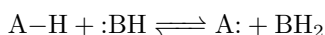


Topic IV

Hard-Soft Acid-Base and Donor-Acceptor Concepts of Transition Metals

IV.1 Module 24: Acid-Base Chemistry

- 2/8: • Brønsted-Lowry Acid-Base Theory of Acids and Bases (1923).
- Acid: Any chemical species (molecule or ion) that is able to lose, or “donate,” a hydrogen ion (proton).
 - Base: Any chemical species that is able to gain, or “accept”, a proton.
 - A base must have a pair of electrons available to share with the proton; this is usually present as an unshared pair, but sometimes is in a π orbital.
 - Acid-base reactions: The transfer of a proton from an acid to a base.



- Protons do not exist free in solution but must be attached to an electron pair.
- Water is amphoteric.
- In the Brønsted-Lowry paradigm, we cannot separate the acids/bases from the solvent (no protons; only *solvated* protons). In a non-aqueous medium such as DMSO, however, we have much broader scope of acids and bases.
- **Carbon acid:** Any molecule containing a C–H bond can lose a proton forming the carbanion.
- Carborane ($\text{H}(\text{CHB}_{10}\text{Cl}_{11})$) is a superacid one million times stronger than sulfuric acid since its conjugate basis is incredibly stable (super easy to delocalize the charge).
- The base dissociation constant or K_b is a measure of basicity. $\text{p}K_b$ is the negative log of K_b and related to the $\text{p}K_a$ by the simple relationship $\text{p}K_a + \text{p}K_b = 14$. The larger the $\text{p}K_b$, the more basic the compound.
- **Superacid:** An acid with acidity greater than that of 100% pure sulfuric acid.
 - In water, the strongest acid you can have is H_3O^+ .
 - The strongest superacids are prepared by the combination of two components, a strong Lewis acid and a strong Brønsted-Lowry acid.
 - Fluoroantimonic acid HF-SbF_5 is 2×10^{19} stronger than 100% sulfuric acid.

- Olah's magic acid ($\text{FSO}_3\text{H}-\text{SbF}_5$) can dissolve paraffin (candle wax; extremely inert), converting methane into the t-butyl carbocation.
- **Hammett acidity function:** Can replace the pH in concentrated solutions. *Also known as H_0 .*

$$H_0 = \text{p}K_{\text{BH}^+} + \log \frac{[\text{B}]}{[\text{BH}^+]}$$
 - Let BH^+ be the conjugate acid of a very weak base B, with a very negative $\text{p}K_{\text{BH}^+}$. In this way, it is rather as if the pH scale has been extended to very negative values.
 - Hammett originally used a series of anilines with EWGs for the bases.
- **Superbase:** A compound that has a high affinity for protons.
 - Again, these do not exist in water.
 - Often destroyed by water, CO_2 , and O_2 .
 - A superbase has been defined as an organic compound whose basicity is greater than that of proton sponge, which has conjugate $\text{p}K_{\text{a}}$ of 12.1.
 - These are valuable in organic chemistry, which abounds in very weak acids.
 - A common superbase is lithium diisopropylamide.

IV.2 Module 25: Acid-Base Chemistry: Lewis Acids and Bases

2/10:

- The second midterm will be Saturday, February 20.
- Talapin will hold office hours this Friday from 3:00-4:00 PM.
- **Lewis acid:** Any species with a vacant orbital.
 - An atomic or molecular species that has an empty atomic or molecular orbital of low energy (LUMO) that can accommodate a pair of electrons.
- **Lewis base:** A compound with an available pair of electrons, either unshared or in a π orbital.
 - An atomic or molecular species that has a lone pair of electrons in the HOMO.
- Examples of Lewis acids:
 - The proton, onium ions (e.g., NH_4^+ , H_3O^+), metal cations (e.g., Li^+ , Mg^{2+}), trigonal planar species (e.g., BF_3 , CH_3^+), and electron-poor π -systems (e.g., enones, tetracyanoethylene [TCNE]).
- **Lewis base adduct:** The product of a Lewis acid reaction.
 - Simple Lewis acids:
 - Examples include BF_4^- in $\text{BF}_3 + \text{F}^- \longrightarrow \text{BF}_4^-$ and BF_3OMe_2 in $\text{BF}_3 + \text{OMe}_2 \longrightarrow \text{BF}_3\text{OMe}_2$, which are Lewis base adducts of BF_3 .
 - In many cases, Lewis base adducts violate the octet rule.
 - In some cases, Lewis acids can bind to two Lewis bases (e.g., $\text{SiF}_4 + 2\text{F}^- \longrightarrow \text{SiF}_6^{2-}$).
 - Complex Lewis acids:
 - Most compounds considered to be Lewis acids require an activation step prior to the formation of the adduct.
 - For example, the reaction $\text{B}_2\text{H}_6 + 2\text{H}^- \longrightarrow 2\text{BH}_4^-$ goes through the intermediate B_2H_7^- .
- The proton (H^+) is one of the strongest but is also one of the most complicated Lewis acids. It is convention to ignore the fact that a proton is heavily solvated (bound to solvent). With this simplification in mind, acid-base reactions can be viewed as the formation of adducts.

- For example, $\text{H}^+ + \text{NH}_3 \longrightarrow \text{NH}_4^+$ and $\text{H}^+ + \text{OH}^- \longrightarrow \text{H}_2\text{O}$.
- In the first reaction, essentially what happens is the frontier orbitals of H^+ and NH_3 are very similar in energy, and thus strongly combine when they form bonding and antibonding orbitals. Two electrons previously in the HOMO of NH_3 drop the bonding energy into the new lower MO of NH_4^+ . More rigorously, we have to account for the change in symmetry group from C_{3v} to T_d , but the above description basically encapsulates what happens.
- Examples of Lewis bases:
 - $\text{NH}_{3-x}\text{R}_x$ where R = alkyl or aryl.
 - $\text{PR}_{3-x}\text{A}_x$ where R = alkyl and A = aryl.
 - Compounds of O, S, Se, and Te in oxidation state 2, including water, ethers, and ketones.
 - Simple anions (e.g., H^- , F^-), other lone pair-containing species (e.g., H_2O , NH_3 , OH^- , CH_3^-), complex anions (e.g., sulfate), and electron-rich π -systems (e.g., ethyne, ethene, benzene).
- To quantify the strength of Lewis acids and bases, compare the standard enthalpies of complexation in kJ/mol.
 - For example, heats of binding of Lewis bases to BF_3 .
 - We do this because we cannot use $\text{p}K_b$ in the same way as we can in the Brønsted-Lowry description.
- **Electrides:** Compounds where the anions are electrons.
 - Put an alkali metal in liquid ammonia; the deep blue color comes from solvated electrons floating freely in solution surrounded by ammonia molecules.
 - If you add a very strong complexing agent for alkali metal ions to the solution, then the electrons will not be able to react back with the alkali metal and we will have crystallized an electride.
 - Electrides are dielectrics.
- The power of the Lewis approach is that it allows us to classify many reactions as acid-base reactions.
 - For example, Wilkinson's catalyst ($\text{Rh}(\text{PPh}_3)\text{Cl}$ where Ph is a phenyl group) is probably the first hydrogenation catalyst to be explained.
 - The Lewis approach allows us to view every step in the mechanism^[1] as an acid-base reaction.

IV.3 Module 26: Hard and Soft Acids and Bases (HSAB) Principle

- HSAB principle can't be derived, but it's been proven to be a very powerful tool for chemists.
- The affinity of hard acids and hard bases for each other is mainly ionic in nature, whereas the affinity of soft acids and bases for each other is mainly covalent in nature.
- Hard acids and hard bases tend to have:
 - Small atomic/ionic radius.
 - High oxidation state.
 - Low polarizability.
 - High electronegativity.
 - Energy low-lying HOMO (bases) or energy high-lying LUMO (acids).

¹We will study this mechanism in depth in CHEM 20200.

- Examples: H^+ , alkali ions, OH^- , F^- .
- Small stabilization energy when bonding, so attraction is mostly electrostatic.
- Soft acids and soft bases tend to have:
 - Large atomic/ionic radius.
 - Low or zero oxidation state.
 - High polarizability.
 - Low electronegativity.
 - Energy high-lying HOMO (bases) or energy low-lying LUMO (acids).
 - Examples: CCH_3Hg^+ , Pt^{2+} , H^- , R_3P , I^- .
 - Large stabilization energy when bonding, so attraction is mostly covalent.
- Hard acids don't readily react with soft bases and vice versa.
- You should be able to intuitively classify acids/bases as hard or soft.

| | | | | | | | | | | | | | | | | | | | | | | |
|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|--|--|----------|----------|--|--|--|
| 1 H | | | | | | | | | | | | | | | | | <div><div></div><div></div><div></div></div> <div>hard soft intermediate</div> | | | | | |
| 3 Li | 4 Be | | | | | | | | | | | | | | | | | | | | | |
| 11 Na | 12 Mg | | | | | | | | | | | | | | | | | 13 Al | 14 Si | | | |
| 19 K | 20 Ca | 21 Sc | 22 Ti | 23 V | 24 Cr | 25 Mn | 26 Fe | 27 Co | 28 Ni | 29 Cu | 30 Zn | 31 Ga | 32 Ge | 33 As | | | | | | | | |
| 37 Rb | 38 Sr | 39 Y | 40 Zr | 41 Nb | 42 Mo | 43 Te | 44 Ru | 45 Rh | 46 Pd | 47 Ag | 48 Cd | 49 In | 50 Sn | 51 Sb | | | | | | | | |
| 55 Cs | 56 Ba | | | 72 Hf | 73 Ta | 74 W | 75 Re | 76 Os | 77 Ir | 78 Pt | 79 Au | 80 Hg | 81 Ti | 82 Pb | 83 Bi | | | | | | | |
| 87 Fr | 88 Ra | | | | | | | | | | | | | | | | | | | | | |

| | | | | | | | | | | | | | | |
|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|
| 57 La | 58 Ce | 59 Pr | 60 Nd | 61 Pm | 62 Sm | 63 Eu | 64 Gd | 65 Tb | 66 Dy | 67 Ho | 68 Er | 69 Tm | 70 Yb | 71 Lu |
|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|----------|

| | | | | | | | | |
|----------|----------|----------|---------|----------|----------|----------|----------|----------|
| 89 Ac | 90 Th | 91 Pa | 92 U | 93 Np | 94 Pu | 95 Am | 96 Cm | 97 Bk |
|----------|----------|----------|---------|----------|----------|----------|----------|----------|

Figure IV.1: Hard vs. soft metal ions.

- The affinity that metal ions have for ligands is controlled by size, charge, and electronegativity.
- This can be refined further by noting that for some metal ions, their chemistry is dominated by size and charge, while for others it is dominated by their electronegativity. These two categories of metal ions have been termed by Pearson as **hard** (metal ions) and **soft** (metal ions).
- Soft Lewis acids are transition metals with unusually high electronegativities (these go against typical periodic trends).
 - Note that QMech can't explain this island of high electronegativities in the transition metals; we need relativistic corrections and Dirac equations.

- The higher electronegativity stabilizes the LUMO of the acid.
- These ions are a treasure trove for catalytic chemistry.
- **Chemical hardness:** One half the difference between the ionization potential I and the electron affinity A .

$$\eta = \frac{I - A}{2}$$

- Note that the above equation cannot be applied to anions because electron affinity cannot be measured for them; the assumption is made that η for an anion X^- is same as that for the radical X^\bullet .
- Related to the **Mulliken electronegativity**:

$$\chi = \frac{I + A}{2}$$

- **HSAB principle:**

- (i) Hard acids prefer to bond to hard bases, and soft acids prefer to bond to soft bases.
 - Note that this rule has nothing to do with acid or base strength but merely says that the product $A-B$ will have extra stability if A and B are both hard or both soft.
- (ii) A soft Lewis acid and a soft Lewis base tend to form a covalent bond, while a hard acid and a hard base tend to form ionic bonds.
- **Solubility:** Hard solvents (e.g., HF, H₂O, and the protic solvents) tend to solvate strong solute bases (e.g., F[−] and the oxygen anions). Dipolar aprotic solvents (e.g., Me₂SO and CH₃COCH₃) are soft solvents with a preference for solvating large anions and soft bases.
 - For example, $LiI + AgF \longrightarrow LiF + AgI$. In this reaction, two mixed reactants recombine exothermically to form a hard-hard product and a soft-soft precipitate. Because of the Lewis definition, we can treat this as an acid-base reaction.

| Classification | | Hard | Intermediate | Soft | |
|----------------|------------------|----------------|-----------------|-----------------|----------------|
| | | F [−] | Cl [−] | Br [−] | I [−] |
| Soft | Ag ⁺ | 0.4 | 3.3 | 4.7 | 6.6 |
| Intermediate | Pb ²⁺ | 1.3 | 0.9 | 1.1 | 1.3 |
| Hard | Fe ³⁺ | 6.0 | 1.4 | 0.5 | - |

Table IV.1: Hard and soft formation constants.

- As another example, we can see in Table IV.1 that soft-soft and hard-hard ions have higher formation constants ($\log K_1$) than any other combination, and hard-soft ions have lower formation constants than any other combination.
- **Coordination chemistry:** Numerous experiments have been done to determine the relative ordering of ligands and transition metal ions in terms of their hardness and softness.
- Au (I) is the softest metal ion.
 - Soft to the extent that compounds such as AuF and Au₂O are unknown.
 - Nevertheless, it forms stable compounds with soft ligands such as cyanide.
- Al (III) is a very hard metal ion.
 - It has very high formation constants with both F[−] and OH[−]; additionally, it has virtually no affinity in solution for heavier halides such as Cl[−].

- Its solution chemistry is dominated by its affinity for F^- and for ligands with negative O-donors.
- **Symbiosis effect:** Taking a borderline element and reacting it with a soft Lewis base to soften it, or a hard Lewis base to harden it.
 - For example, B^{3+} can be reacted with $3H^-$, we make the soft compound BH_3 . If we react it with $3F^-$, we make the hard compound BF_3 .
 - Now BH_3 and F^- do not react, but we can further soften it in the reaction $BH_3 + H^- \longrightarrow BH_4^-$.
 - Similarly, BF_3 and H^- don't react, but we can further harden BF_3 by reacting it with F^- to form BF_4^- .
- There is a preference against combining hard and soft ligands in the first coordination sphere. Examples:
 - $CH_3F + CF_3I \rightleftharpoons CH_3I + CF_4$ since CH_3^+ and I^- are soft while CF_3^+ and F^- are hard.
 - $Co(NH_3)_5F$ is stable since it has all hard ligands, but $Co(NH_3)_5I$ will react with H_2O to produce $Co(NH_3)_5OH$ since the former combines hard and soft and the latter does not.
 - Thiocyanate is **ambidentate**. Thus, we have Fe–N bonds in $Fe(NCS)_6^{3+}$ (with the hard Fe (III) ion) and Au–S bonds in $Au(SCN)_2^-$ (with the soft Au (I) ion).
 - Note that intermediate metal ions tend to bond to thiocyanate through nitrogen; indeed, intermediate Cu (II) forms $Cu(NCS)_4^{2-}$.
- **Ambidentate** (ligand): A polyatomic ligand that can bind through more than one of its constituent atoms.

IV.4 Module 27: Steric Effects in Inorganic Chemistry

2/12:

- Chemistry is a game of two players: Electronic and steric effects.
- Gallanes, gallenes, cyclogallenes, and gallynes:
 - Various compounds with gallium.
 - React gallium (III) chloride with a Grignard reagent containing mesityl (2,4,6-trimethylphenyl; Mes; $Me_3C_6H_2$): $3(Me_3C_6H_2)MgBr + GaCl_3 \xrightarrow{-3MgBrCl} Ga(Me_3C_6H_2)_3$.
 - A molecule with even more steric hindrance: a three gallium ring with a 2,6-dimesitylphenyl groups attached to each gallium.
 - The triphenylcyclopropenium cation is aromatic but the cyclogallene dianion is metalloaromatic.
 - In another compound, a $Ga \equiv Ga$ triple bond holds two massive groups together.
 - The debate over this is a good example of chemistry not being black and white.
 - Acetylene and gallyne both show band structure.
 - Additionally, the former is linear while the latter is trans-bent with a donor-acceptor bond lacking a sigma bond.
 - To learn more about such bonding, read Power (2010).
- Steric effects are much more common in Lewis acid-base reactions in which larger acids are used.
 - With a small acid, increasing size in the base (i.e., replacing hydrogens with methyl groups) correlates with increasing strength.
 - With a large acid, decreasing size in the base correlates with increasing strength.
- Frustrated Lewis pairs:
 - Excessive steric bulk can prevent acids and bases from reacting to form an adduct.

- However, you can also have alternate reactions that yield compounds with Lewis acidic *and* Lewis basic regions.
- Frustrated Lewis pairs can activate (cleave) H_2 .
 - They can also activate CO_2 ; this has applications in the realm of CO_2 sequestration.

IV.5 Chapter 6: Acid-Base and Donor-Acceptor Chemistry

From Miessler et al. (2014).

- IUPAC calls hydronium “oxonium” and uses “hydrogen ion” instead of “proton.”
- In the Brønsted-Lowry paradigm, “the equilibrium always favors the formation of *weaker* acids and bases” (Miessler et al., 2014, p. 171).
- Common amphoteric solvents: sulfuric acid, hydrofluoric acid, water, acetic acid, methanol, ammonia, and acetonitrile.
- H_3O^+ and OH^- are the strongest acid and base, respectively, that^[2] can exist in H_2O .
- **Leveling effect:** Acids stronger than H_3O^+ cannot be differentiated by their aqueous ionization.
- Because of leveling, the strength of the strong acids cannot be differentiated in aqueous solution.
 - Thus, we use more strongly acidic solvents, such as **glacial** acetic acid.
 - The use of this solvent allows us to determine that $\text{HClO}_4 > \text{HCl} > \text{H}_2\text{SO}_4 > \text{HNO}_3$.
 - Note that basic solvents similarly permit the differentiation of strong bases.
- **Glacial** (substance): A 100% pure, concentrated sample of a substance.
- “Nonamphoteric solvents...do not limit solute acidity or basicity because the solute does not react with the solvent. In these solvents, the inherent solute acid or base strength determines the reactivity, without a leveling effect” (Miessler et al., 2014, p. 173).
- **Superelectrophilic activation:** The result of generating small organic ions bearing a large amount of positive charge.
- Introduces the Hammett acidity function^[3].
- Protonating CH_4 is of particular interest since the natural abundance of methane (from natural gas) makes it attractive as a starting point for synthesis of more complex molecules.
 - CH_5^+ has been isolated; CH_6^{2+} and CH_7^{3+} have been proposed.
- Dissolving SO_3 in H_2SO_4 results in **fuming sulfuric acid**, which contains higher polysulfuric acids such as $\text{H}_2\text{S}_2\text{O}_7$, all of which are stronger than H_2SO_4 .
- Water is a strong base in superacid media: Adding H_2O to superacids produces hydronium salts.
 - As a consequence, we can form $[\text{H}_3\text{O}][\text{Ln}][\text{AsF}_6]_3$ where Ln is a lanthanide element and AsF_6 is part of a superacid.
- Thermodynamic measurements:
 - Hess’s law is often used to express the enthalpy of weak acid reactions in terms of reactions that do go to completion; this is not perfect, but it’s an ok starting point.

²Errata: Miessler et al. (2014, p. 173) has a typo: “than” instead of “that.”

³Errata: The – sign should be a + sign.

- One can also measure K_a via titration curves at different temperatures and use the van't Hoff equation

$$\ln K_a = \frac{-\Delta H}{R} \frac{1}{T} + \frac{\Delta S}{R}$$

- According to the above equation the plot of $\ln K_a$ vs. $\frac{1}{T}$ will be linear with a slope that allows us to determine ΔH and a y -intercept that allows us to determine ΔS .
- Brønsted basicity scale: Enthalpy changes associated with protonation by the superacid fluoro-sulfonic acid (HSO_3F).
- The best measure of acid/base strength is gas-phase acidity/basicity, since there are no solvent effects.
 - Wrt. $\text{HA}(\text{g}) \longrightarrow \text{A}^-(\text{g}) + \text{H}^+(\text{g})$, ΔG = Gas-Phase Acidity (GA) and ΔH = Proton Affinity (PA).
 - There exist analogous definitions for bases.
 - Modern measurement techniques can measure these values very accurately for a select few molecules; from these molecules and Hess's law, we can build pretty good approximations of other reactions.
- **Superbase**: A base with a gas-phase proton affinity greater than 1000 kJ/mol.
 - Examples such as Grignard and organolithium reagents are ubiquitous in organic synthesis.
- Considers inductive effects.
- "Solvation [of amines] is dependent on the number of hydrogen atoms available to form $\text{O} \cdots \text{H}-\text{N}$ hydrogen bonds with water" (Miessler et al., 2014, p. 181).
- Steric effects are less obvious — for example, 2,6-dimethylpyridine is more basic than 4-methylpyridine, but the latter is more basic than 2-*t*-butylpyridine (see Figure IV.2).
- Binary hydrogen compounds:
 - Acidity increases left-to-right across a period and down a group (wrt. the non-hydrogen atom).
 - The three heaviest hydrohalic acids are all equally acidic in water due to leveling.
- Considers relative strengths of various chlorine-based oxyacids (e.g., HClO_4 vs. HClO_3).
 - As the number of oxygens increases, the electronegativity of the terminal oxygen increases. Less and less electron density binds the hydrogen, making the bond increasingly weak and susceptible to heterolytic cleavage.
- In polyprotic acids, the $\text{p}K_a$ increases about 5 units with each successive proton removal.
- Aqueous cations get surrounded by H_2O .
 - Then these complexes react in acid-base reactions.
 - When the concentration is sufficiently high, two complexes can combine via hydroxide or oxide bridges between the metal atoms.
 - Metal ions with charges of 4+ or higher are the really acidic ones. They're acidic to the extent that they do not exist by themselves in solution, but are present in oxygenated forms such as permanganate (MnO_4^-).
- **Coordinate covalent** (bond): A bond that links a Lewis acid and base into their adduct. *Also known as dative* (bond).
- **Coordination compound**: A Lewis acid-base adduct involving one or more metal ions.
- "In most Lewis acid-base reactions, the HOMO-LUMO combination forms new HOMO and LUMO orbitals of the product" (Miessler et al., 2014, p. 186).

- When the symmetries match and the energies are close, a stable adduct forms.
- Since every molecule has a HOMO and LUMO, technically any molecule can act as acid, base, oxidizing agent, or reducing agent when combined with the right reactants.
 - For example, H_2O can oxidize Ca metal, forming $\text{Ca}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) + \text{H}_2(\text{g})$. H_2O^{-} is not formed because the addition of electrons to antibonding orbitals in H_2O weakens one O–H bond.

2/14:

- Section on spectroscopic support for frontier orbital interactions.
 - Very confusing; electron transitions' colors in I_2 when it acts as a Lewis acid.
- **Lewis basicity**: The thermodynamic tendency of a substance to act as a Lewis base.
 - It would be ideal to measure in the gas phase, but this is hard to do.
 - Determining a good reference acid is difficult.
 - When in solution, we want to make sure that we choose a solvent that will not react significantly as a Lewis acid with the solutes.
 - In other words, we need a solvent that primarily solvates the solutes through dispersion forces (i.e., nonpolar solvents).
 - Once we have a solvent, we quantify the Lewis basicity of a substance by finding the K_b and $\log K_b$ values for its complexation with a reference acid (such as I_2). The absolute basicity will vary based on the solvent, but the Lewis basicity ranking will be the same overall.
- The standard scale for Lewis basicity is its BF_3 affinity, or $-\Delta H^\circ$ in the following reaction, corrected for the enthalpy of BF_3 dissolving in the solvent.



- **Halogen (bond)**: A coordinate covalent bond formed by a halogen X_2 (e.g., I_2) or interhalogen XY (e.g., ICl) to a Lewis base.
 - Exhibit approximately 180° angles about the halogen donor atom, supporting use of σ^* LUMO.
- I_2 affinity values are commonly determined in heptane, and while there is interest in creating an I_2 affinity scale analogous to the BF_3 one, as of yet, various experimental designs have not been reconciled.
- Since the formation of a halogen bond involves donation into the halogen σ^* LUMO, the X–Y bond weakens and lengthens. This also decreases the stretching frequency. Measuring changes in stretching frequency induced in I_2 , ICN , and ICl via Raman spectroscopy has correlated reasonably well to basicity, but not perfectly since these stretching bands are affected by the presence of the Lewis base, too.
- Inductive effects:
 - Adding more electronegative substituents decreases Lewis basicity (in C_{3v} Lewis bases).
 - Adding more alkyl substituents increases Lewis basicity (in C_{3v} Lewis bases).
 - Adding more electronegative substituents increases Lewis basicity (in D_{3h} Lewis bases) since the increase in bond lengths draws π -bonding electrons farther from the central boron.
- Steric effects on Lewis acidity and basicity: **front strain**, **back strain**, and **internal strain**.
- **Front strain**: Bulky groups interfering directly with the approach of an acid and a base to each other. *Also known as F strain.*
- **Back strain**: Bulky groups interfering with each other when VSEPR effects force them to bend away from the other molecule forming the adduct. *Also known as B strain.*

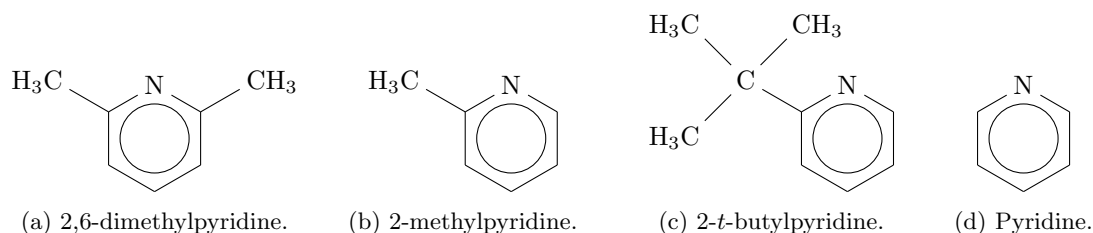


Figure IV.2: Four Lewis bases.

- **Internal strain:** Effects from electronic differences within similar molecules. *Also known as I strain.*
- An example of differences caused by F strain:
 - When reacting with H^+ , 2,6-dimethylpyridine > 2-methylpyridine > 2-*t*-butylpyridine > pyridine.
 - When reacting with BF_3 or BMe_3 , pyridine > 2-methylpyridine > 2,6-dimethylpyridine > 2-*t*-butylpyridine.
- An example of B strain: The fact that tri(*t*-butyl)boron vs. H^+ elicit a reversed order of the basicity of NH_3 , MeNH_2 , Me_2NH , and Me_3N .
- **Frustrated Lewis pair:** A lone pair on a Lewis base that does not form a traditional adduct with a Lewis acid because both molecules are hindered by excessive steric bulk. *Also known as FLP.*
- Example (refer to Figure IV.3 throughout the following discussion):

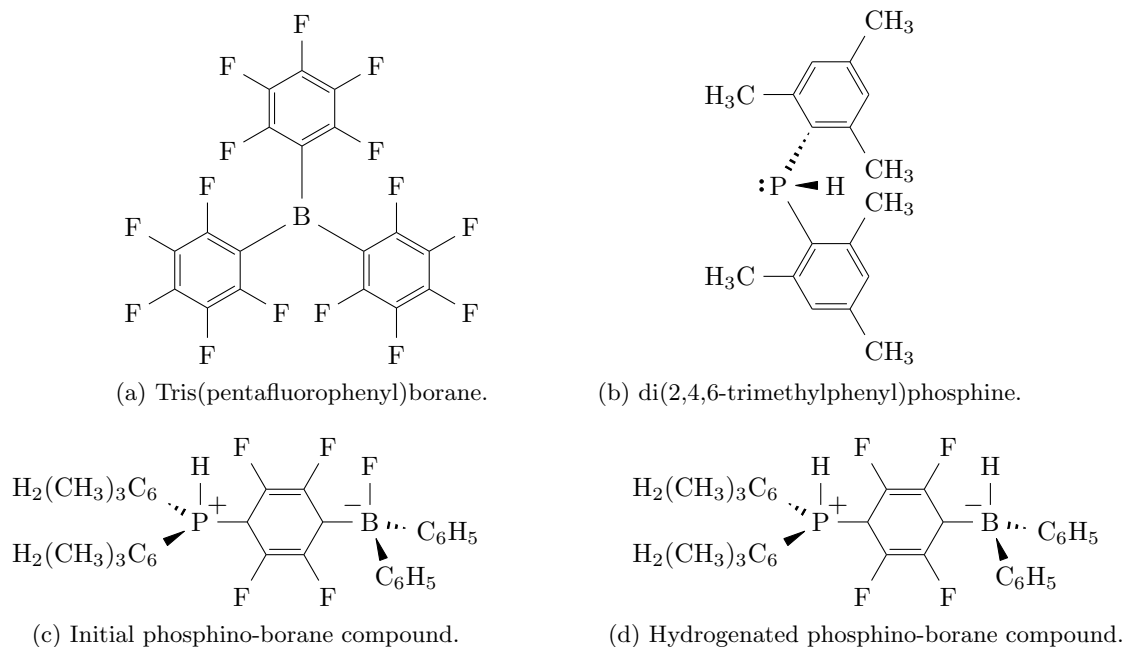


Figure IV.3: FLP phosphino-borane compound.

- Although the FLP in di(2,4,6-trimethylphenyl)phosphine does not attack the boron in tris(pentafluorophenyl)borane, it reacts with a para carbon of the borane to create a **zwitterionic** species after fluoride migration.
- This fluoride can be substituted for a hydrogen, creating a species that releases hydrogen gas upon heating and reacts with hydrogen gas at ambient temperature to reform the zwitterion.

- The phosphino-borane is the first non-transition metal species that can reversibly activate the H–H bond in H₂.
- **Zwitterion:** A species that contains at least one formal positive and negative charge.
- More small molecules have been activated with FLPs, including CO₂ and N₂O.
- **Hydrogen (bond):** An X–H···B bond formed from an attraction between an X–H unit (where the electronegativity of X is greater than that of H) and a donor atom B.
 - The components may be incorporated into larger molecular fragments.
 - Hydrogen bonds can be either intermolecular or intramolecular.
- Hydrogen bonds can be described on the basis of the varying relative contributions from three components:
 - Electrostatic contribution: The polarity of X–H.
 - Partial covalent character and charge transfer: The donor-acceptor nature of the interaction.
 - Dispersion forces.
- Experimental stipulations that must be satisfied for an interaction to be declared a hydrogen bond.
 - Bond angle close to 180°.
 - Red-shifted IR stretching frequency for the X–H bond (indicative of the bond weakening and, thus, donation into antibonding orbitals).
 - High deshielding of NMR chemical shift (extremely sensitive probe for hydrogen bonding).
 - Magnitude of ΔG for bond formation must exceed the thermal energy of the system.
- **H-bond puzzle:** Predicting the strength of hydrogen bonds solely on the basis of the structures of the participating molecules.
 - Based in the fact that structurally very similar hydrogen bonds can exhibit massive differences in strength.
 - For example, the O–H···O bond between hydronium and water is roughly six times stronger than the comparable bond between two water molecules.
- **pK_a equalization** (paradigm): An approach to predicting the strength of hydrogen bonds in aqueous solution by envisioning H-bond strength as associated with the 3-way proton transfer equilibrium between X and B.
 - The better matched the pK_a values, the stronger the predicted H-bond.
 - In effect, a strong H-bond minimizes $\Delta pK_a(X-H \cdots B) = pK_a(HX) - pK_a(BH^+)$.
- **Receptor-guest interaction:** An interaction between molecules with extended pi systems where their pi systems interact with each other to hold the molecules or portions of the molecules together. *Also known as receptor-substrate interaction, host-guest interaction.*
 - One example is a double-concave hydrocarbon buckycatcher that forms a ball-and-socket structure with C₆₀. This is an example of an **inclusion complex**.
- The hard vs. soft distinction is largely a result of polarizability, with soft acid having high polarizability and hard acids having low polarizability.
- Soft transition metals have *d* electrons available for π -bonding.
- **Exchange reaction:** A reaction involving the exchange of a water and a base B.

- Is OH^- or S^{2-} more likely to form insoluble salts with 3+ transition-metal ions? Which is more likely to form insoluble salts with 2+ transition-metal ions?
 - OH^- and 3+ are hard, and thus will form insoluble salts.
 - S^{2-} and 2+ are soft, and thus will form insoluble salts.
- Both hardness/softness and strength must be considered to determine reactivity.
- MO diagrams for hard-hard and soft-soft interactions reveal that there is a lower covalent contribution in hard-hard interactions, but this is compensated for by a strong ionic contribution, so don't think that hard-hard interactions are weaker.
- Wrt. chemical hardness η (Greek "eta"), ionization energy is assumed to measure the energy of the HOMO ($E_{\text{HOMO}} = -I$) and electron affinity the energy of the LUMO ($E_{\text{LUMO}} = -A$); thus, a hard species is one with a high difference in energy between its HOMO and LUMO.
- **Chemical softness:** The inverse $\sigma = \frac{1}{\eta}$ of hardness.
- Reviews a quantitative system of acid-base parameters to account for reactivity by explicitly including electrostatic and covalent factors.