# Experiment 20 —

Pericyclic reactions

Pre-lab preparation. (1) Read text section 22.6. (2) Write the potential Diels-Alder reactions of maleic anhydride and anthracene, and (3) estimate the reaction enthalpy changes for these and for the analogous D-A rxn of naphthalene — the body of this write-up will provide the necessary guidance. Based on your calculations, predict the product of the reaction and explain the difference in Diels-Alder reactivity between naphthalene and anthracene. (4) Based on the quantities of reactants in the procedure, determine the theoretical yield of product. (5) Look up the boiling points of xylenes and of ethyl acetate. What's the purpose of rinsing the crystals with ethyl acetate before drying? (6) According to Wikipedia, what excited states are produced in the sequence of reactions that leads to chemiluminescence from luminol?

In the first part of this lab we will do a Diels-Alder cycloaddition; in the second we will carry out a chemiluminescent reaction whose key step is a cycloreversion akin to a retro-Diels-Alder reaction.

Diels-Alder reaction of anthracene and maleic anhydride. The Diels-Alder cycloaddition is a concerted 6-electron reaction that involves a 4-π-e^- component, usually a diene, and a 2-π-e^- component, called the *dienophile*. The text has several examples; one additional example, the reaction of cyclopentadiene and maleonitrile, is illustrated below.
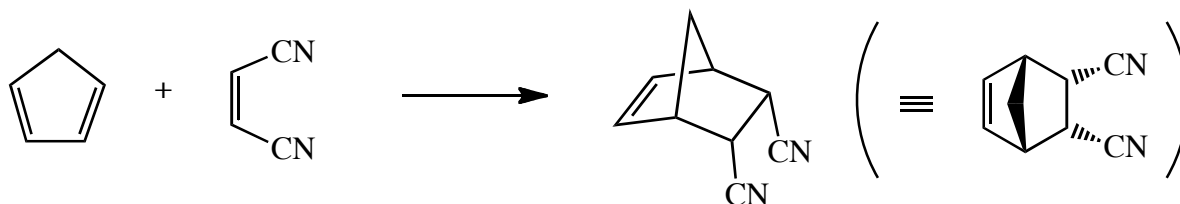

The Diels-Alder reaction is classified as a thermally "allowed" pericyclic process, meaning that the concerted pathway is expected to have a relatively low-energy transition state. The reason that such a process should be especially favorable can be understood in several

different ways. The most straightforward analysis recognizes that the transition state involves 6 electrons in a 6-orbital cycle. This is shown for the Diels-Alder cycloaddition of ethylene and 1,3-butadiene below. Because this transition state orbital array has the same basic topology (more on that in class) and the same number of electrons (4n+2) as the π-system of benzene, it is aromatic. Although the transition state almost certainly has less aromaticity than benzene, the stabilization is still significant.

An alternative analysis recognizes that there is a favorable (all in-phase) interaction between the "frontier orbitals" of the diene and dienophile. For example, interaction of the diene HOMO and dienophile LUMO (below) results in a pair of electrons going down in energy. Similarly, the diene LUMO interacts favorably with the dienophile HOMO, sending a second pair of electrons down in energy. Both of these filled-empty orbital interactions contribute to a stabilization of the transition state and a favorable reaction. The predictions made by these two methods are almost always in agreement, and almost always in accord with experimental observations as well.
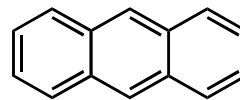
electron-pushing   **6-p-orbital array**   **frontier orbital analysis**
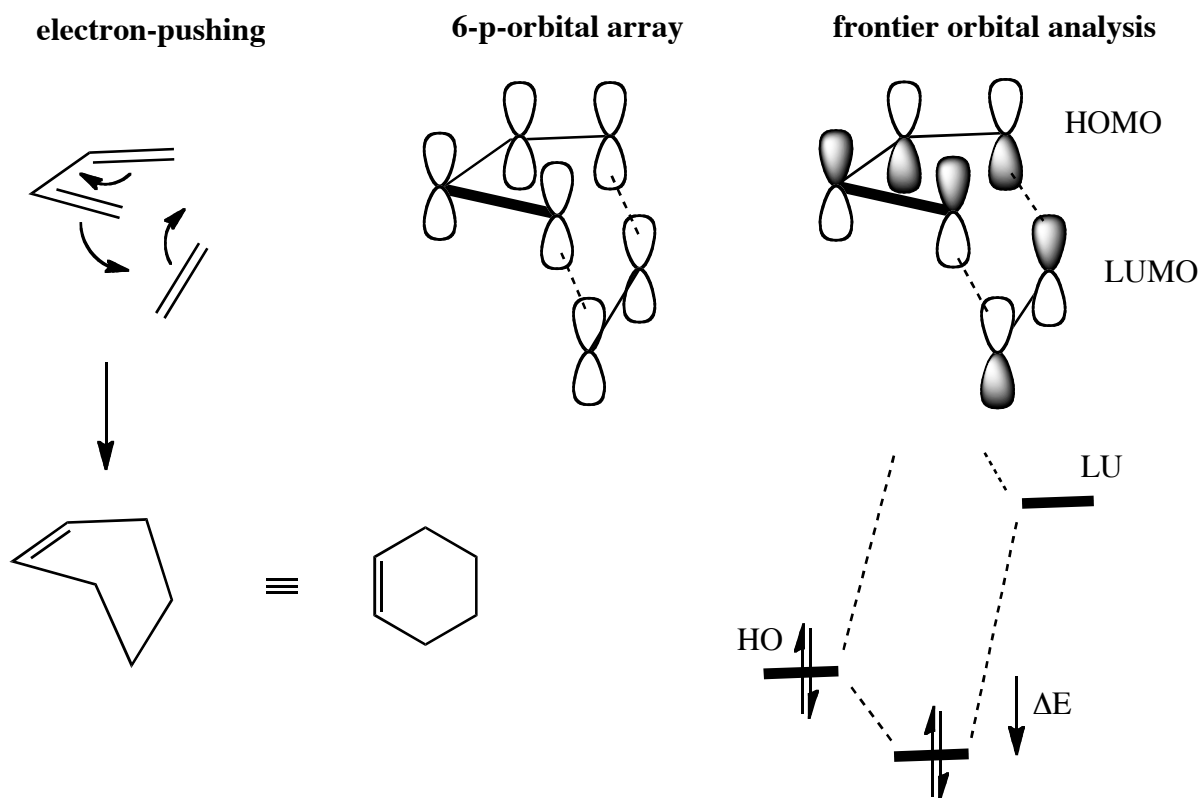
In a typical Diels-Alder reaction, neither π-system is part of an aromatic ring. But aromatic compounds — benzene for example — have resonance structures with embedded "diene" units that could, at least in principle, participate in Diels-Alder reactions. So it may at first seem surprising that benzene is completely unreactive, as is naphthalene. Of course the reason must have something to do with the desire of these compounds to retain their aromatic cycles. The aromatic stabilization energy of benzene (relative to an *isolated π-bond* reference) is 32 kcal/mol, and that of naphthalene is 53 kcal/mol. But, surprisingly, the next member of the series, anthracene, which has an even greater stabilization energy of 72 kcal/mol, *does* participate as the diene component in Diels-Alder reactions. Why?
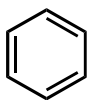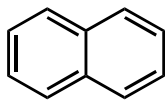
benzene   **naphthalene**   **anthracene**

Anthracene has 4 resonance structures, one of which is shown above. One could imagine several 4-carbon units within the anthracene π-system participating as the diene component in a Diels-Alder reaction, but let's consider just the two possibilities indicated below — reaction of the dienophile at positions a and at positions b — which would result in two different products. Draw the Diels-Alder cycloadducts (the products of the cycloaddition) of anthracene and maleic anhydride that would result from reaction at these positions.
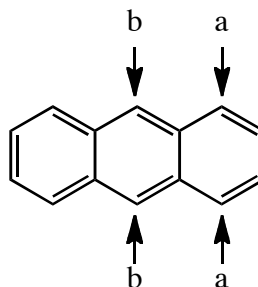
Now let's estimate ΔH° for these two reactions. There are three factors that need to be accounted for: (1) Changes in bonding, (2) changes in strain energy, and (3) changes in aromaticity.

First, let's evaluate the contribution of bonding changes to the reaction $\Delta H°$. The simplest possible Diels-Alder reaction is that between ethylene and 1,3-butadiene, shown below. The reaction involves a net exchange of 2 π-bonds (worth about 65 kcal/mol each) for 2 σ-bonds (worth about 80 kcal/mol each). This process is therefore exothermic by about 30 kcal/mol, i.e., $\Delta H° \approx -30$ kcal/mol. It is this exchange of π-bonds for σ-bonds that makes the Diels-Alder reaction thermodynamically favorable.

break 3 π bonds   make 1 π and 2 σ bonds
3 x +65 kcal/mol   1 x –65 + 2 x –80 kcal/mol
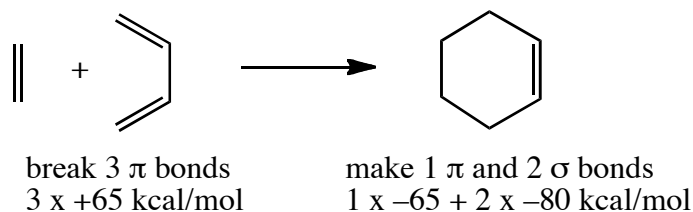
Of course, because two molecules are coming together to make one, the entropy change is unfavorable, so the reaction is not quite 30 kcal/mol downhill in free energy, though it is still considerably downhill in $\Delta G°$. Since we can't easily estimate entropy changes, we'll continue to focus our attention on understanding the enthalpy changes in these cycloadditions.

In the example above, strain and aromaticity were not factors. But when a dienophile adds to a cyclic 6-membered ring diene, e.g. 1,3-cyclohexadiene, or to a "diene" unit embedded in a benzenoid aromatic ring, a bicyclo[2.2.2] ring system is created. This ring system has about 10 kcal/mol of strain energy. So given that the bonding changes are responsible for a 30 kcal/mol decrease in enthalpy (as above), but the bicyclic product is strained by 10 kcal/mol, we can estimate $\Delta H°$ for a D-A reaction with a 6-membered ring diene as about –20 kcal/mol. We can use this as our "base" value, and add to it the contributions from changes in aromaticity.

Before we attempt to estimate the enthalpy changes for the two possible D-A reactions of anthracene, let's make the estimate for the D-A reaction of benzene. If a dienophile were to add to benzene, there would be a 32 kcal/mol loss of aromaticity. We would then estimate $\Delta H°$ to be –20 + 32 kcal/mol = +12 kcal/mol, i.e. endothermic, or *uphill*, by 12 kcal/mol. Again, that's –20 to account for changes in bond strengths and strain energy, and +32 to account for the loss of aromatic stabilization.

Now it's your turn. Estimate the enthalpy changes for the cycloadditions of maleic anhydride to one of the outer rings (the "a" positions) and to the central ring (the "b" positions) of anthracene. Also estimate the $\Delta H°$ for an analogous D-A cycloaddition to naphthalene. Based on your results, explain why anthracene does D-A reactions, while naphthalene does not, and predict the product of the cycloaddition of maleic anhydride to anthracene. Note that only the C=C unit of maleic anhydride reacts, and the small strain present in its 5-membered ring does not change, so whether the dienophile is maleic anhydride or ethylene makes no difference as far as the $\Delta H°$ estimate is concerned.

You are encouraged to complete these calculations as part of your pre-lab write-up, but if you feel that you cannot do them properly, you may do them in lab *before beginning* your experiments and turn them in as part of your report.

Procedure. *Maleic anhydride and the product of the reaction are severe skin irritants. Be sure to wear gloves and avoid contact with these compounds.*

In a dry round-bottom flask, combine 1.0g of anthracene and 0.5 g of maleic anhydride, add 12 ml of xylenes (a mixture of dimethylbenzenes), and reflux the reaction mixture for about 30 min. Use an IR lamp as the heat source. Allow the mixture to cool to room temperature. Cooling in ice for 5 - 10 min should complete the crystallization of product. Collect the crystals in the usual way, and rinse with a few ml of cold ethyl acetate. Then make the usual filter paper sandwich to sop up the excess solvent, and pop 'em into the over (ca. 100°C) for 15 min.

Measure the mass of your dried solid and determine the yield. If you get more than a 50% yield, you will get an NMR spectrum of the product, from which you can determine which of the two possible modes of cycloaddition actually occurred. In addition to the usual procedural write-up with data and observations, your report — *due at the end of the period* — should include a detailed analysis of the spectrum and an explanation of the structure assignment. If you get less than a 50% yield, you'll need to (a) do the reaction again, (b) draw the two possible cycloadducts on the board 50 times each, or (c) clean all the glassware in the stockroom. The instructor will determine which option is appropriate or may elect to change "or" to "and".

If you're lucky enough to earn an NMR spectrum, you'll find that the aromatic protons display splitting of a type that we haven't encountered before — the 4-line multiplets come from a phenomenon called "virtual coupling". To understand this splitting, the sets of strongly-interacting equivalent Hs have to be treated as "spin system", and the interaction of these spins, as a group, with neighboring Hs leads to the observed patterns. But to determine the structure of the D-A adduct from the spectrum, you don't need to be concerned with understanding this in detail — focus on the number of signals, their chemical shifts, and their relative integrations.

Chemiluminescence from luminol. Chemiluminescence is light emission produced chemically, i.e. from reactions of molecules in their ground electronic states. This is a *very uncommon* phenomenon that happens only in very special cases — the *vast* majority of chemical reactions most certainly *do not* produce light! Chemiluminescence from luminol was first observed over 80 years ago. The light emission is triggered by reaction of luminol dianion with oxygen (perhaps molecular oxygen, perhaps something that creates the equivalent of an oxygen molecule). A possible mechanism for this process is shown below. This oxidation is usually initiated by the reaction of hydrogen peroxide with a metal-based catalyst. Since its discovery, luminol has become one of the most well-known chemiluminescent compounds, mainly due to its use by crime scene investigators to detect traces of blood. When luminol and hydrogen peroxide are sprayed on a surface, the iron in hemoglobin catalyzes the oxidation that leads to detectable light emission.
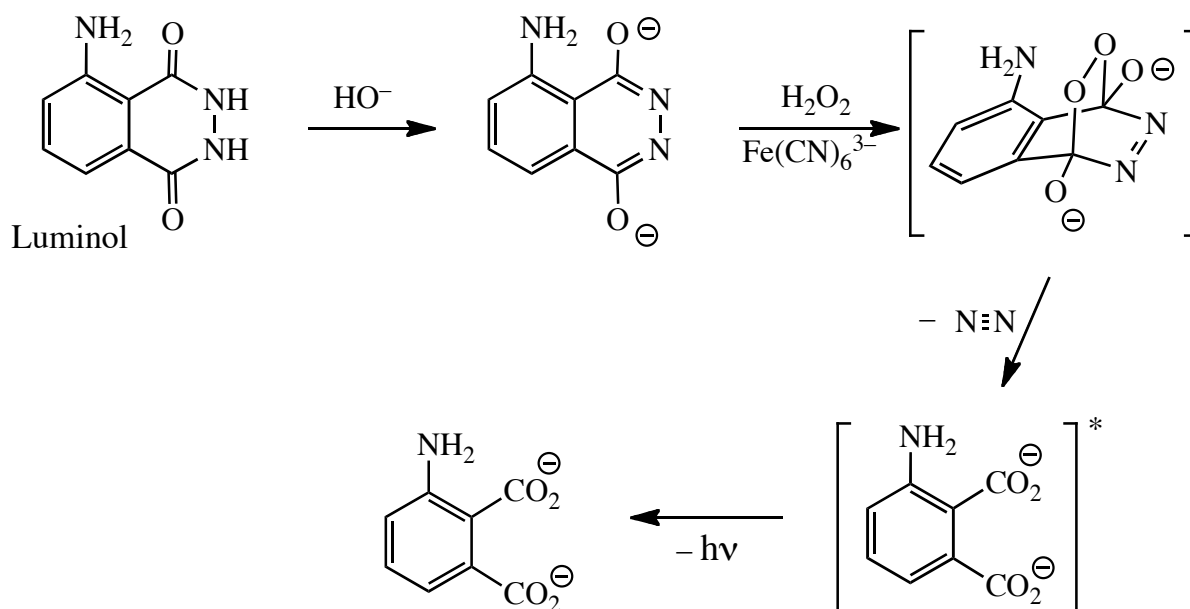
Procedure. You can do this one while you're waiting for the Diels-Alder reaction. Dissolve 100 mg of luminol in 2 ml of 10% NaOH, and dilute this solution with water to a volume of about 40 ml. In a separate flask, combine 4 ml of 3% aq $K_3Fe(CN)_6$ (potassium ferricyanide) and 4 ml of 3% aq H_2O_2, then dilute this mixture with water to a volume of about 40 ml.

Summon the instructor/TA so we can watch, and take the two solutions into the dark chemical storage room. Pour one into the other. Or you can make the effect even more dramatic by pouring them simultaneously through a funnel into a clean flask. If your instructor can verify light emission, he may decide to waive the penalties for getting less than a 50% yield from the D-A part. And — shocking but true — there's no write-up required for this part of the lab!

When finished, (1) discard the solutions into the waste, (2) place your triphenylmethanol and any other solid products into the appropriate collection bottles (note that we'd like to save these for use in future labs), and discard any other compounds that you have squirreled away over the year. Appropriate solid and liquid waste containers will be provided, as always.

Next week will be check-out, with a special surprise during the pre-lab lecture period for those who show up on time.