Loudon Ch. 3 Review: Acids/Bases/Curved Arrows

Jacquie Richardson, CU Boulder – Last updated 9/16/2020

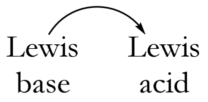
There are two different definitions of acids and bases that show up in this chapter:

1. Lewis acids (a.k.a. **electrophiles,** E or E^+) accept an electron pair; Lewis bases (a.k.a. **nucleophiles,** Nu or Nu^-) donate an electron pair
2. Brønsted-Lowry acids (a.k.a. **acids**) donate a proton (H^+ ion); Brønsted-Lowry bases (a.k.a. **bases**) accept a proton

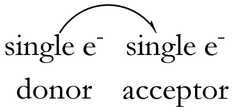
Brønsted-Lowry acids/bases are a subset of Lewis acids/bases. If something is a Brønsted-Lowry base it must also be a Lewis base but the reverse is not true. When an organic chemist says "acid" or "base", what we normally mean is Brønsted-Lowry acid or base. We use the terms **nucleophile** for Lewis bases (things that love nuclei), and **electrophile** for Lewis acids (things that love electrons). Confusingly, chemists will often use "nucleophile" to mean "only-nucleophile-but-not-a-base". Officially, though, bases are a subset of nucleophiles, and acids are a subset of electrophiles. We're going to look some more at the differences between them, but first, here are some tools to help explain what's going on during a reaction: curved arrows, and frontier molecular orbitals.

Curved Arrows Show Electron Movement

We can show the flow of electrons to a Lewis acid by using curved arrows. These are used to show the movement of **electrons only**, not anything else! Two electrons moving as a pair (like in a reaction between a Lewis acid and base) are shown as a full-headed arrow. (One electron moving by itself – a radical – is shown as a half-headed or fishhook arrow. We'll cover these later.) The Lewis base "attacks" the Lewis acid by pushing electrons towards it.



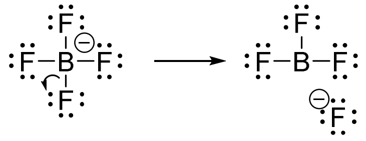
2 electrons move: full-headed arrow



1 electron moves: half-headed (fishhook) arrow

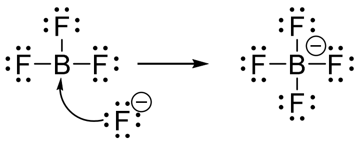
In other words, these arrows are used to take electrons from electron-rich areas (nucleophiles), and give them to electron-poor areas (electrophiles). There are only three legal moves for curved arrows. (This is also a good time to mention that organic chemists will usually neglect to show lone pairs if they're not involved in a particular step of a mechanism. But generally, if you're showing one LP on an atom, you should show all LPs on that atom.)

1. Bond to an atom → lone pair on that same atom (in this case, electrons from B-F bond become a new lone pair on F.)



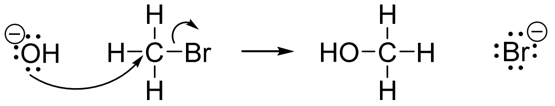
   The electrons come out of the F-B bond and recreate the fourth lone pair on F. Note that the total charge is conserved: there's an overall -1 charge before the reaction, so there has to be an overall -1 charge after the reaction. Since B gave up electrons during the attack (it's at the back end of the arrow), it loses its negative charge. Since F gained electrons during the attack (it's at the front end of the arrow), it gains a negative charge.

2. Lone pair on an atom → bond to that same atom (in this case, lone pair on F becomes a new B-F bond.)

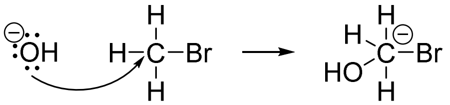


This is the reverse of the previous example: one of the lone pairs on F$^-$ is sent towards B, creating a new F-B bond afterwards. Again, charge is conserved, and the atom where the arrow ends (the B) gains negative charge while the bottom F loses its negative charge. Note that before this reaction, B has an unfilled octet and needs to accept electrons to fix that. Unfilled-octet species are a common example of Lewis acids.

Here's another example of this type. In this case, the lone pair on O becomes a new O-C bond. Note that there's also a second arrow happening here that converts the C-Br bond into a LP on Br at the same time.



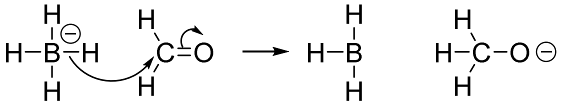
Not all Lewis acids have unfilled octets. Some have a full octet but also a δ$^+$ charge due to polar bonds, like the C atom shown above. It's tempting to show the same kind of reaction as in the first example, but if we only add new electrons to carbon then we'll violate the octet rule by having ten electrons around carbon.



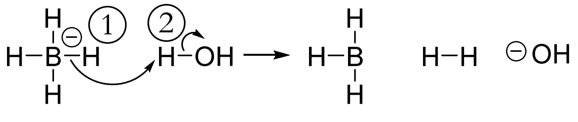
Octet violation = bad!

Instead, something else needs to leave the carbon to free up some space. This makes the reaction an **electron pair displacement**. In this case the lone pairs on O are attacking the C and displacing the electrons out of the C-Br bond.

3. Bond to an atom → different bond to that same atom (in this case, B-H bond becomes a new H-C bond.) This is another displacement reaction, but this time, we're breaking a π bond (the C=O) instead of a σ bond. To keep its octet, the C must lose a bond when it gets attacked, and the easiest bond to break is the π bond. Even though B has a negative charge, it does not have a lone pair to attack with. Instead, it attacks with the electrons out of the B-H σ bond. Note that this is not an acid-base reaction, because there is no proton being moved, but there is a hydride. The difference is that a hydride (H$^-$) brings electrons with it to do an attack, while a proton (H$^+$) doesn't.



Here's another example of bond-to-bond electron transfer. Again, the B-H bond electrons are being used to form a new H-H bond, which displaces the OH group. This reaction has both a hydride (1) and a proton (2).



1: hydride (H$^-$) keeps the electrons it started with

2: proton (H$^+$) doesn't keep the electrons it started with

Here's a more complicated example; this is part of a reaction we'll see in Ch. 5. Here, the alkene's π electrons attack one of the Br atoms (bond to bond), which displaces the other Br and sends it off with a negative charge (bond to LP). Meanwhile, the LP electrons on the first Br are attacking one of the C atoms from the alkene (LP to bond).

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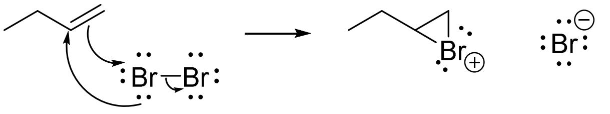
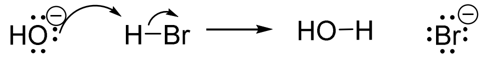
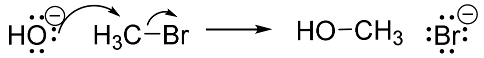


Bases vs. Nucleophiles

The only difference between the two reactions below is what group is being attacked by the HO⁻ - if it's a proton, the HO⁻ is acting as a base, and if it's anything other than a proton (in this case a methyl or CH_3 group), the HO⁻ is acting as a nucleophile only.

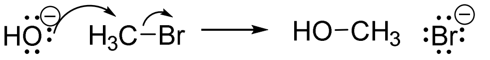


Base Acid

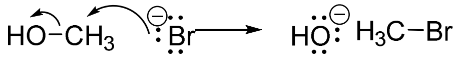


Nucleophile Electrophile

The group that gets displaced is the **leaving group** (LG). This group takes the electrons from its former bond with it when it leaves. This means that if you showed the same reaction in reverse, the leaving group would become the base/nucleophile and vice versa.



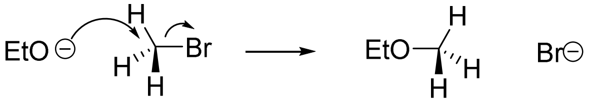
Nu^- E LG



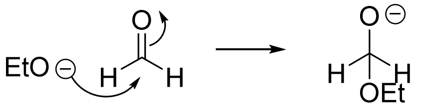
E Nu^- LG

Frontier Molecular Orbital Theory

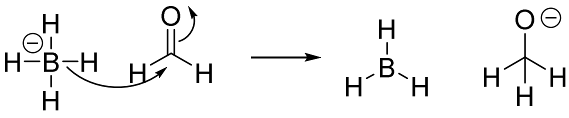
We know that electrons have to be located in orbitals, so which orbitals exactly are involved in a mechanism? This is described by Frontier Molecular Orbital (FMO) Theory. Whenever a mechanistic step occurs, electrons are being transferred from one orbital to another during the attack. The orbital where electrons are attacking from is the **highest occupied molecular orbital** (HOMO) and the orbital that is being attacked is the **lowest unoccupied molecular orbital** (LUMO). Often the HOMO is a lone pair (nonbonding or nb) orbital, but it can also be a σ bond or π bond. The LUMO will often be an antibonding orbital – in the first example below, the C-Br σ bond breaks as a result of the attack, so this means the incoming electrons must have gone into the σ* orbital to reduce its bond order to zero. This can also happen to π bonds, as in the second and third examples.



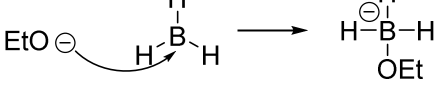
HOMO = nb (lone pair) on O
LUMO = C-Br σ*



HOMO = nb on O
LUMO = C-O π*



HOMO = B-H σ
LUMO = C-O π*



HOMO = nb on O
LUMO = nb on B

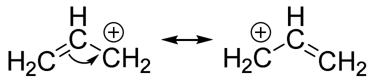
It won't always be so clear-cut as this. In the example at the top of the previous page, there's no immediately obvious HOMO or LUMO because the two molecules are attacking each other at once.

Showing Resonance with Curved Arrows

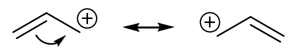
We've already seen how to draw resonance structures for simple molecules in Ch. 1, but for more complicated molecules it can sometimes be difficult to figure out multiple Lewis dot structures when starting from scratch. For this reason, curved arrows are helpful to move electrons around and generate a new resonance form from a structure you already know. Some rules to know are:

1. The number of total electrons must stay the same. If there's an overall negative charge on one form, there must be an overall negative charge on all forms.
2. Arrows follow the same rules for normal arrow pushing, so only three moves are legal: bond to LP, LP to bond, or bond to bond.
3. Nuclei don't move around, only electrons. So you can't change the connectivity of a structure by moving an atom to somewhere else. This means that π bonds and charges tend to be the parts that move around most often. (There are some resonance forms that involve breaking σ bonds, but these mostly only come up in highly polarized bonds like in H-Cl).
4. As above, if you're trying to add electrons to an atom with a full octet, then that atom has to relinquish electrons to somewhere else.
5. Out of all the resonance forms, if some have filled octets on all atoms and some don't, then the filled-octet forms will be the most stable.
6. If all resonance forms have the same number of unfilled-octet atoms, then the most stable forms will be the ones with the fewest total charges, or with negative charges on more EN atoms and positive charges on less EN atoms.

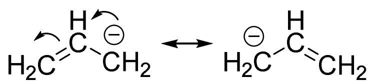
With this in mind, we can look at an example. For clarity, this is shown two different ways: with the atoms explicitly written, and as skeletal structures. Here, the positive carbon on the right end of the molecule is electron-deficient, so the π bond donates some electron density to it. This leaves the leftmost carbon electron-deficient afterwards.



or



As another example, we can look at an electron-rich molecule. In this case the negative carbon is sending its electrons to create a new π bond with the central carbon, but this means the central carbon has to drop it existing π bond, which becomes a new lone pair elsewhere.



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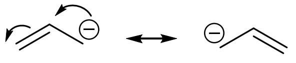

These two examples together show a very common pattern. Often a molecule will have a three-atom group where two of the atoms are connected by a double bond, and the third atom has a charge or lone pair. You can use resonance arrows to show the two end carbons changing so that the one that was double-bonded before now has a charge and vice-versa. In the example below, note that each step involves a three-atom group (shown with bold lines) where the double bond swaps places with the charge.

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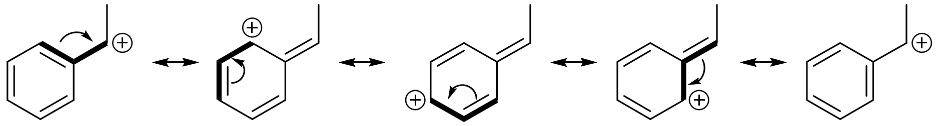
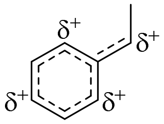
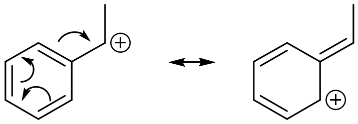


Average form:
(charge on
alternating
atoms)

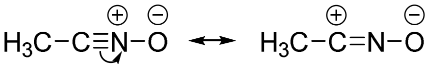


Arrows can be combined
to get from any resonance
form to any other:



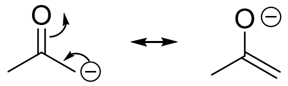
This tends to give us patterns where the charge is only located on alternating carbons, but this will not always be the case if there are heteroatoms (anything other than C and H) involved. Note that we can combine arrows to get from any resonance form to any other resonance form, so writing the structures in any order is acceptable so long as the arrows make sense.

Here are some examples that cover the last two rules.



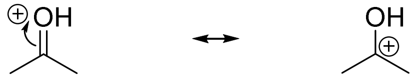
All atoms have filled octet
More stable

Carbon has unfilled octet
Less stable



Negative charge on C
Less stable

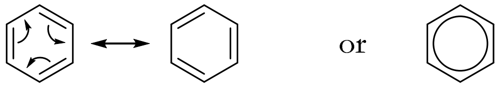
Negative charge on O
More stable



Positive charge on O, but everything has filled octet
More stable

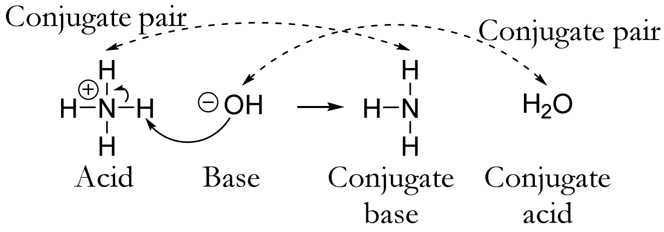
Positive charge on C, but C has unfilled octet
Less stable

One more example: something with no charges, only π bonds. (This shows that all the C-C bonds in benzene are equivalent, which is why it's drawn with a circle.)



Brønsted-Lowry Acids and Bases

These can be shown with the same curved arrows as before, but now we're specifically looking at attacking a proton. When an acid loses a proton, it becomes a conjugate base. When a base gains a proton, it becomes a conjugate acid. Whether a structure is gaining or losing a proton, the two forms are called a conjugate pair. The stronger an acid is, the weaker its conjugate base is, and vice versa.



Compounds that can act as either an acid or base are called **amphoteric**. Water is the most common example. This means it's part of two conjugate pairs at once.

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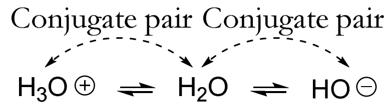


Measuring Acid and Base Strength

This is determined by the equilibrium position for an acid-base reaction with water. The equilibrium constant, K_{eq}, is determined the way it normally is: all of the product concentrations, divided by all of the reactant concentrations. However, since H_2O is a reactant and its concentration is assumed to be constant, this term can be multiplied out to give K_a, the acidity constant. Note that the equilibrium arrow has two lines, unlike the resonance arrow. These are two different things: equilibrium means there is an actual reaction taking place reversibly, while resonance means we're just drawing the same structure in multiple different ways.

$$HA + H_2O \rightleftharpoons A^{\ominus} + H_3O^{\oplus} \qquad K_{eq} = \frac{[A^{\ominus}][H_3O^{\oplus}]}{[HA][H_2O]} \qquad K_a = K_{eq}[H_2O] = \frac{[A^{\ominus}][H_3O^{\oplus}]}{[HA]} \qquad pK_a = -\log(K_a)$$

Stronger acids have an equilibrium that favors the products more, so the equilibrium will lie farther to the right. This means that **stronger acids have higher K_a values**. However, the possible range of K_a values covers so many orders of magnitude that we usually put it into more manageable terms. We can do this by converting K_a to pK_a. The "p" of something is the negative base-ten log of it (this is the same as with pH, the negative log of $[H^+]$). Since a large K_a means a small pK_a, this means that **stronger acids have lower pK_a values**. Since stronger acids have weaker conjugate bases, this also means that **weaker bases have lower pK_a values**. You can think of pK_a as a rating of how badly a compound wants to hang on to its proton or get its proton back – a large number means it really wants to keep its proton, which makes it a very weak acid and (if it does get deprotonated) a very strong base. Reactions will always favor making the weaker acid and base over the stronger acid and base. The textbook has a table of pK_a values – you do not need to memorize these, but it will be helpful in future to get a feel for where compounds rank relative to each other. (The exact numbers will also vary from book to book.) Here are some examples:

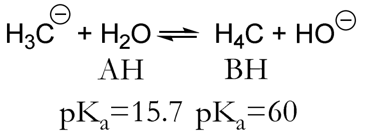
pKa	Acid	Conj. Base
-10	HI	I^-
-1.7	H_3O^+	H_2O
9.3	NH_4^+	NH_3
15.7	H_2O	HO^-

pK_a tables normally only show the acid form explicitly, but it's important to remember that they're describing the balance between the acid and its conjugate base. In many reactions, the strength of the conjugate base is the determining factor in outcome, so you need to mentally convert the compounds in the table to their conjugate bases before comparing them to your compound. (Note: The only pK_a values that can be directly measured in comparison to water are the ones that lie between the pK_a values for water acting as an acid or base (-1.7 to 15.7). Anything outside of this range is harder to measure directly, but it can be measured in relation to an acid with a known pK_a in a solvent other than water. pK_a values can shift by several units depending on conditions, but the relative rankings tend to stay the same.)

pK_a values can be used to predict the equilibrium in acid-base reactions. The general way to do this is with this equation:

$$AH + B^{\ominus} \rightleftharpoons A^{\ominus} + BH \qquad K_{eq} = 10^{(pKa_{BH} - pKa_{AH})}$$

For example, we can calculate that in a reaction between CH_3^- and H_2O, the equilibrium will massively favor the products.

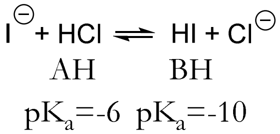


$$\underset{\substack{\text{AH}\\ pK_a=15.7}}{H_3C^{\ominus} + H_2O} \rightleftharpoons \underset{\substack{\text{BH}\\ pK_a=60}}{H_4C + HO^{\ominus}}$$

$K_{eq} = 10^{(60 - 15.7)} = 10^{44.3}$ = huge!

Equilibrium favors products exclusively

In the reaction between I^- and HCl, the reactants are favored instead.



$$\underset{\substack{\text{AH}\\ pK_a=-6}}{I^{\ominus} + HCl} \rightleftharpoons \underset{\substack{\text{BH}\\ pK_a=-10}}{HI + Cl^{\ominus}}$$

$K_{eq} = 10^{(-10 - -6)} = 10^{-4}$ = small

Equilibrium favors reactants

For a reaction to be considered "quantitative" or ~100% complete, the K_{eq} has to be at least 10^2. This means that less than 1% of the reactants are left.

The K_{eq} for a reaction can also be used to find the energy change for the reaction. This is given in terms of Gibbs free energy, or $\Delta G°$. Since it's an exponential equation, a small change in $\Delta G°$ makes a very large change to the K_{eq} for the reaction.

$$K_{eq} = 10^{-\Delta G/(2.3RT)}$$

Relationships between Structure and Acidity

All of these are based on the same principles: **More stable negative charge on conjugate base = weaker conjugate base = stronger conjugate acid = lower pKa.** Charges tend to be more stable when they're at least partially cancelled out by opposite charges, or if they're spread out over a greater area. Even if we do not know the pK_a for a compound, we can estimate it based on several trends. You can remember these as **CARDIO**: **C**harge, **A**tom, **R**esonance, **D**ipole **I**nduction, and **O**rbitals.

1. **Charge** effect: Positive charge on something makes it more acidic and negative charge makes it less acidic. This is because an acid becomes more negative after losing a proton, so the more positive it is to begin with, the more stabilization it will gain. As an example, H_3O^+ has a pK_a of -1.7, while H_2O has a pK_a of 15.7.
2. **Atom** effect (or element effect in Loudon): if you look at the atom that is losing a proton and its place in the periodic table, it will be more acidic as you move left to right and top to bottom in the table.

increasing acidity →

H–CH$_3$	H–NH$_2$	H–OH	H–F
60	35	15.7	3.2
		H–SH	H–Cl
		7	-3
			H–Br
			-5
			H–I
			-9

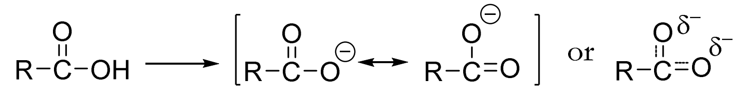
increasing acidity ↓

If we assumed base strength was determined solely by electronegativity, then the left-to-right trend makes sense (F is more stable with a negative charge than C, so it's more willing to act like an acid by giving up a proton). But the top-to-bottom trend does not make sense – F is more electronegative than I, but HI is more acidic. This is because the process of an acid dissociating from a proton can be broken into 3 hypothetical steps:

1) $\text{H-A} \longrightarrow \text{H}^{\bullet} + \text{A}^{\bullet}$ Energy depends on H-A bond strength
2) $\text{H}^{\bullet} \longrightarrow \text{H}^{\oplus} + e^{\ominus}$ (constant for all acids)
3) $\text{A}^{\bullet} + e^{\ominus} \longrightarrow \text{A}^{\ominus}$ Energy depends on electronegativity of A

The reason HI is such a strong acid is because even though I is not that electronegative, the H-I bond is so weak that breaking it costs very little energy. This effect is larger than the effect caused by the difference in electronegativities. So vertical differences within the periodic table are caused mostly by different costs in step 1, while horizontal differences are caused mostly by differences in step 3.

3. **Resonance** effect: This is based on conjugate bases having more than one valid Lewis dot structure, and so more than one valid location for the charge. Since the true molecule is a weighted average of its resonance forms, this means that the charge is spread out over more than one atom. We'll use carboxylic acids as examples. When these lose a proton, they take on a negative charge that can be distributed out through resonance. This makes the charge more stable, which means that they're more acidic than the same molecule without resonance (for instance, a deprotonated alcohol).



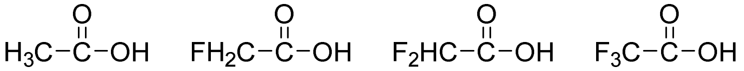
pKa = 0-5

$\text{R}-\text{CH}_2-\text{OH} \longrightarrow \text{R}-\text{CH}_2-\text{O}^{\ominus}$

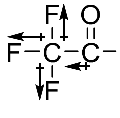
pKa = 16-18

No resonance to stabilize charge = less stable with charge = stronger base/weaker acid = higher pKa

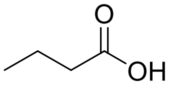
4. **Dipole Induction** effect (or polar/inductive effect in Loudon): This is caused by polar bonds further away from the acidic site (the atom that loses a proton). We can look at how different R groups will change the pK_a of a carboxylic acid. We'll start with acetic acid, where R is just a CH_3 group, and replace more and more of its Hs with Fs to see how the pK_a changes.

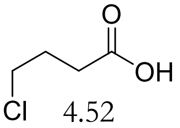


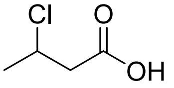
4.76 2.66 1.24 0.23



As it turns out, having more F atoms drops the pK_a by several units. This effect is due to the F-C polar bonds pulling some electron density down the chain away from the carboxylic group, which reduces and stabilizes the negative charge on the conjugate base form. As another example, we can look at the change in pK_a values when the same group is moved closer to the acidic site.


4.82


4.52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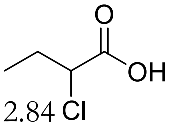
2.84


Again, this is because the dipole is being pulled along the chain towards the Cl, but the effect gets weaker the further the Cl gets from the carboxylic acid. F and Cl are two examples of groups that are inductively electron-withdrawing. It is also possible to have groups that are inductively electron-donating, so they destabilize the negative charge and make the compound less acidic instead of more. (Note that separating resonance and induction is sometimes difficult – for example in the two molecules shown under the "Resonance effect" heading, you could also make the argument that the second oxygen on carboxylic acids helps to stabilize the charge by induction. These effects often work together but not always.)

5. **Orbital** effect: This is based on the hybridization of the atom bearing the negative charge. We already know that hybrid orbitals are averaged-together from s and p orbitals – for example, an sp^3 hybrid is 25% s and 75% p. The more s character an orbital has, the more stable it will be when holding a negative charge, since the s orbitals are closer in to the nucleus. For this reason, a negative charge on sp-hybridized carbon is much more stable than on sp^3-hybridized carbon.

$R{-}C{\equiv}C{-}H$

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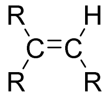

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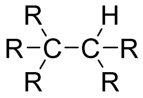

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