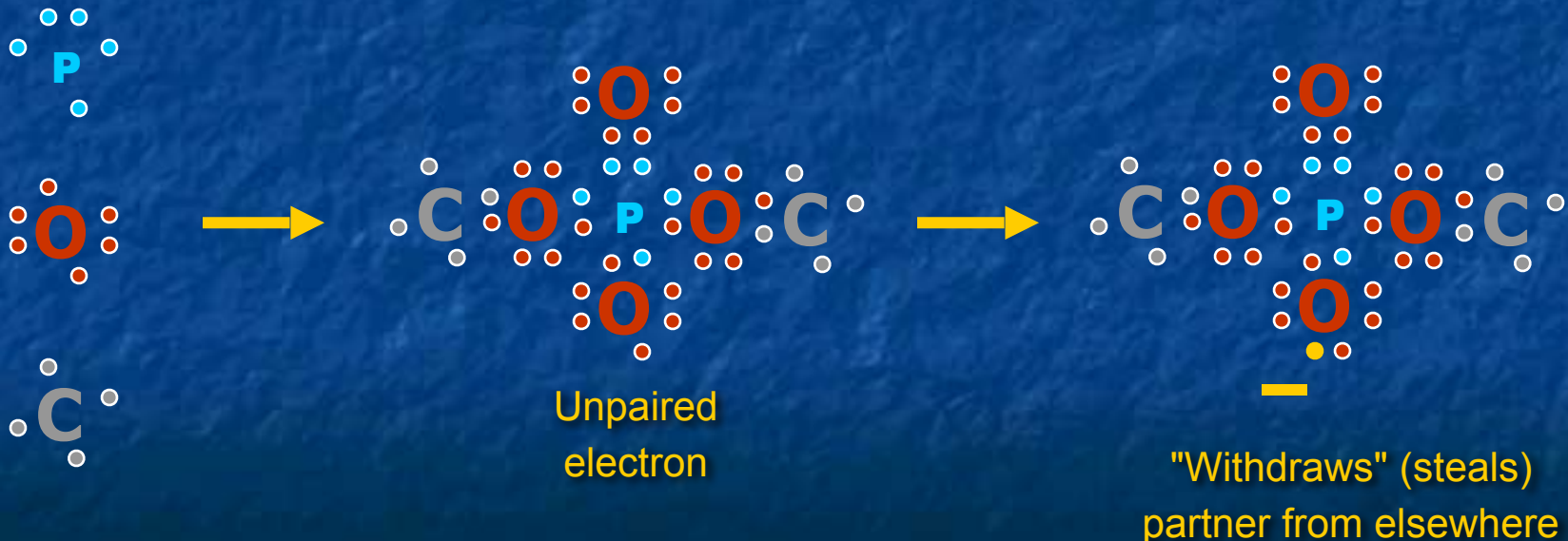


OK, but teach me something I don't sort of already know!

- DNA is incredibly floppy, and yet it maintains a taught double helix structure
- DNA bases are hydrogen bonded, and yet DNA splits apart at just 90C
- DNA fingerprinting is based upon the movement of CHARGES in electric fields

All explained by the Lewis structure of DNA's phosphate groups

and the negative charge that drives them apart:

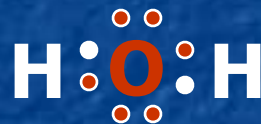


Onward: We also need to figure out the SHAPE of molecules

For that, key fact is:

Pairs of electrons, even if unbonded, act in same way as bonded pairs

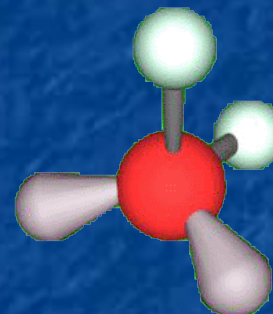
Prime example is in H₂O:



Is this, in fact, a linear molecule?

NO, there are 4 pairs of electrons: Two bonded pairs, two unbonded pairs

Four independent repelling groups → tetrahedron



*Pedantic chemistry teacher: "It's not an **exact** tetrahedron!"*

John: "Show me where 1-2 degrees matters and I'll care!"

Gray lumps =
unbonded pairs

Leading to bonding structure of more complex molecules:

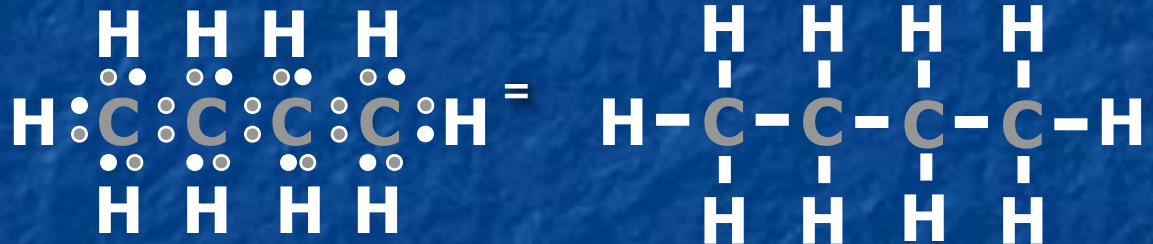
Carbon: Comes with 4 bonding electrons - But has 8 easy places to put electrons

Simplest self-assembly = singly bonded chain ("alkane"):



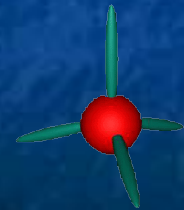
But that leaves LOADS of unbonded, unpaired electrons:

Solution = Add hydrogens

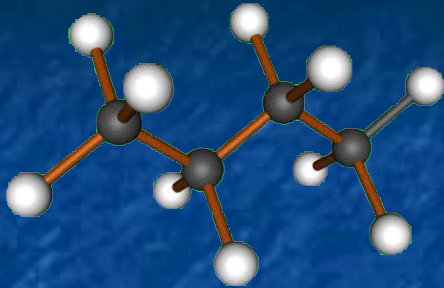


But is this really accurate? Are bonds all at right angles, in a single plane?

Of course not! We have four naturally repelling electron pairs →



So real singly bonded carbon chain must look more like:



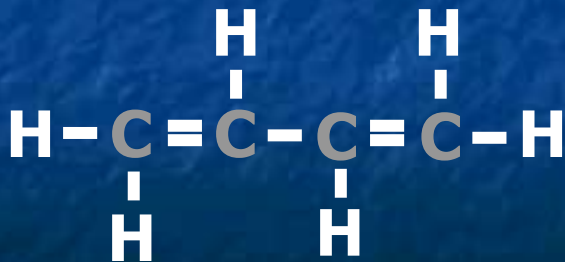
4 independent repelling bonds → Tetrahedra

Figure out number of independent bonds → Spread them out as far as possible

Physicists and chemists go on and on about sp^2 , sp^3 . . . hybridization

But that is just the mathematical process of doing the same thing!!

Example: What if carbon chain had alternating single and double bonds:

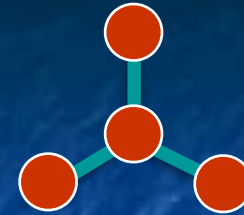


Added enough H's to give each C necessary 4 bonds

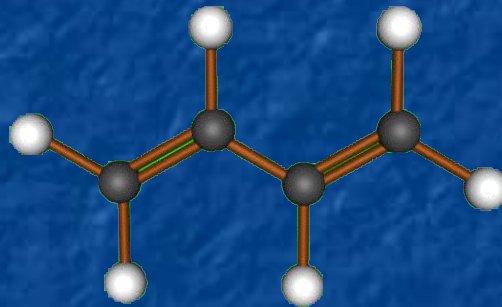
=> only 3 independent bonds: 2 single + 1 double

Three independent bonds repelling one another:

Planar, 120 degrees apart



So true structure of single-doubly bonded chain must be planar:



Then what about a single-triple bonded chain: $\text{H-C}\equiv\text{C-C}\equiv\text{C-H}$

Well, every C now has two independent bonds: 1 single + 1 triple

Will spread in opposite directions = a line: **Exactly as in formula: $\text{H-C}\equiv\text{C-C}\equiv\text{C-H}$**

Everyday Application: "Free Radicals" & "Antioxidants"

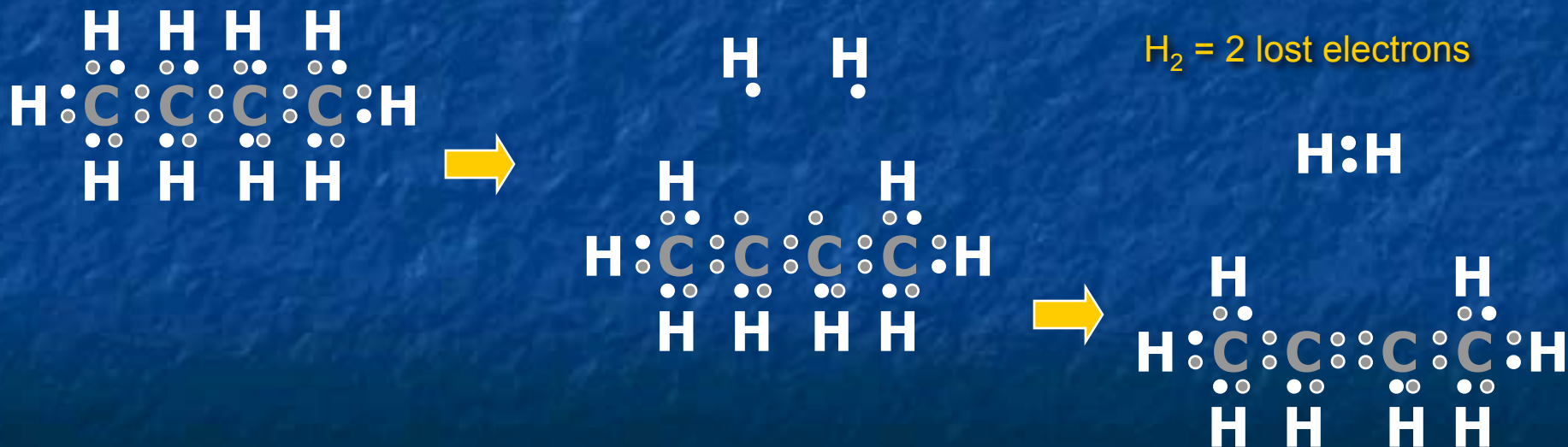
These terms are all over the news (or at least in "health food" advertising)!

Radicals also come up in discussion of possible nanoscience hazards

What's it all about?

Oxygen tends to steal electrons => "Oxidation" is defined as loss of electrons

For organics, this can be the same as losing Hydrogens:



Representing bonds as lines, restructuring of the molecule is more obvious:



But what might make Hydrogens suddenly leave? FREE RADICALS

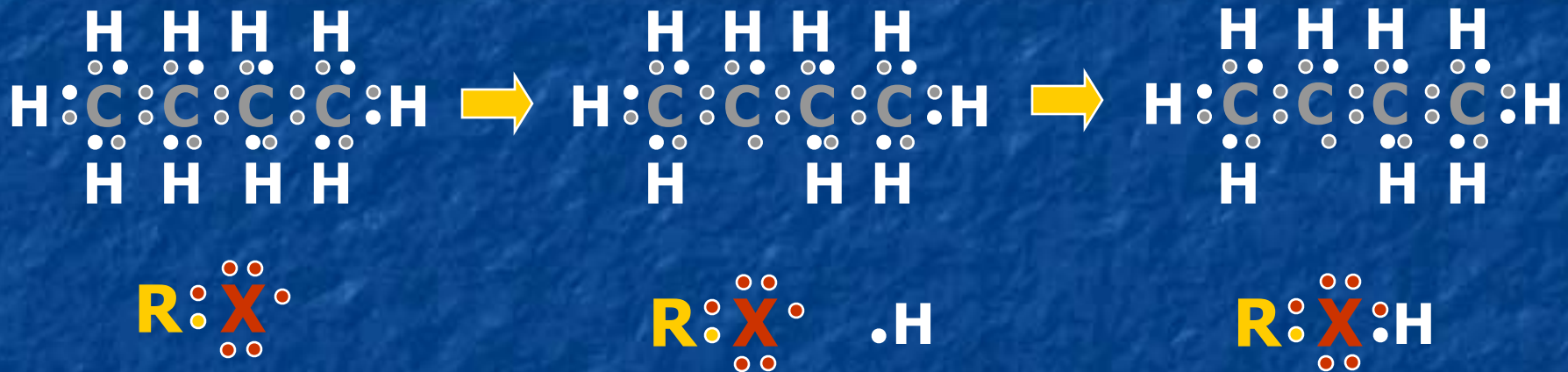
Molecule ("R") with atom ("X") having unpaired electron:

Possible source of radicals? Surface of Nanotubes & C60 (more on this later)

How would these "Free Radicals" promote loss of Hydrogen (a.k.a. oxidation)?

Free Radical Induced Oxidation

Free Radical approaching organic molecule:

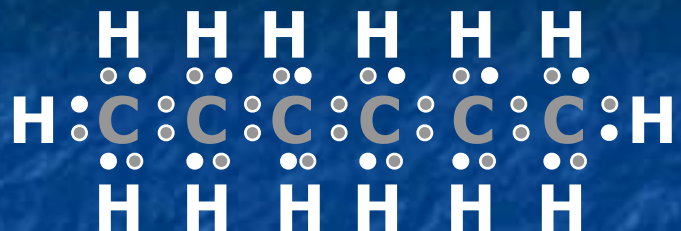


If Hydrogen bond to radical is stronger, it can “steal” a hydrogen!

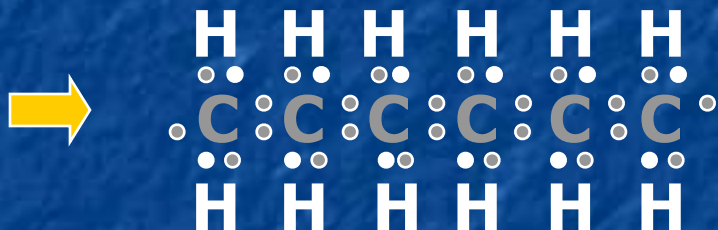
As organic loses electrons it will be driven to reconfigure its bonds

Oxidation => Major restructuring of molecules = Change in their biological function

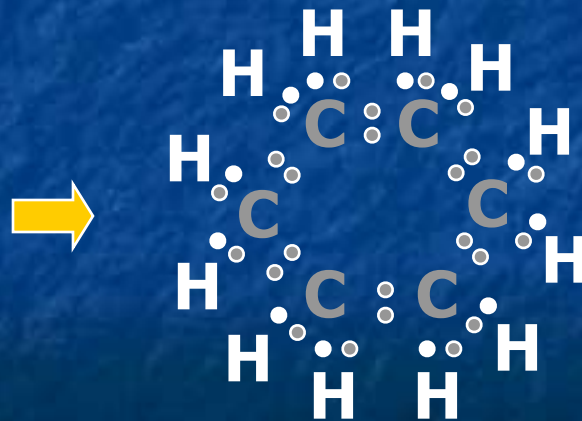
Oxidation could, for instance, convert chains to rings:



Oxidizing two hydrogens on ends of chain



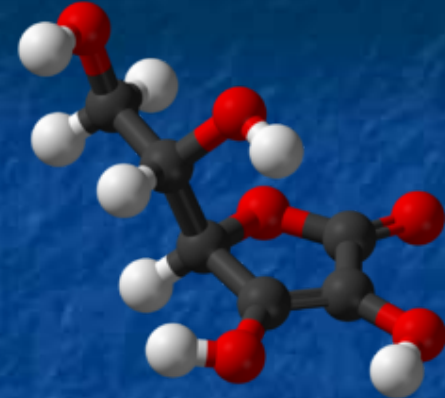
Connect now radicalized end carbons



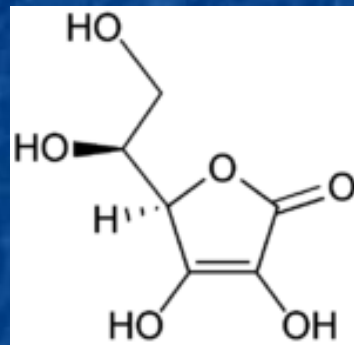
Antioxidant? = Molecules that can SUPPLY electrons

Or, supply Hydrogens that bring along their electrons

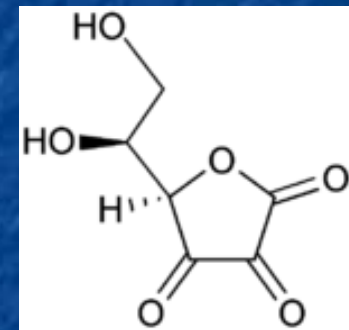
Classic example is VITAMIN C (ascorbic acid):



Normal configuration:



After giving up two Hydrogens:



So “antioxidants” don’t STOP oxidation. They undergo oxidation!

Acting as sacrificial volunteers!

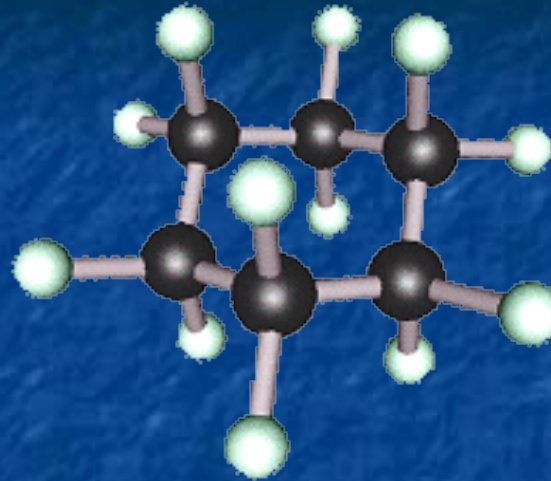
Or, moving toward nanotech headlines, consider rings:

Cyclohexane:

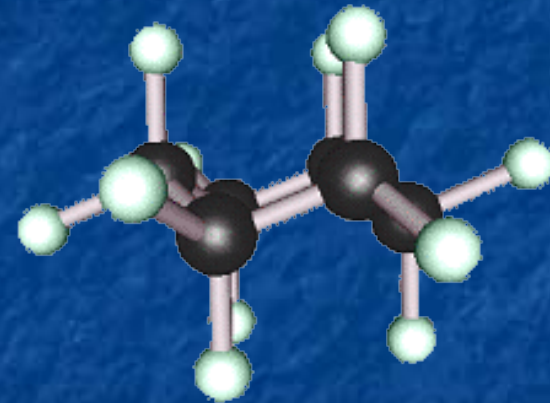
Cyclo = it's a ring

hex = with six atoms

ane = singly bonded



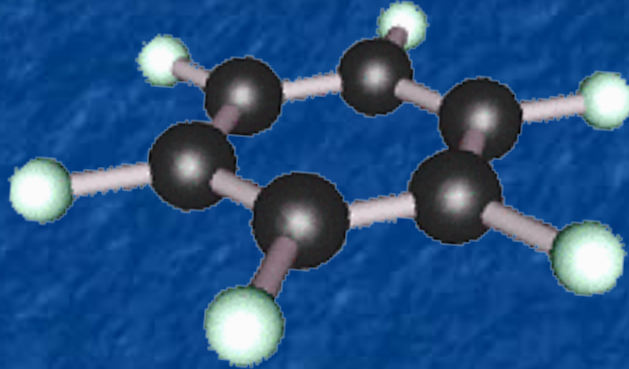
But notice how electrons/bonds
are still in an almost undistorted
tetrahedral geometry:



Figures on this (and the six pages to follow) are screenshots of interactive / manipulatable 3D models on WeCanFigureThisOut.org: WeCanFigureThisOut.org/VL/Nanocarbon.htm

What if we strip off half of the hydrogens and try to flatten ring?

Carbon atoms are not very happy as each now has only three bonds!



Leaving each carbon as (including its neighbor's bonding electrons):



What happens to that very unhappy unpaired electron (at the top)?

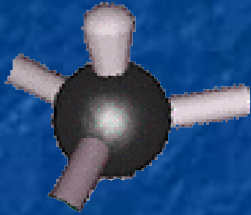
But it's not a B-B, it's more of a cloud

Other electrons are in plane

So leftover electron wants to get OUT of that plane

Two ways to get leftover electron out of that plane:

Alternative #1: Squeeze it upward (or downward) out of the plane



Produces unpaired electron pointing outward

=> **FREE RADICAL** => Oxidizer => Disorganizer

Possible on all "graphitic" surfaces (graphene, nanotubes, Buckyballs)

Alternative #2) Squeeze electron cloud both upward and downward

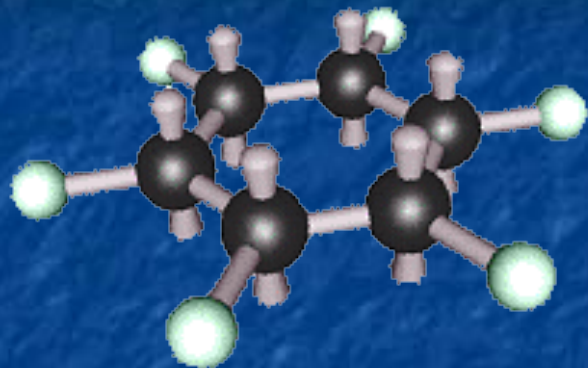
Like squeezing a water balloon in the middle:



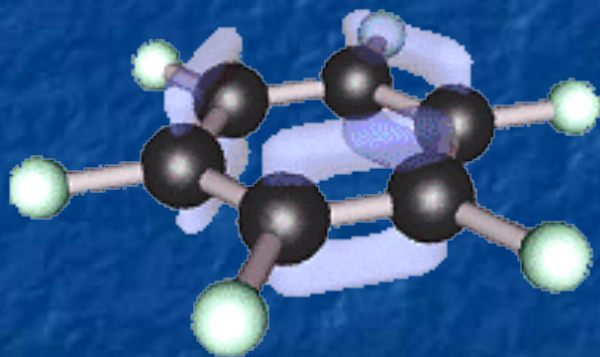
DOES get leftover electron out of the crowded plane

But STILL leaves electron unpaired – how is this any better?

Because this configuration DOES allow for pairing/bonding:

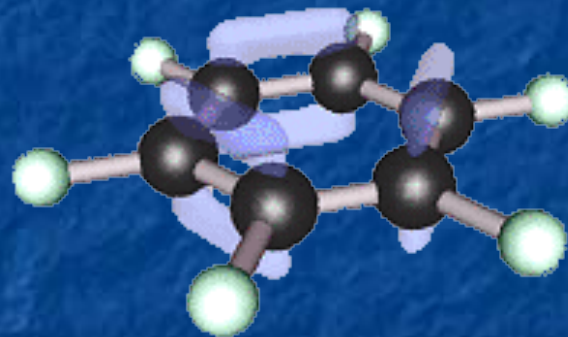


Because those six electrons can now link up (bond) above and below plane:



Pairing #1

OR:

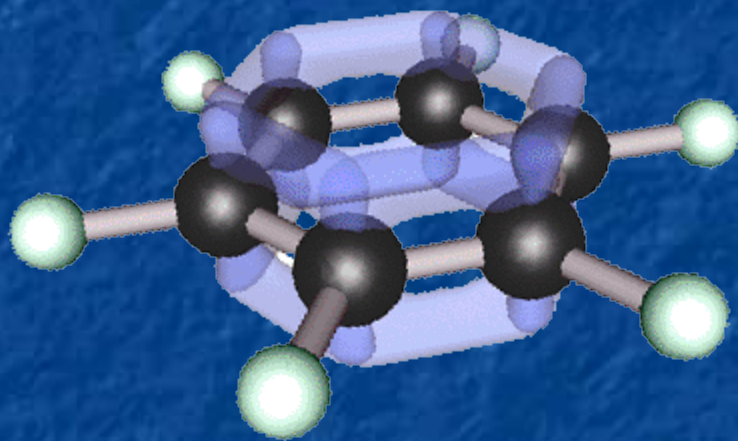


Pairing #2

But which pairing is right?

BOTH! They both do the same thing and save the same amount of energy!

So "nature" USES both - switching rapidly between them giving:



It's called a "resonance bond"

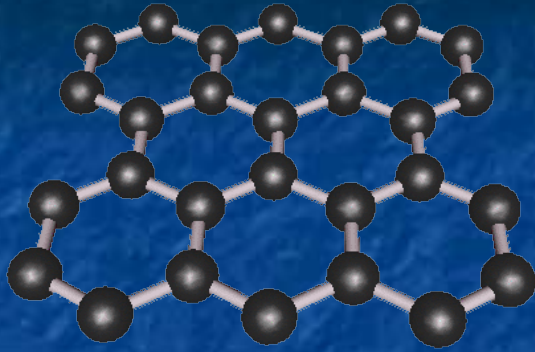
It "resonates" rapidly back and forth between the two linkages!

Significance? "Resonant" electrons are no longer locked into a single position

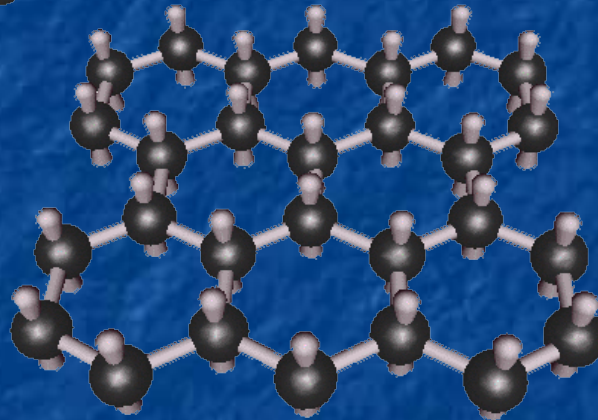
They can move around (for instance under the influence of an electric field!)

This "**Benzene**" molecule can conduct electricity (where cyclohexane cannot!)

Connect the rings and get "Graphene"

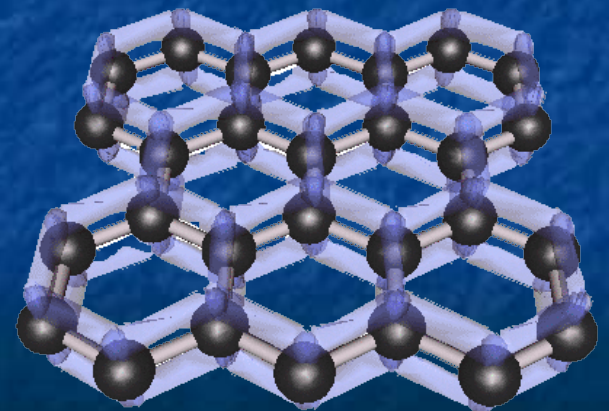


Simple "ball and stick" representation

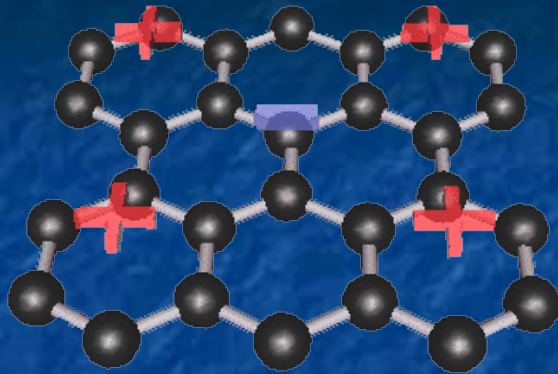


Accounting for all the valence electrons

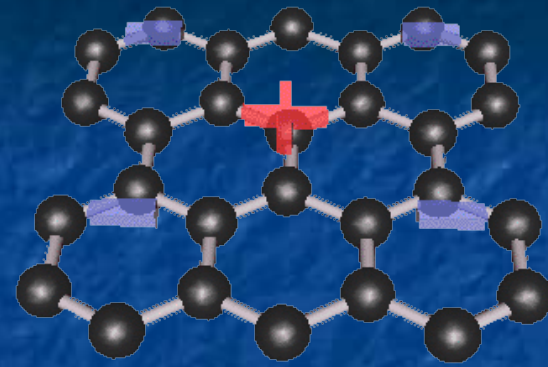
Or showing resonant bonding that gives graphene its electrical conductivity:



But mobile electrons can slosh around forming pockets of +/- charge:



OR:

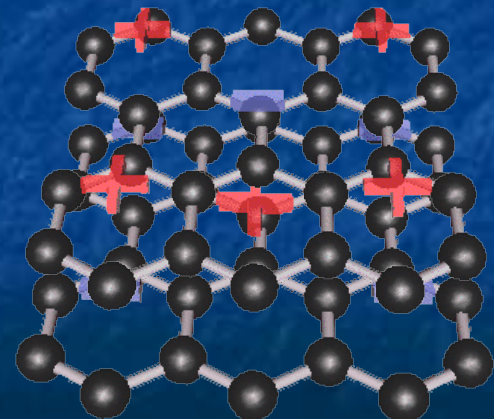


What then happens if one sheet of graphene is positioned above another?

If top graphene sheet randomly assumes charge configuration left/above

Then complementary charging pattern will be **INDUCED** on bottom sheet

And sheets will then be attracted to one another:



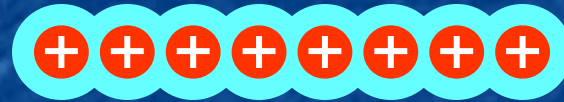
It's called Van der Waals bonding

Very different than the covalent bonding we've been talking about

Van der Waals bonding is weaker: Not based on tight electron magnetic coupling

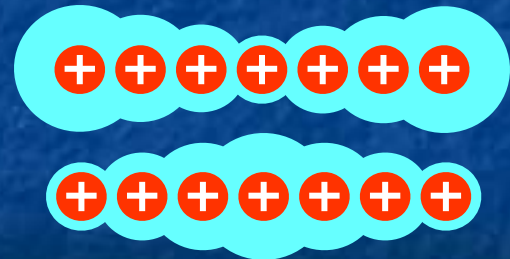
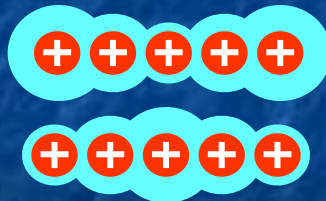
Van der Waals bonding is also not localized: It occurs over regions

What happens if molecule gets larger?



More room for electrons to slosh →

Bigger sloshes (larger deviations from neutrality):



So Van der Waals "bonding" gets stronger in larger molecules!

Proof? Boiling points of similarly structured molecules:

Alkanes



-42.1 C



0.5 C



36.1 C

Alcohols



78 C



97.4 C

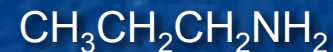


117.3 C

Amines



16.6 C



47.8 C



77.8 C

Within a given family of molecules (column) longer the chain → higher the boiling point

Move between families (columns), different atoms "slosh charge" differently

→ *Shifts whole family's boiling points up or down*

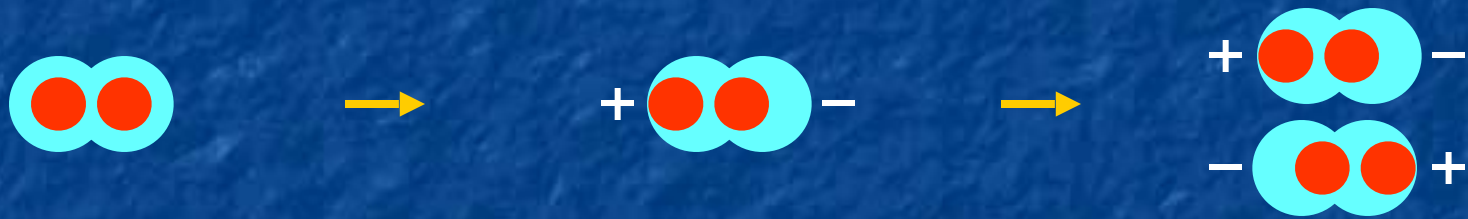
Better name for Van der Waals bonding = "INDUCED-DIPOLE BONDING"

Does this only occur for conductive molecules?

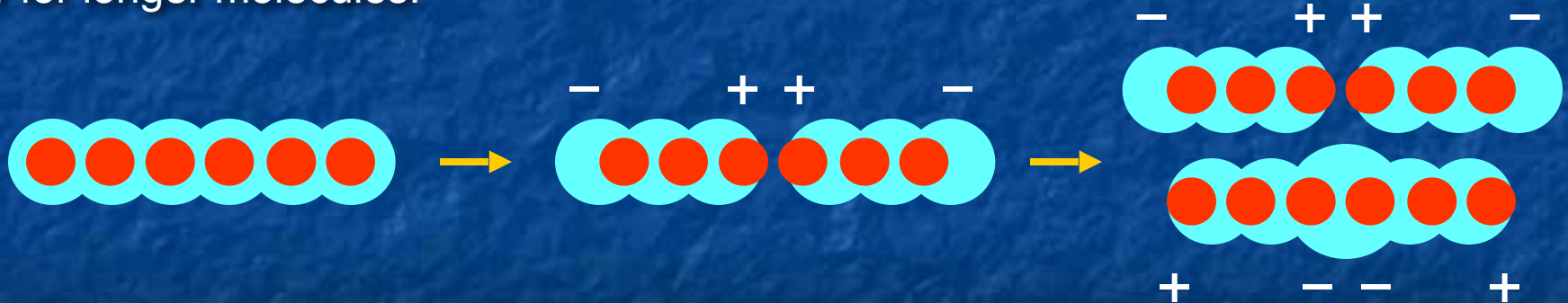
In preceding figures charge shifted from atom to atom

Possible in conductive, semiconductive, or molecules with “conjugated” bonds

But can also get Van der Waal / induced-dipole effects via simpler polarization:



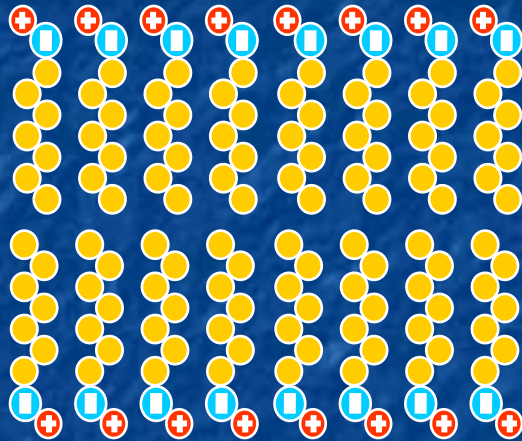
Or for longer molecules:



So polarization of insulating molecules also yields VDW, growing stronger with length

Van der Waals bonding helps hold cell membranes together:

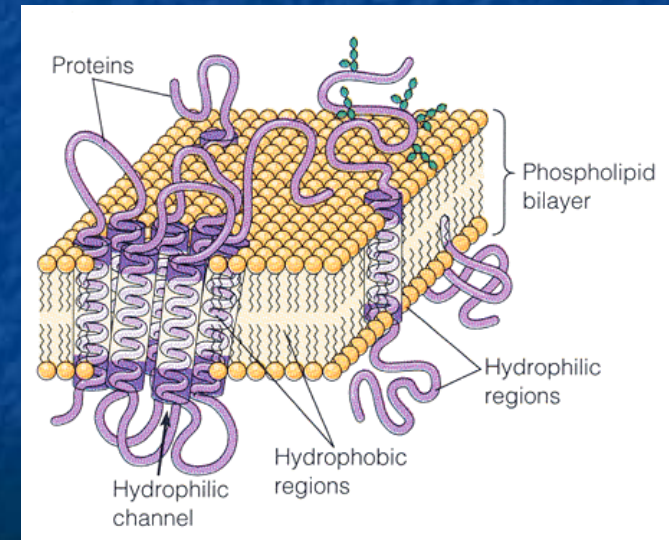
Remember "lipid bilayers" in water where hydrophobic tails gathered together?



Electrons in tails polarize ("slosh")
producing Van der Waals bonding

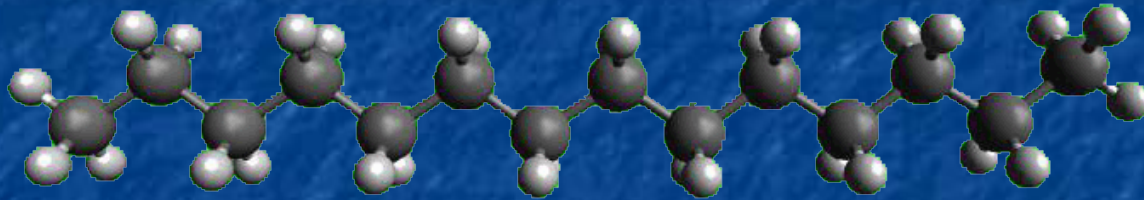
So cell wall self-assembly has two components:

- 1) Hydrophilic heads orient out toward water
- 2) Hydrophobic tails drawn together by VdW

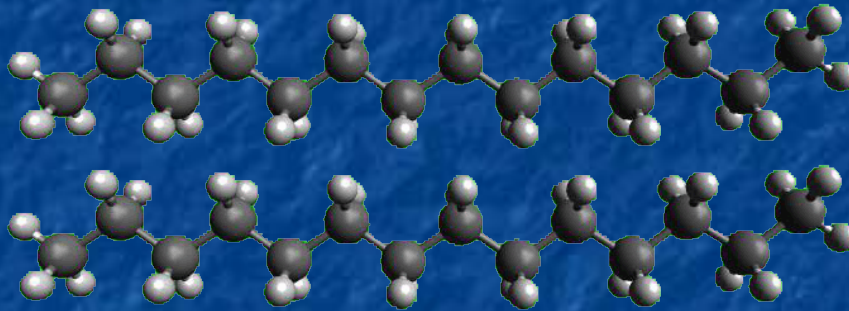


Everyday application: "Saturated Fats"

Here "saturated" means saturated with hydrogens: Saturated carbon chain (fat):



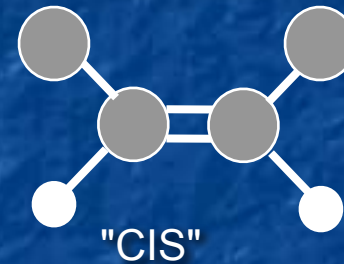
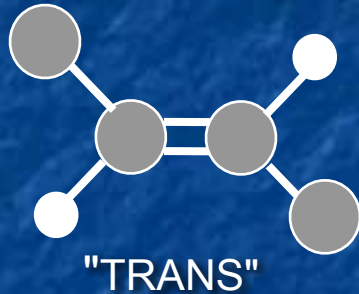
Straight saturated chains pack closely => Strong VdW bonding => High melting T



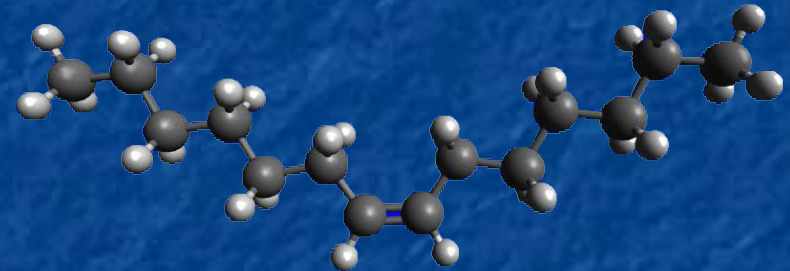
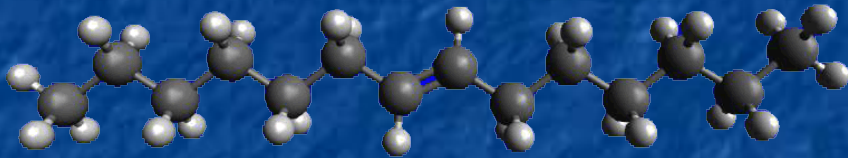
So they can SOLIDIFY within your arteries => Heart Attacks / Strokes

Change to "unsaturated" by removing pair of hydrogens:

Carbons that lost hydrogens form double bond. But 2 configurations are possible :



Producing these alternative full fat chain structures:



TRANS FATS are straight, like saturated fats

Which do pack closely

So they too solidify in our arteries

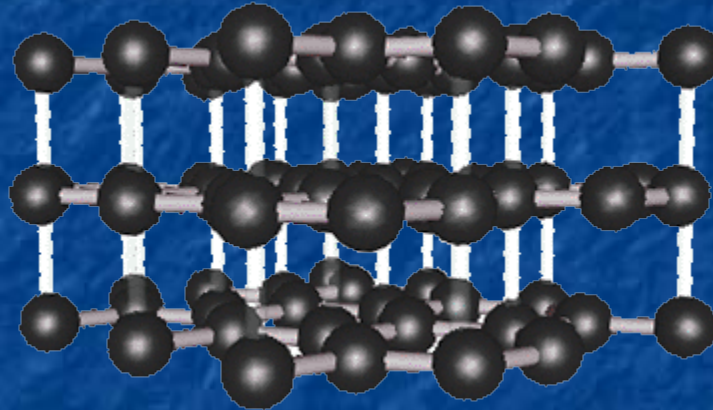
But **CIS FATS** are kinky

Which don't pack closely

So they stay liquid in our arteries

Getting back to graphene

Van der Waals bonding between graphene sheets forms GRAPHITE:



Although not shown, there are also "resonant bonds" → *electrical conductivity*

Weak Van der Waals interplanar bonds → Sheets slip or separate easily

Giving Graphite it's well known lubricating properties

And providing the tape trick we use to obtain fresh "HOPG" layers for STM in lab

What about other deliberately *man-made* molecules?

Field is called "Synthetic Chemistry" It is incredibly powerful!

But kept secret from those of us who never got beyond Chem 101

To describe its richness, we must first learn Organic Chemist's shorthand:

1) Carbon is EVERYWHERE, so get lazy and don't draw it

If no atom identified at intersection of bonds, ASSUME it is C:



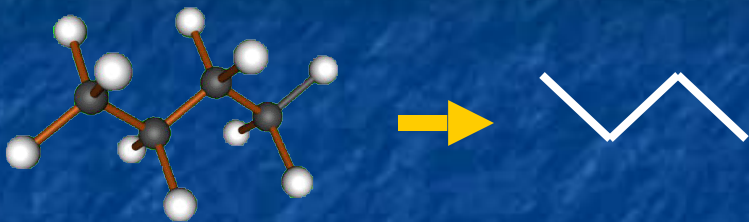
2) In water, hydrogen is ALWAYS available to fill in any bonds left over, so get even lazier

Don't show hydrogen atoms OR their bonds, just infer their presence

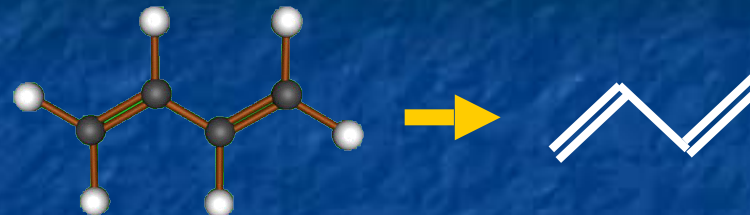
(Can work out number because KNOW carbon must have 4 electrons in bonds)

Which produces shorthand diagrams like the following:

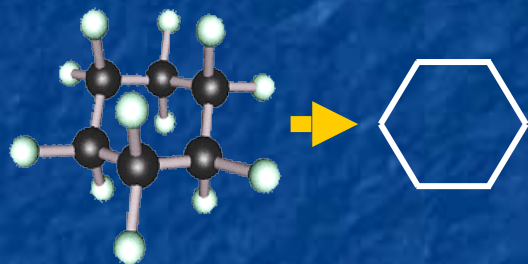
"Fully saturated" alkane:



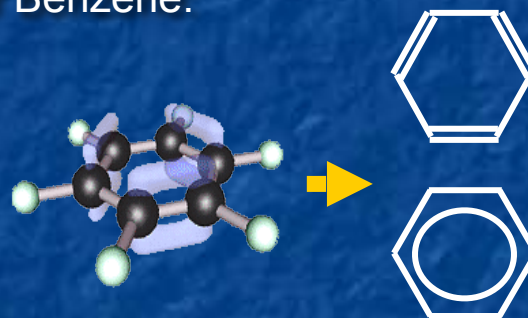
"Conjugated" alternation of single / multiple bonds:



Rings of Cyclohexane:



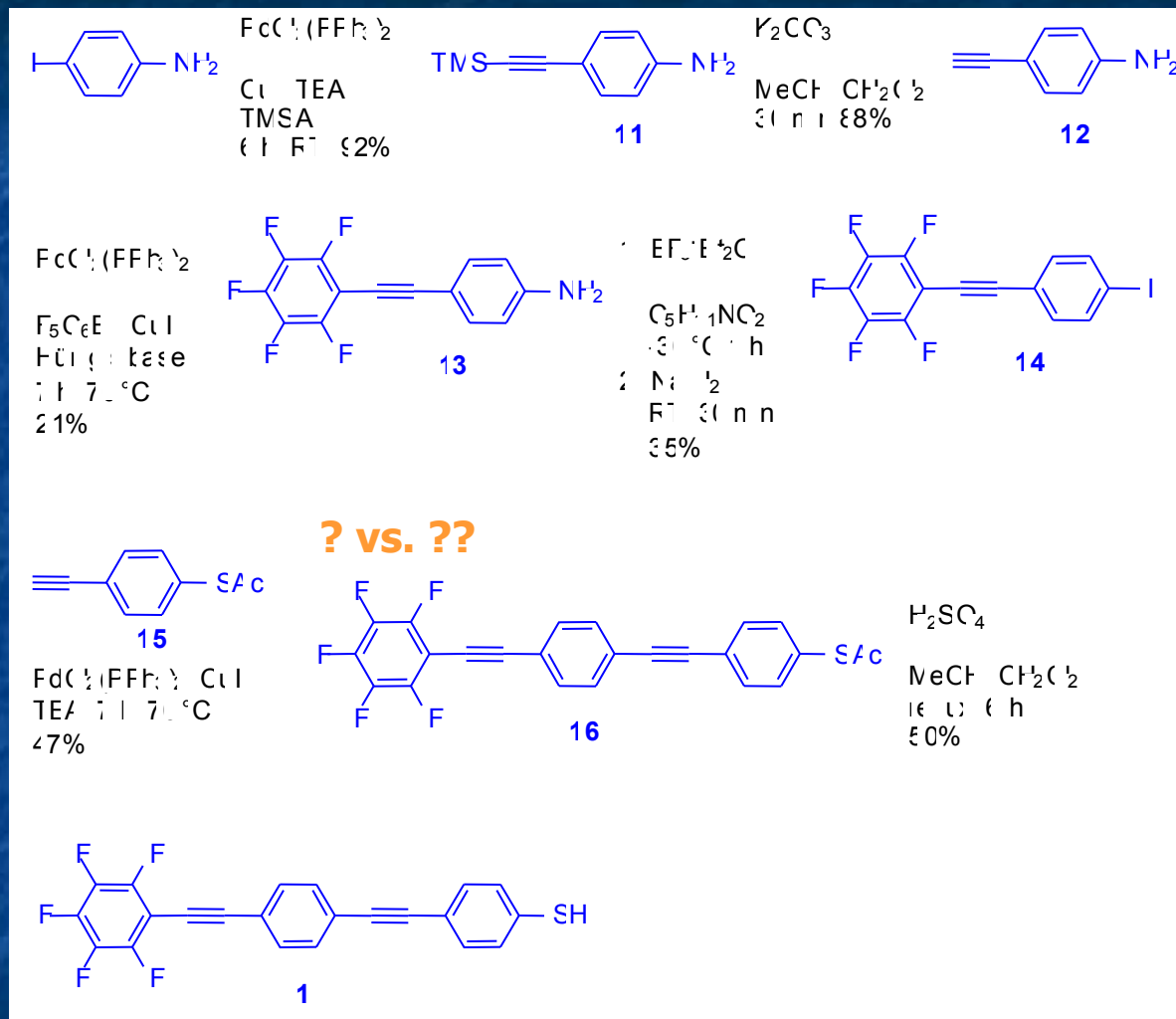
Or Benzene:



Or propyl alcohol (with extra OH group at end represented explicitly):



Chemists build smaller molecules by addition and subtraction:



Step by step they add reagents (black) to remove or add pieces to the starting molecule

Gradually transform it into final product shown at bottom

Can actually pinpoint atoms such as which F is removed from benzene ring (**? vs. ??**)

How? Depending on ring position, F's bond slightly more strongly or weakly

Courtesy of Professor Jim Tour, Rice University
(from our DARPA MOLEapps collaboration)

Followed by purification:

Because steps like those on previous slide are only 30-80% successful

That is, 30-80% of molecules do what was drawn, others do something else

So you must sort out intended from unintended products using:

1) **Distillation:** Different molecules = Different boiling points

Condense different components in progressively cooler regions

Exploits Van der Waals bonding!

2) **Crystallization:** Like molecules try to fit into single crystals, excluding unlike

Remember trying to freeze juice into popsicles? Sugar on outside - ice inside

3) **Chromatography:** Diffuse the mixture through paper, gel . . .

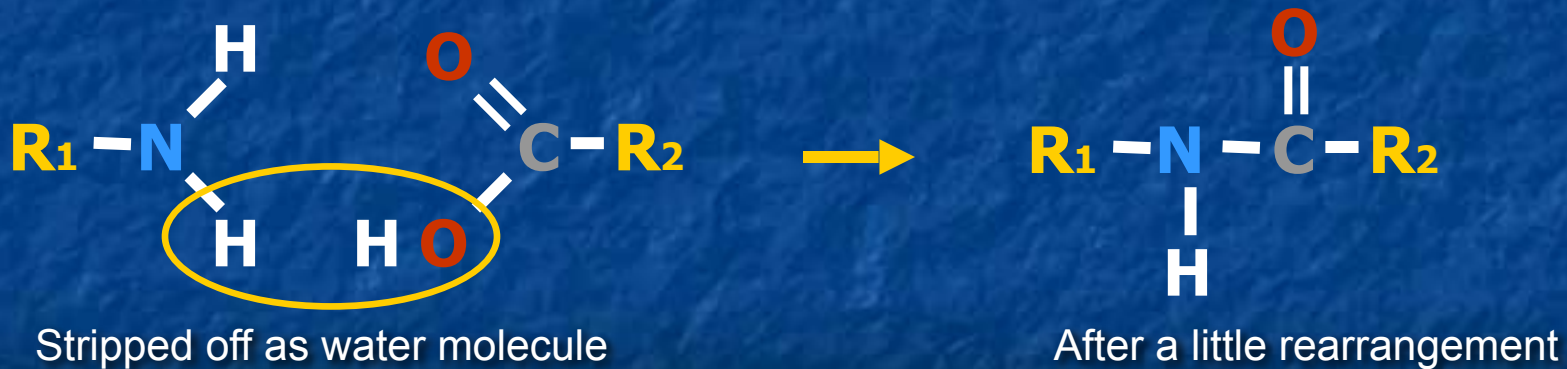
Same thermal energy per molecule = $\frac{1}{2} mv^2$, smaller molecules \rightarrow higher v !

But to self-assemble larger structures, need other tricks

One useful trick employs an amine (left) and a carboxylic acid (right):



Where the following reaction can occur, linking the two sub-molecules:

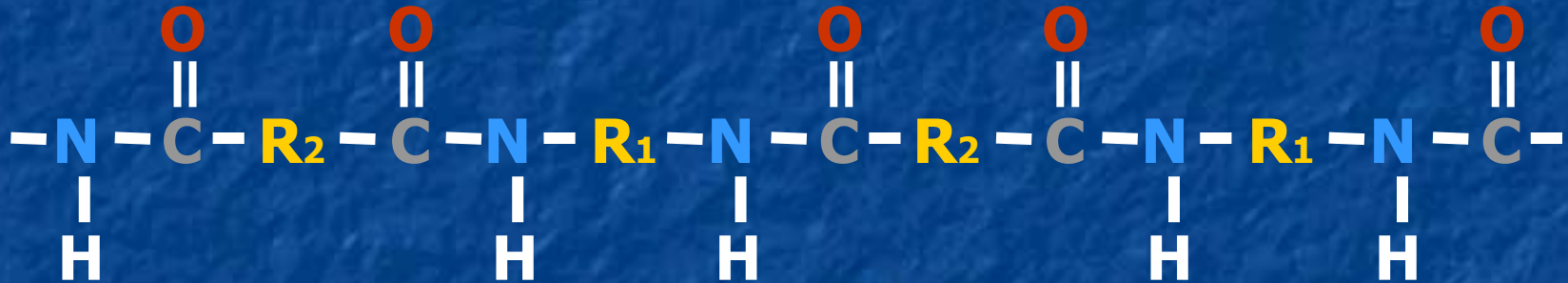


Two every day applications of (just this one!) trick:

Double up reactive groups, one at each end of subunits:



Mix subunits together to produce a self-assembled polymer:



R₁, R₂ = C6 alkane => NYLON

R₁, R₂ = benzene => KEVLAR

That gives you a glimpse of organic chemistry's power

And, via standing waves, finally explains chemists' mysterious "orbitals"

The remaining challenge is to cope with all of the field's crazy names

And to do that, I offer: [John's Organic Chemistry Cheat Sheet](#)

(Derived from my own close collaboration with an organic chemist)

But for the penultimate form of nanoscale self-assembly,

we have got to follow this path all the way into biology

Which will bring us in the next hour to:

Nature's ultimate tool for self-assembly - DNA

Credits / Acknowledgements

Funding for this class was obtained from the National Science Foundation (under their Nanoscience Undergraduate Education program).

This set of notes was authored by John C. Bean who also created all figures not explicitly credited above.

Copyright John C. Bean

(However, permission is granted for use by individual instructors in non-profit academic institutions)