

Organometallic Chemistry and Homogeneous Catalysis

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Lecture N2

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Properties

I. Nature of M

II. Nature of L

II.a. Ligands with C-atom attached to M

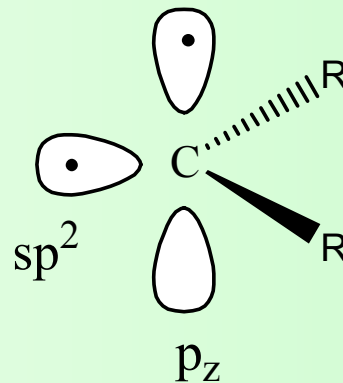
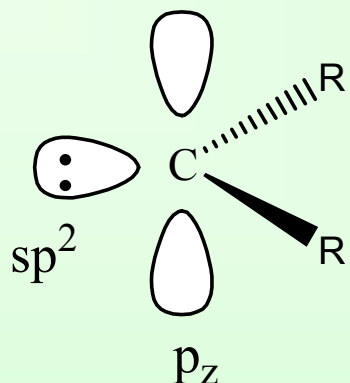
II.b. Ligands with other atoms attached to metal

Alkylidene (carbene) ligand



A free carben CR_2 has two spin states (not resonance forms):

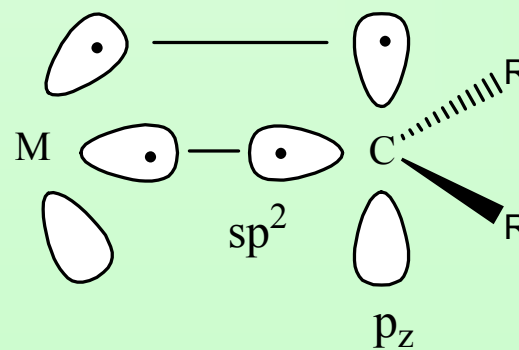
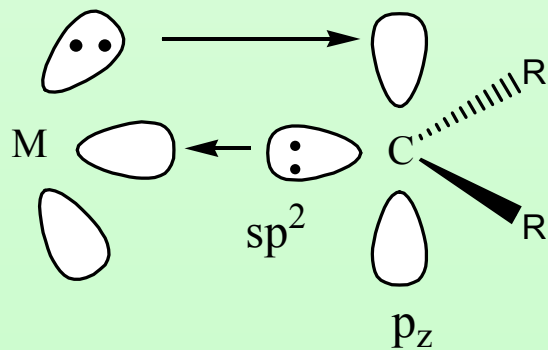
singlet ($\uparrow\downarrow$) and **triplet** ($\uparrow\uparrow$)



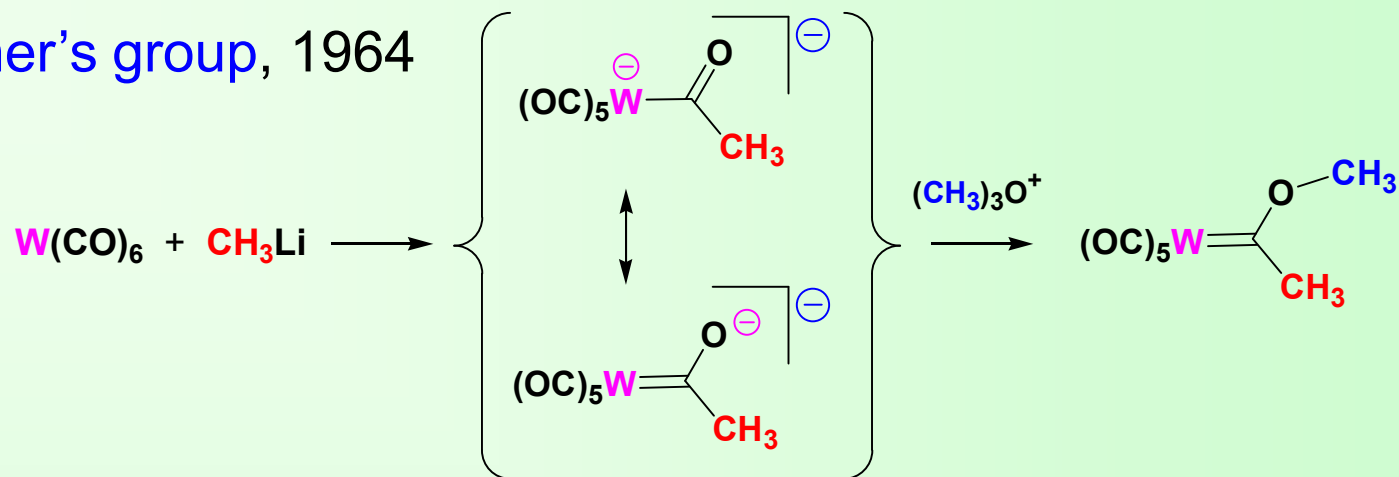
Two types of coordinated carbene can be distinguished:

Fischer type carbene

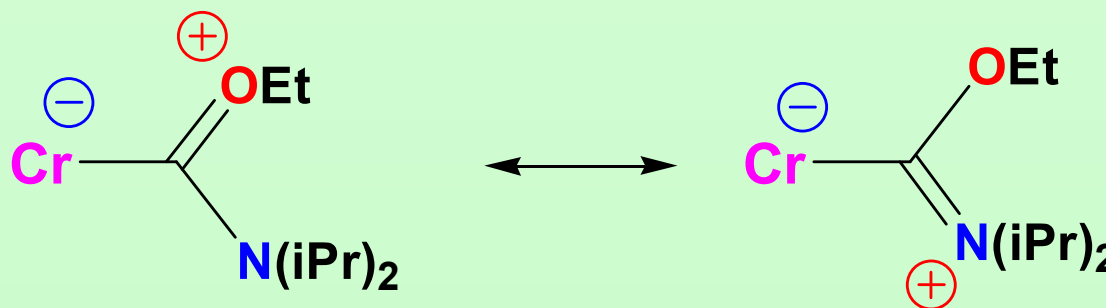
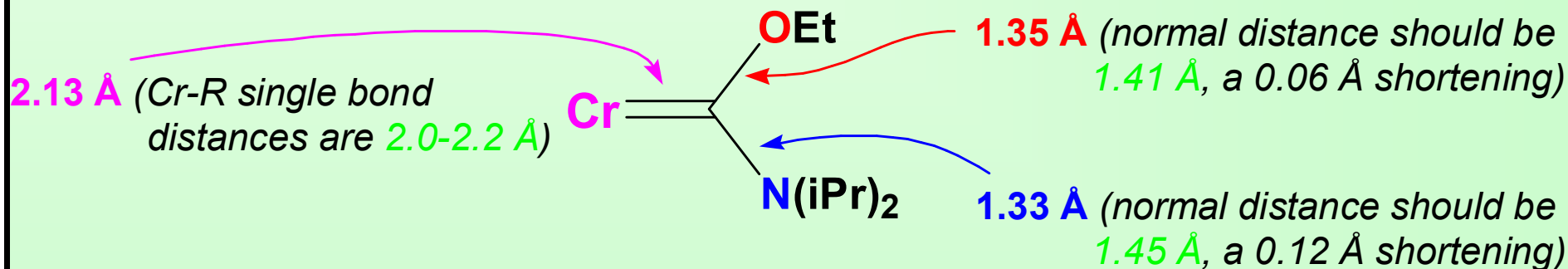
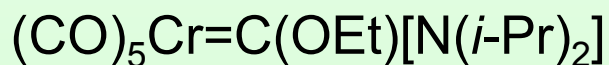
and **Schrock** type carbene



Fischer's group, 1964

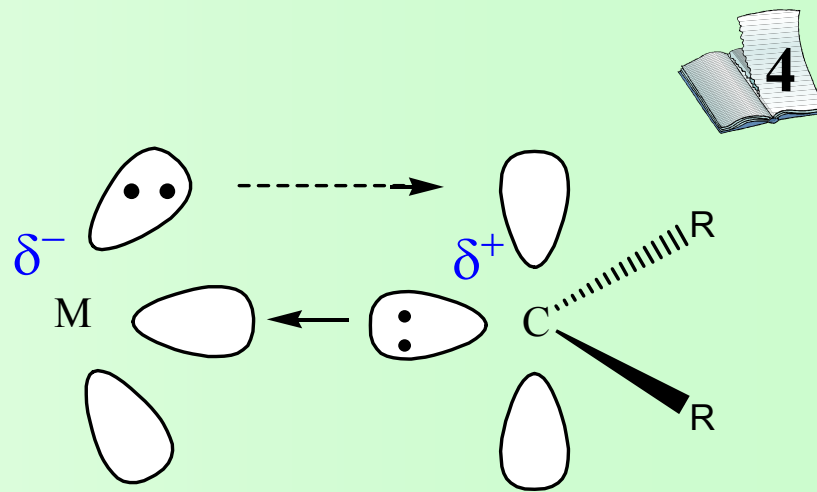


Structural features of Fischer type carbenes:

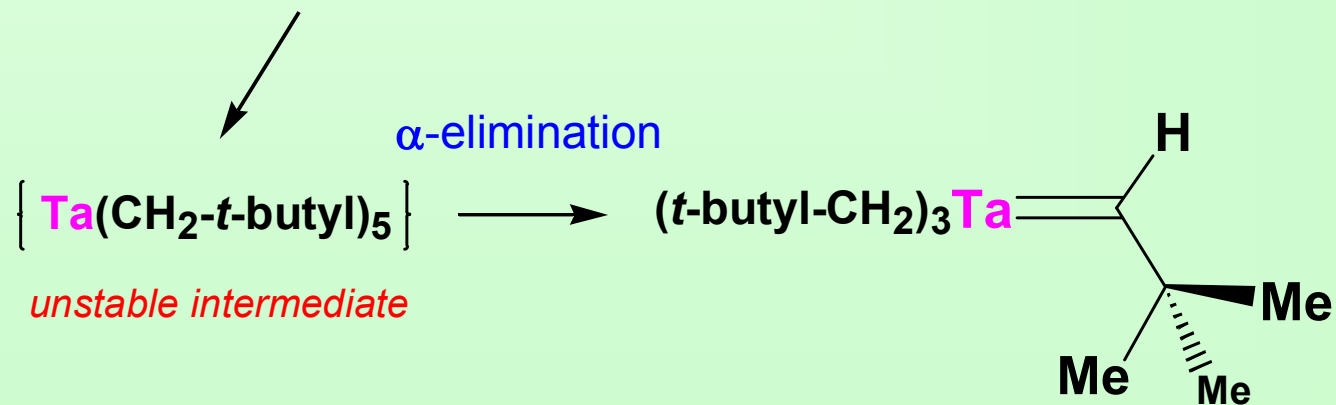


Fischer type carbene:

- middle to late transition metals
- low oxidation state
- π – acceptor ligands at M (CO, NO)
- π – donor substituents (OMe, NMe₂)

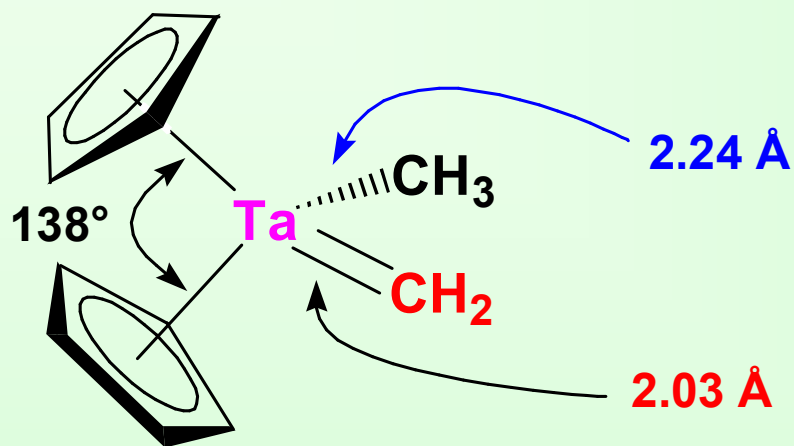
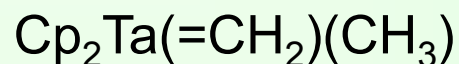


Schrock, 1973



+ neopentane

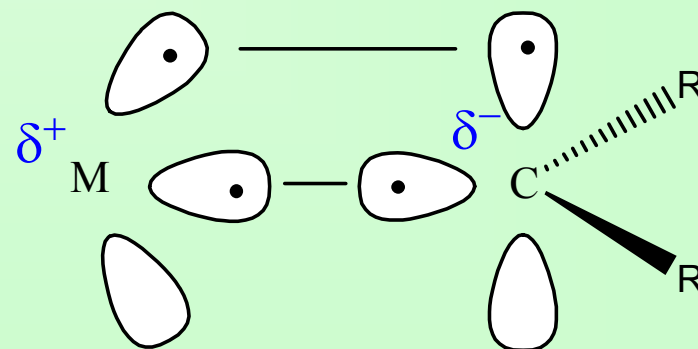
Structural features of Schrock type carbenes:



The Ta=CH₂ bond is distinctly shorter than the Ta-CH₃ single bond!

Schrock type carbene:

- early transition metals
- high oxidation state
- ligands with strong σ , π – donor capacity
- substituents with no π – acceptor capacity (H, CR₃, SiR₃)



Let's compare carbenes of **Fischer** and **Schrock** type:

Fischer Carbenes	Schrock Alkylidenes
Nucleophilic attacks at carbon atom of carbene (carbon is electron deficient)	Electrophilic attacks at carbon atom of alkylidene (carbon is electron-rich)
Electrophilic attacks on metal center (metal is more electron-rich, often d^6 18 e^- system)	Nucleophilic attacks on metal center (metal is electron-deficient, usually d^2 or d^0 16 or 14 e^- count)
Carbene is <u>stabilized</u> by heteroatom groups that can π -bond to it. Likes NR_2 , SR, OR, or Ph groups.	Alkylidene is <u>destabilized</u> by heteroatom groups that can π -bond to it. Strongly prefers H or simple alkyl groups.
Later transition metals favored, especially with d^6 counts (carbene as neutral $2e^-$ donor ligand)	Early transition metals favored, especially with d^0 centers (alkylidene as dianionic $4e^-$ donor)



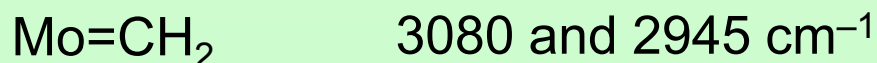
NMR data for carbenes:

^{13}C resonance for carbenes occurs to low field: 200-350 ppm,

^1H resonance for $\text{M}=\text{CHR}$ occurs at +10...+20 ppm

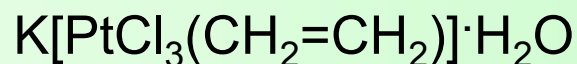
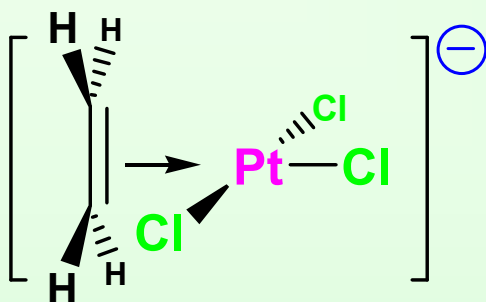
Compound	^{13}C , δ (ppm)	Class
$\text{Cp}_2\text{Ta}(\text{=CH}_2)(\text{Me})$	224	Schrock
$(\text{t-BuCH}_2)_3\text{Ta}(\text{=CH}(\text{t-Bu}))$	250	Schrock
$(\text{OC})_5\text{Cr}(\text{=CH}(\text{NMe}_2))$	246	Fischer
$(\text{OC})_5\text{Cr}(\text{=CPh}(\text{OMe}))$	351	Fischer
$(\text{OC})_5\text{Cr}(\text{=CPh}_2)$	399	Fischer

IR data: carbene complexes are characterized usually only through ν_{CH} IR absorption ($\nu_{\text{M}=\text{C}}$ is seldom unambiguously identified)

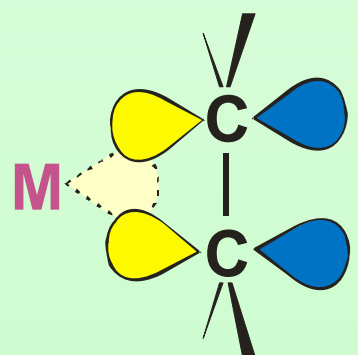


Olefinic ligands

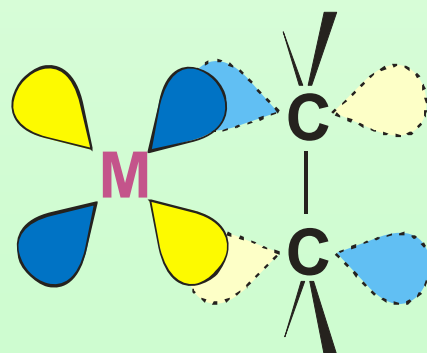
First organometallic compound – Zeise salt, 1831



Alkenes act as **neutral $2e^-$ donors** (per C=C double bond). Due to the presence of empty π^* antibonding orbitals, there is the possibility of some π -backbonding:



σ -donation via the filled alkene π -system



π -back donation via the empty alkene π^* -system

Dewar-Chatt-Duncanson bonding model (1953)

	$d_{\text{C-C}}$, nm
$\text{CH}_2=\text{CH}_2$	1.337
$\text{K}[\text{PtCl}_3(\text{CH}_2=\text{CH}_2)]$	1.375
$\text{Cp}^*_2\text{Ti}(\text{CH}_2=\text{CH}_2)$	1.438
CH_3-CH_3	1.540



IR-spectroscopy in analysis of olefinic complexes

Electronic effects with **alkenes** can often be somewhat easily monitored using IR-spectroscopy. The more **π -backbonding**, the **weaker** the C=C double bond and the **lower** the C=C stretching frequency in the IR-spectrum

The table shows a series of alkene compounds with differing amounts of **σ -donation** and **π -backbonding**

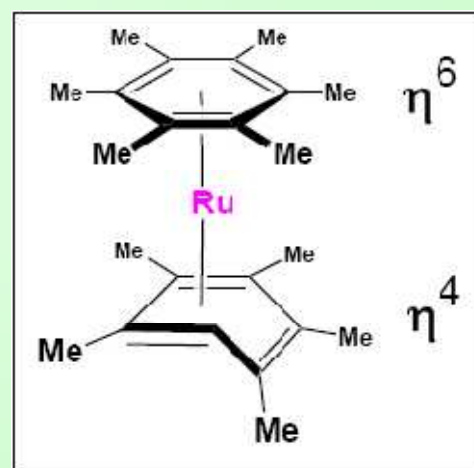
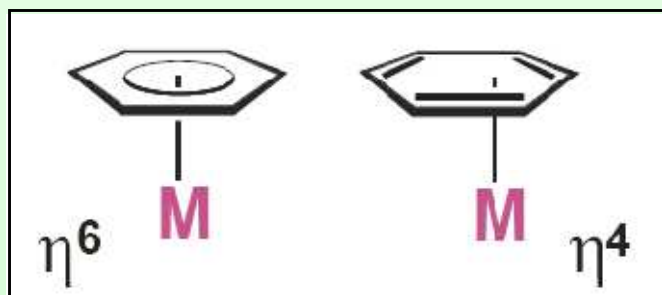
Ethylene Complex	$\nu_{\text{C=C}}$ (cm^{-1})
Free Ethylene	1623
$[\text{Ag}(\text{H}_2\text{C}=\text{CH}_2)_2]^+$	1584
$\text{Fe}(\text{CO})_4(\text{H}_2\text{C}=\text{CH}_2)$	1551
$[\text{Re}(\text{CO})_4(\text{H}_2\text{C}=\text{CH}_2)_2]^+$	1539
$[\text{CpFe}(\text{CO})_2(\text{H}_2\text{C}=\text{CH}_2)]^+$	1527
$\text{Pd}_2\text{Cl}_4(\text{H}_2\text{C}=\text{CH}_2)_2$	1525
$[\text{PtCl}_3(\text{H}_2\text{C}=\text{CH}_2)]^-$	1516
$\text{CpMn}(\text{CO})_2(\text{H}_2\text{C}=\text{CH}_2)$	1508
$\text{Pt}_2\text{Cl}_4(\text{H}_2\text{C}=\text{CH}_2)_2$	1506
$\text{CpRh}(\text{H}_2\text{C}=\text{CH}_2)_2$	1493

Factors influence *thermodynamic stability* of olefinic complexes:

- 1) Electron-withdrawing groups on the alkene generally increase the strength of the metal-alkene bonding
- 2) In cases where *cis-trans* isomerism is possible, the more stable complex is almost always formed by the *cis*-alkene (steric factor)
- 3) Third-row metals form the strongest bonds and most stable complexes

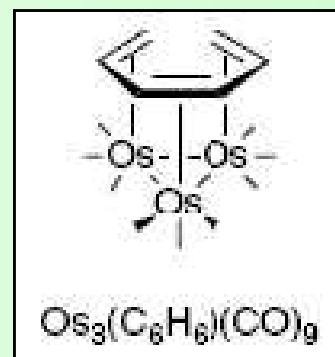
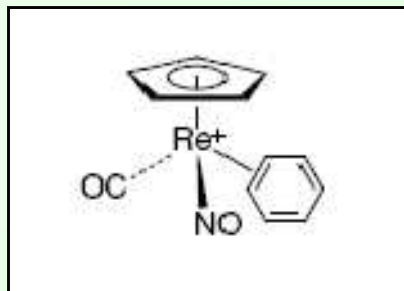
Arene ligands

Arenes typically coordinate in an η^6 fashion and as such are **neutral 6 e- donors**, although they can adopt lower coordination modes (η^4 and η^2).

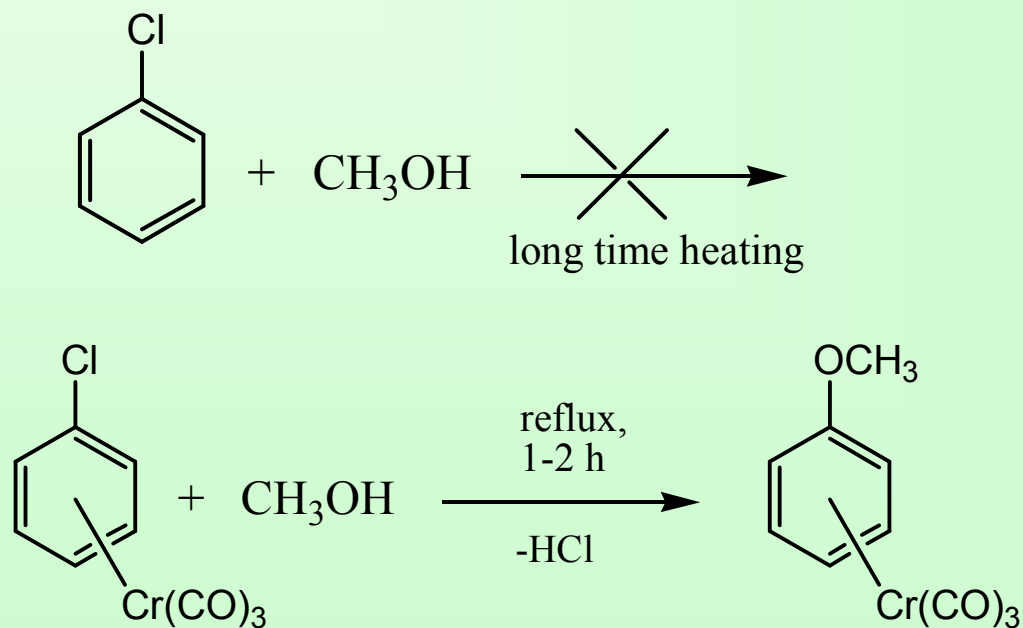


An interesting aspect of metal-arene complexes is that π -back donation plays a relatively important role in the bonding and chemistry.

Other types of coordination:

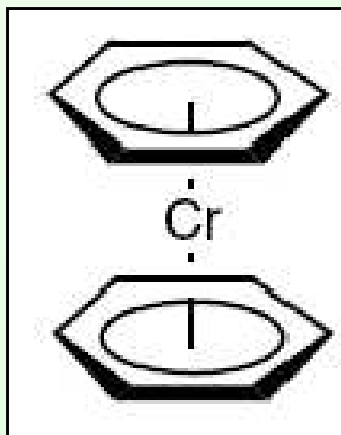


The reactivity of arenes in metals complexes differs from the “free” state:

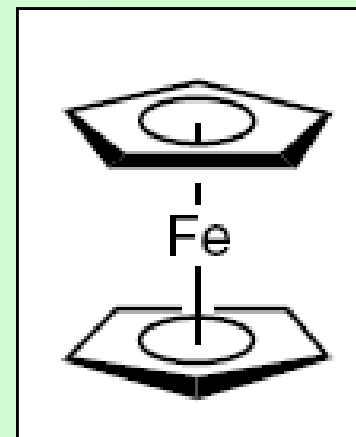


Electron-withdrawing Cr(CO)₃ or Mn(CO)₃ fragments facilitate nucleophilic addition and substitution and other processes.

Sandwich bis-arene complexes



Bis-arene complexes are close to bis-cyclopentadiene complexes, but arene ligand is less tightly bonded to the metal center.

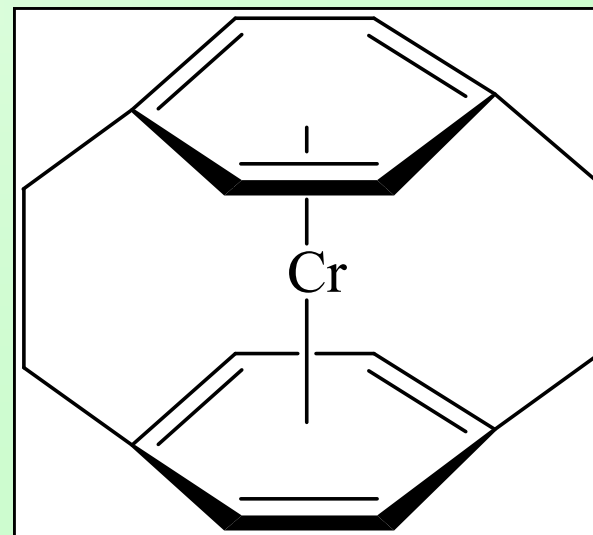


Cr-C₆H₆ bond energy is 40 kcal/mol

Fe-C₅H₅ bond energy is 52 kcal/mol

Ansa-metallocenes and metallocenophanes:

When two rings (Cp or arene) are bridged together we have new class of organometallic compounds – ansa-metallocenes. They have an application in the synthesis of stereoregular polymers, treating cancer and so on.



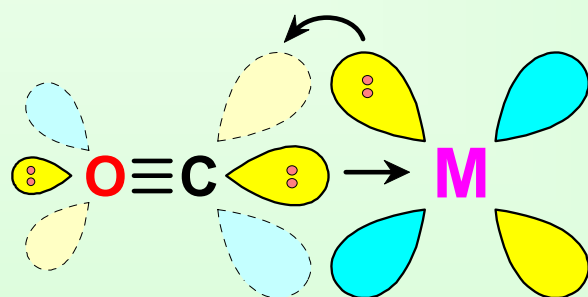
Carbonyl ligand



Ludwig Mond, 1888

Examples of complexes: $\text{Fe}(\text{CO})_5$, $\text{Ni}(\text{CO})_4$, $\text{Co}_2(\text{CO})_8$, $\text{Mn}_2(\text{CO})_{10}$, $\text{Cr}(\text{CO})_6$

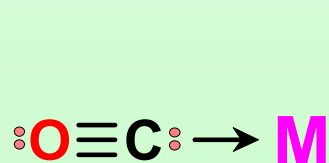
empty π^ -acceptor orbitals on carbonyl*



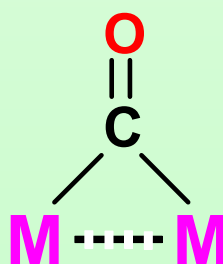
powerful π -acceptor ligand!

excellent ligand, therefore, for stabilizing **electron-rich** low-valent metal centers

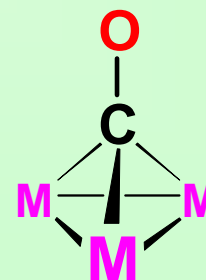
Standard Bonding Modes:



terminal mode
 $2e^-$ neutral donor



μ_2 -bridging mode
 $2e^-$ neutral donor



μ_3 -bridging mode
 $3e^-$ neutral donor

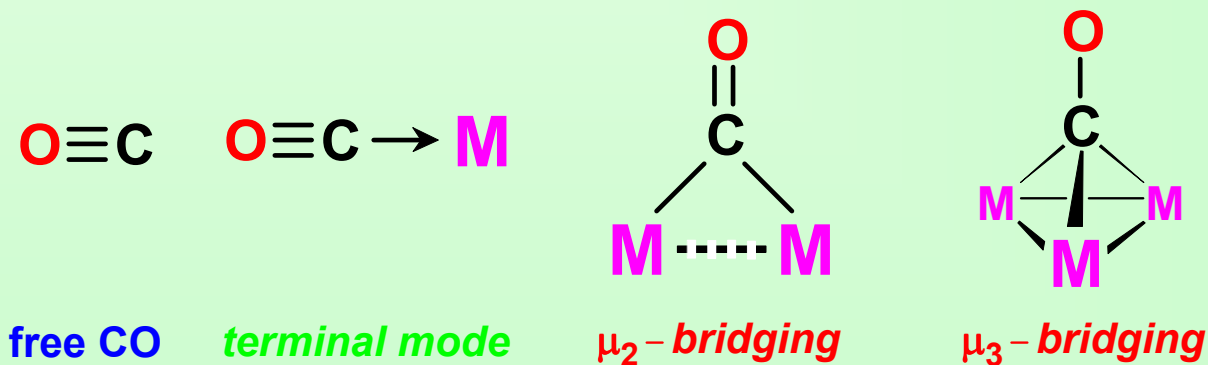
Carbonyl IR Stretching Frequencies



The **position** of the carbonyl bands in the IR depends mainly on the bonding mode of the CO (terminal, bridging) and the amount of electron density on the metal being π -backbonded to the CO.

The **number** (and intensity) of the carbonyl bands one observes depends on the number of CO ligands present and the symmetry of the metal complex.

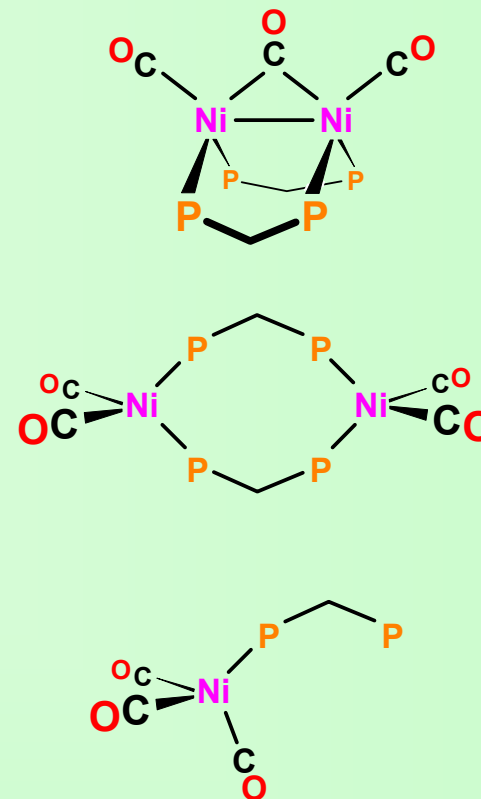
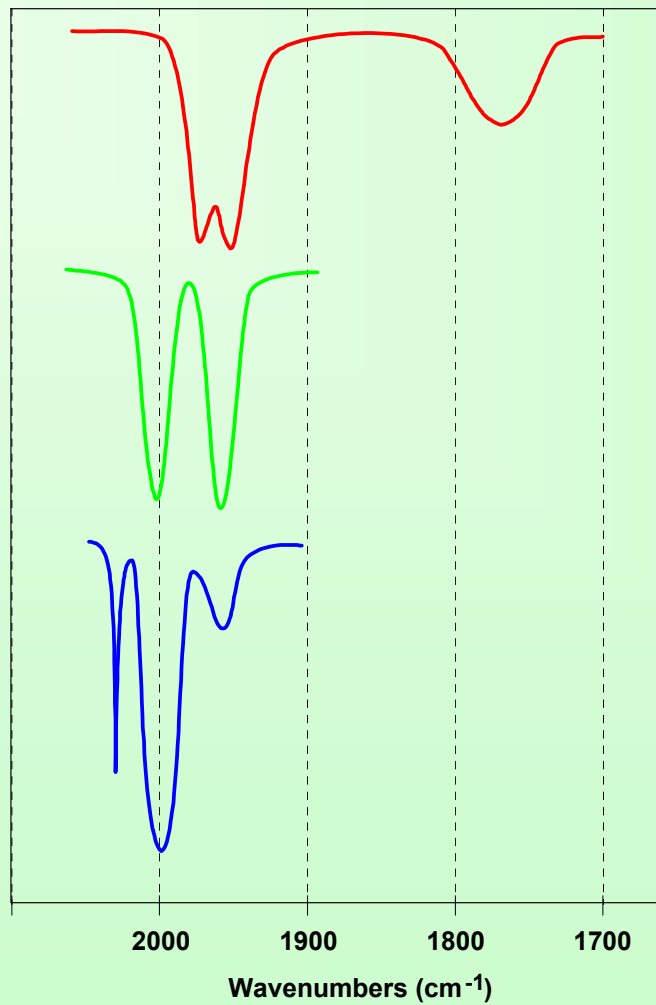
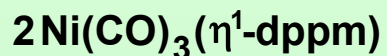
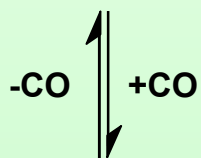
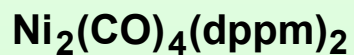
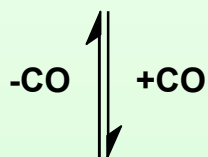
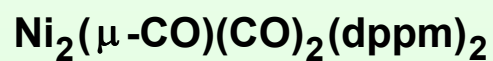
Bonding Modes: As one goes from a terminal CO-bonding mode to μ_2 -bridging and finally μ_3 -bridging, there is a relatively dramatic drop in the CO stretching frequency seen in the IR.



ν_{CO} IR (cm^{-1})	2143	2120 - 1850	1850 - 1720	1730 - 1500
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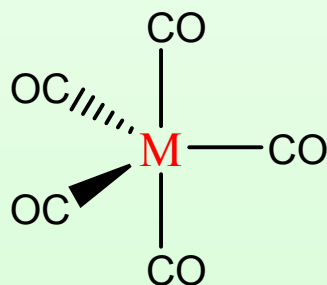
(for neutral metal complexes)

The carbonyl region in the IR spectrum can be very distinctive and useful for help in assigning structures and for indicating the relative amount of electron density present on the metal:



Carbonyl NMR Shifts

Terminal carbonyl ligands generally give rise to ^{13}C NMR resonances in the region 160-230 ppm, the chemical shift (δ) being most significantly dependent on the metal center. In general, heavier metals cause shifts to the higher field. Bridging carbonyls typically resonate to lower field.

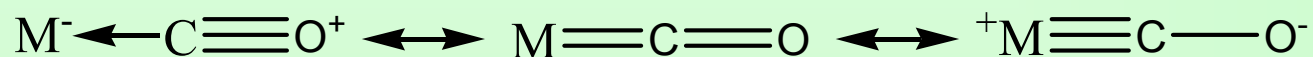


^{13}C NMR shifts: M = **Fe** $\delta = 211$ ppm

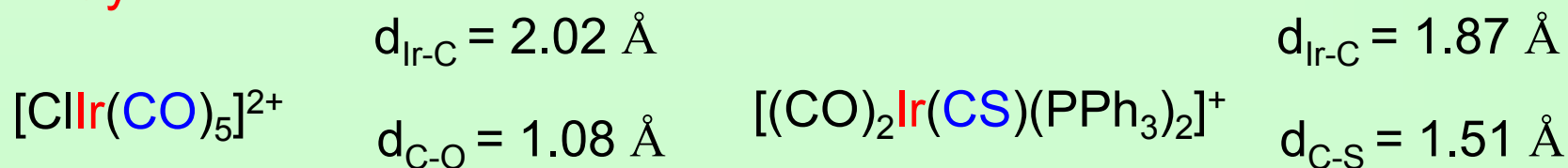
M = **Ru** $\delta = 200$ ppm

M = **Os** $\delta = 183$ ppm

Bonding M-CO could be drawn as a set of resonance forms:



X-Ray:



II.b. Ligands with other atoms attached to metal (H, halogens, CO, phosphines)

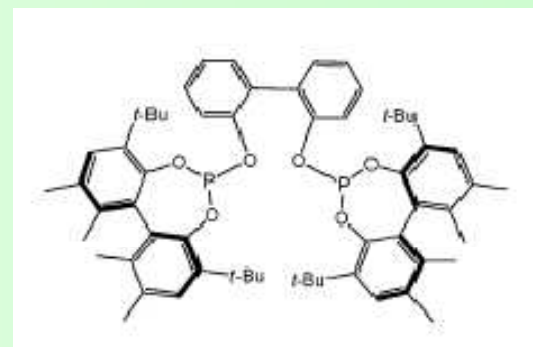
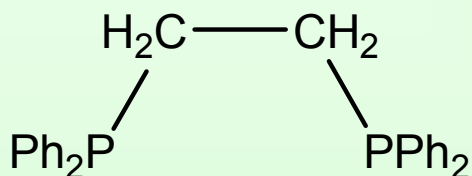


Phosphorus(III) ligands

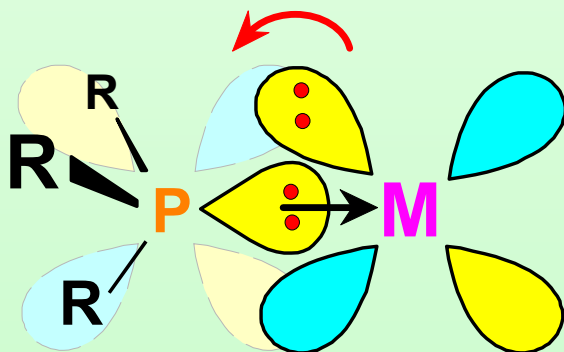
Examples:

Phosphines: PMe_3 , PPh_3

Phosphites: P(OMe)_3 , P(OPh)_3



M-P bond formation is close to M-(CO):



Phosphine ligands

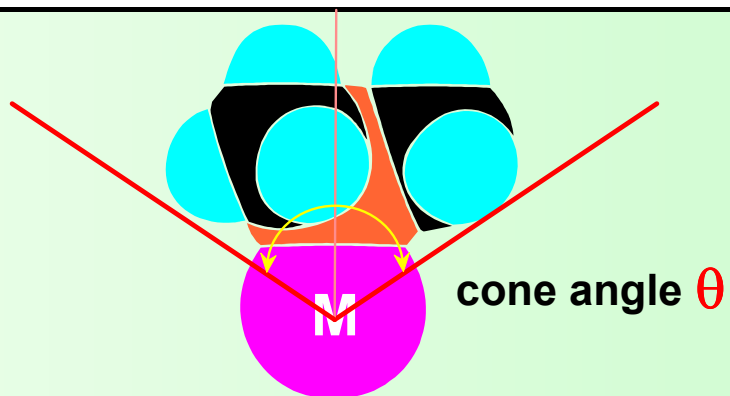
*excellent soft-donor ligands
with a wide variety of easily adjusted
steric and electronic factors*

neutral $2e^-$ donor

But in compare with CO phosphines have more donor ability:

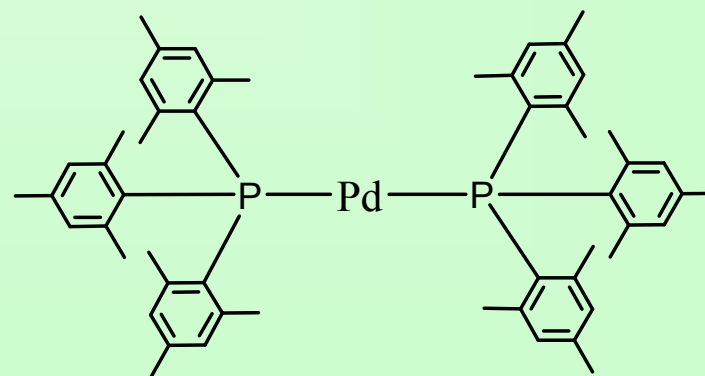
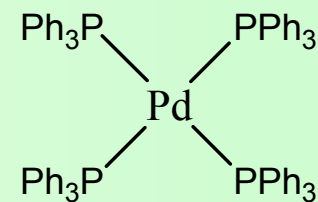
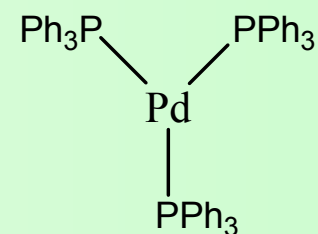


Tolman, 1977



cone angle θ

Phosphine/Phosphite	Cone angle
PH_3	87°
PF_3	104°
$\text{P}(\text{OMe})_3$	107°
PMe_3	118°
PPh_3	145°
$\text{P}(\text{mesityl})_3$	212°



Structural features of phosphine metal complexes

Some typical first row M - PR_3 average bond distances:

Ti-P 2.6 Å

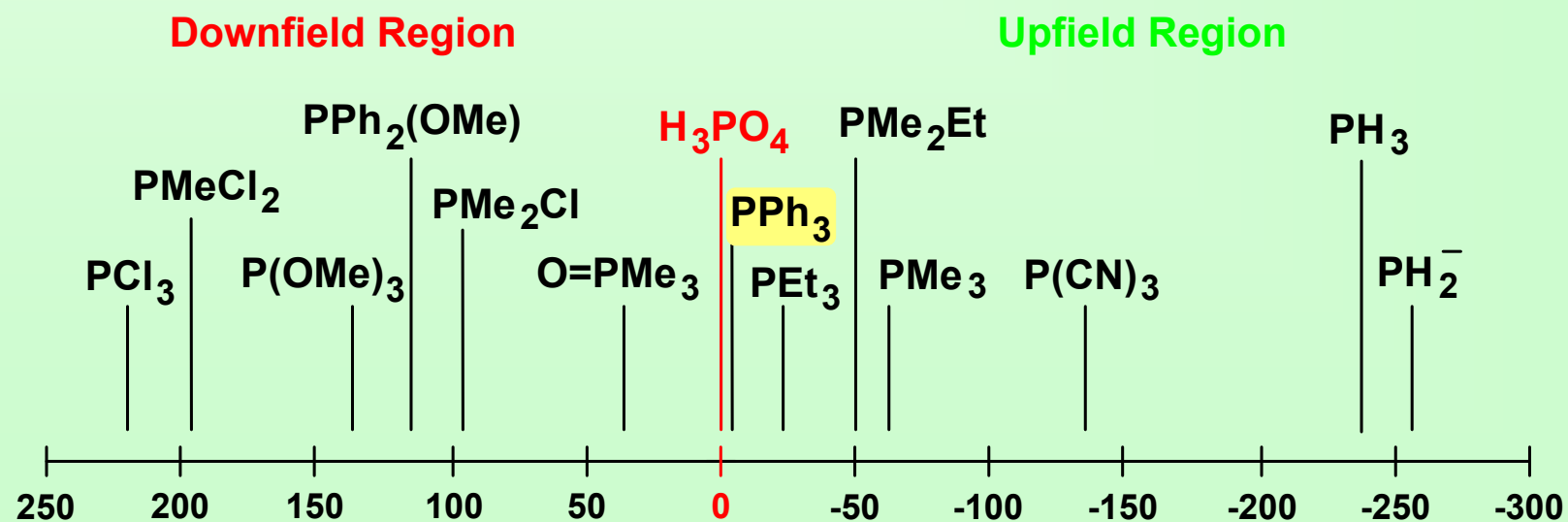
V-P 2.5 Å

Cr-P 2.4 Å

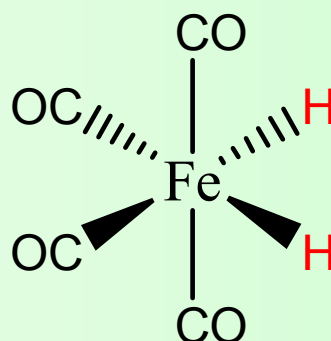
Ni-P 2.1 Å

NMR – ^{31}P characteristics of phosphine complexes

^{31}P NMR shifts: + 245...-163 ppm



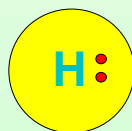
Hydride ligand



Walter Hieber, 1931

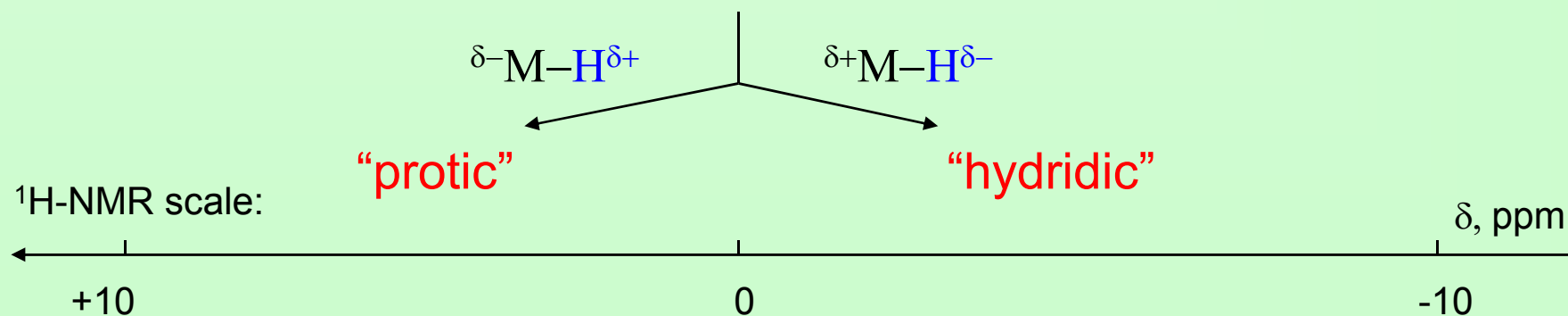
Hydride complexes – quite a common case, many of hydride complexes are stable at ambient conditions

H^- - hydride ligand



anionic $2e^-$ donor

Depending on the TMC, the nature of this ligand can vary from acidic ($\text{HCo}(\text{CO})_4$ is a strong acid) to hydride (like in NaH):

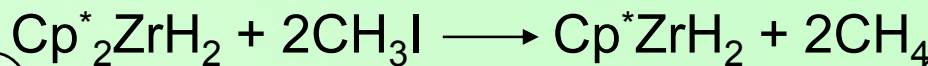
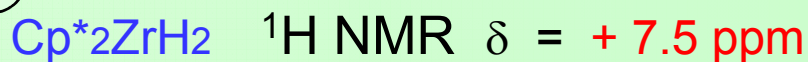
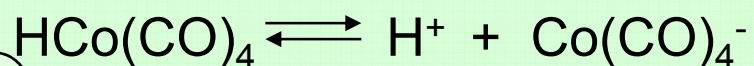
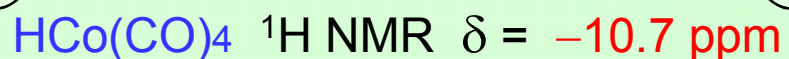


NMR – characterization of hydride complexes:

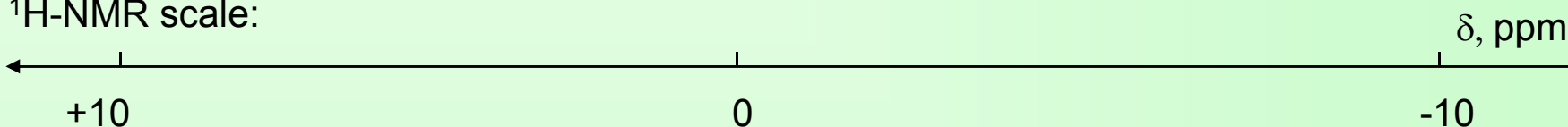


Most hydride ligands show ^1H NMR signal in hydride region – from -3 to -25 ppm, although it could be found in a wider range – from +25 to -60 ppm.

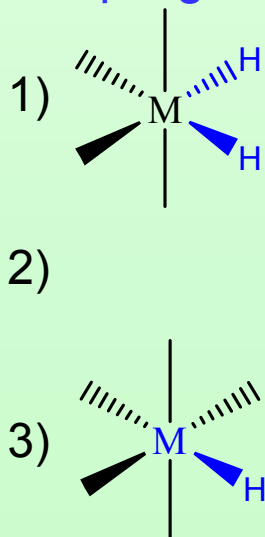
But it could not be a valid criteria of the “acidity” of metal-hydrogen bond



^1H -NMR scale:



Coupling constants – hydrides can couple with the atom, where this has $\frac{1}{2}$ spin:



with other hydride ligand (if they are not equivalent)

$$^2J_{\text{HH}} = 1-10 \text{ Hz}$$

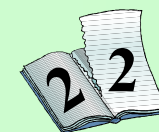
with phosphorus atom (phosphines, phosphites)

$$^2J_{\text{HP}} = 15-30 \text{ Hz for } \textit{cis}\text{-}, 90-150 \text{ Hz for } \textit{trans}\text{-orientation}$$

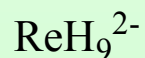
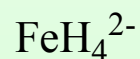
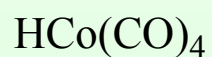
with metal atom (Pt, Rh, W)

$$^1J_{\text{PtH}} = 500-1500 \text{ Hz}$$

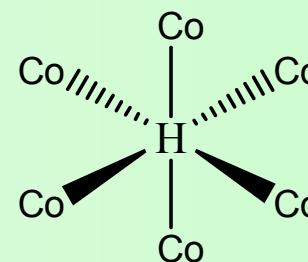
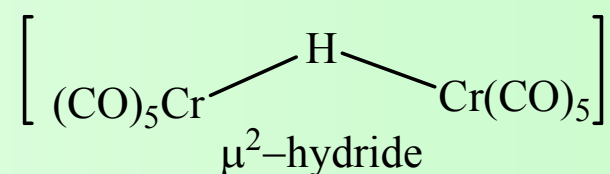
IR – characterization and types of hydride complexes



Terminal hydrides :



Bridged hydrides:



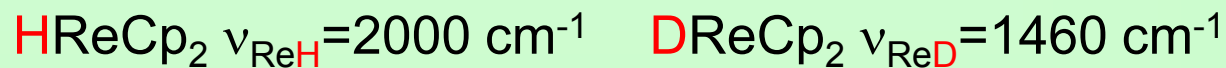
μ^6 -hydride in $[\text{HCo}_6(\text{CO})_{15}]^-$

IR Spectra: 2200 - 1500 cm^{-1}
can be weak or absent

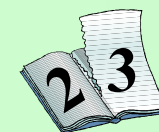
1600 - 800 cm^{-1}
broader (weak or absent)

D-test:

$$\nu_{\text{MH}} = \sqrt{2} \nu_{\text{MD}}$$



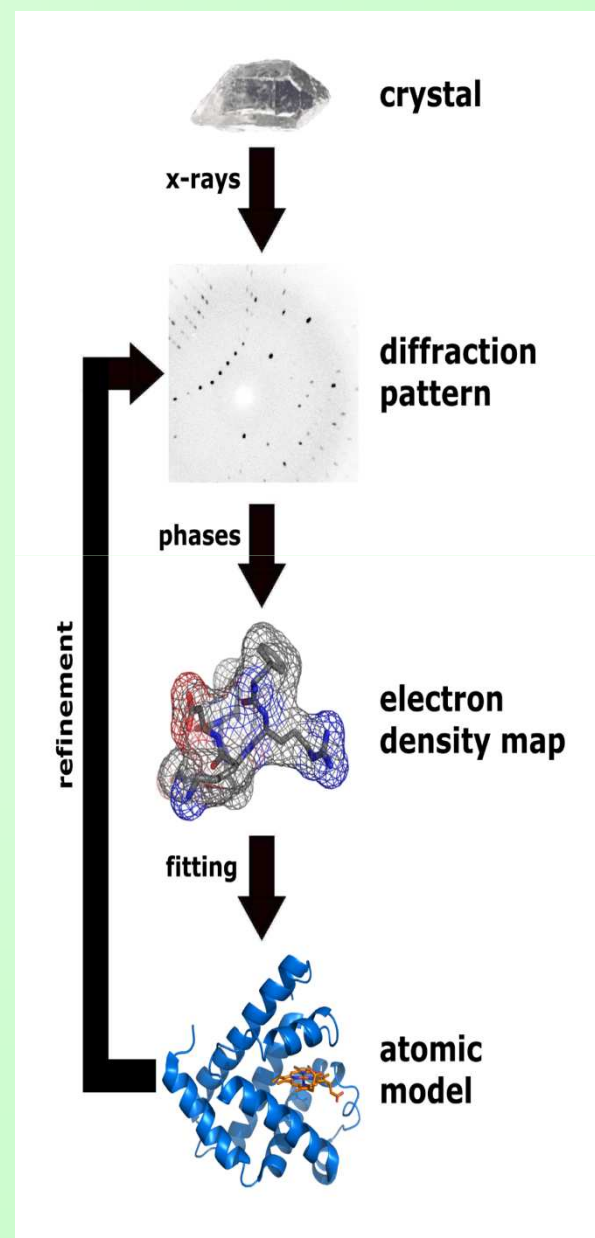
Structural Features (X-ray, neutron diffraction)



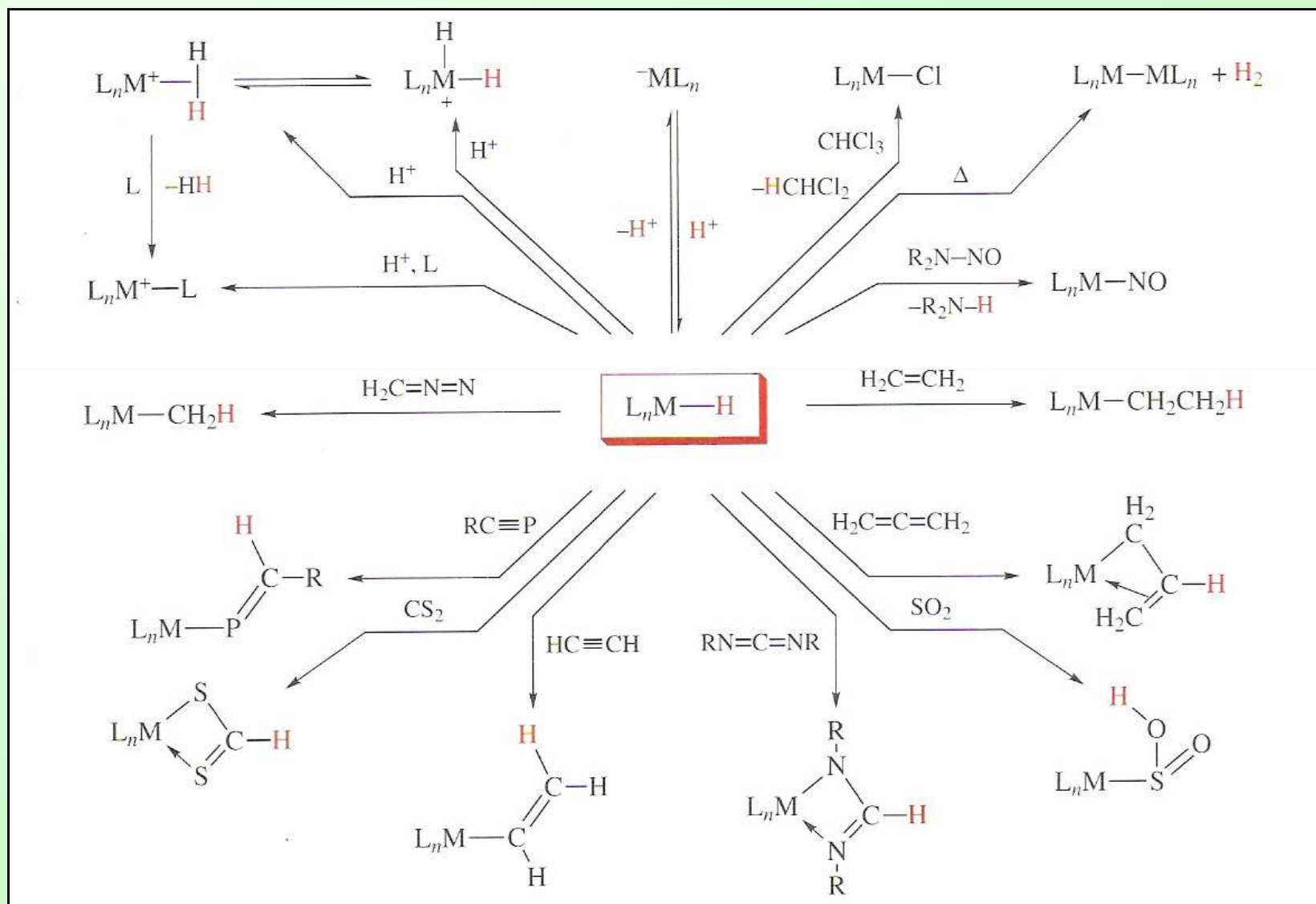
Hydride is the smallest ligand and as a result, M-H distances are typically quite short: 1.8 to about 1.5 Å, depending on the metal.

Hydrides can be quite difficult to observe via X-ray diffraction (the most common technique used to determine structures) due to the very small number of electrons on the hydride vs. adjacent atoms, especially the metal. Since X-rays are scattered by electron density, not by atomic nuclei it is the M-H bonding electrons that are detected, so that X-ray methods systematically underestimate the true M-H internuclear distance by approximately 0.1 Å

Therefore, neutron diffraction studies are considered best for accurately locating and identifying hydrides on metal centers (proton itself scatters neutrons relatively efficiently), even so much larger crystals are usually needed for neutron work (1 mm³ vs. 0.01 m³)

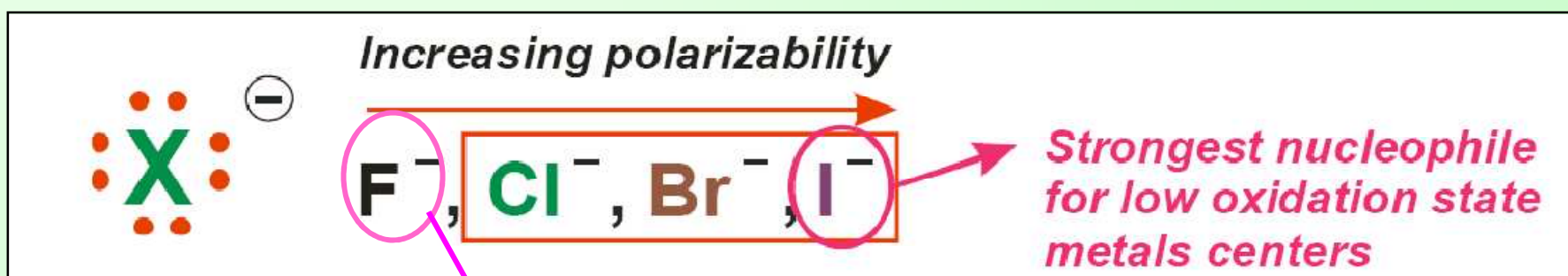


Reactivity of hydride complexes



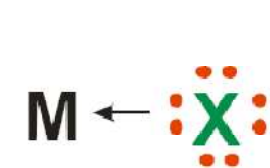
Halogen ligands (halide donors)

The halides are anionic donors that generally only donate $2e^-$ to a metal center. Due to their relatively high electronegativity they are not especially good σ -donor ligands. Although they can theoretically act as π -donor ligands, once again, the higher electronegativity limits them to simple $2e^-$ donor ligands.

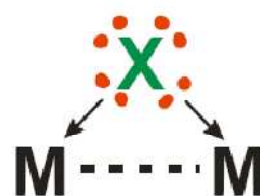


Fluoride is generally NOT a good ligand except for very high oxidation state metal centers. It is too electronegative to donate much of its electron density.

Bonding modes:



2e- terminal

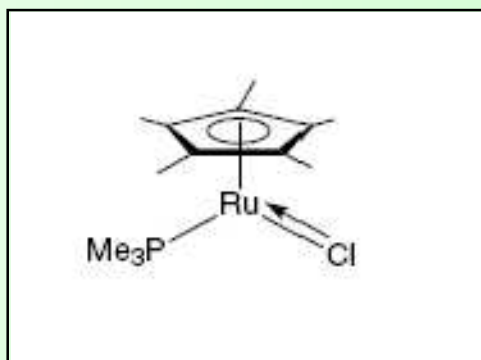


4e- μ -bridging

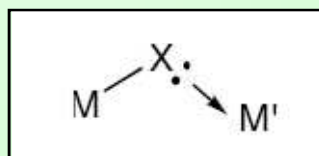


6e- μ_3 -bridging

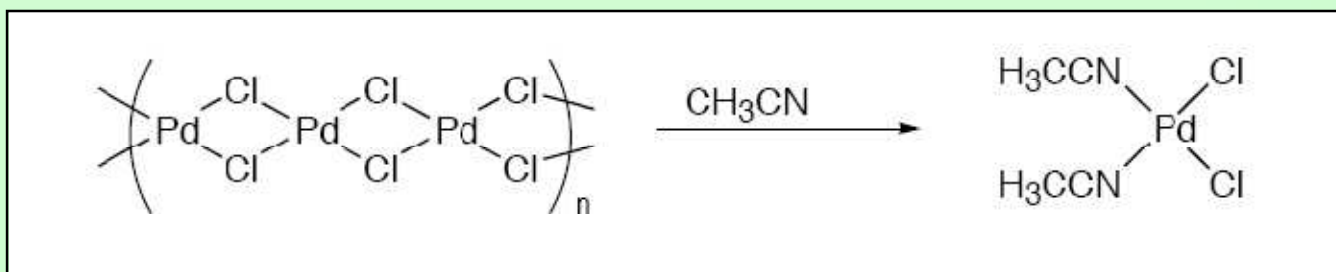
When coordinating to a single metal center usually we regard halides as $2e^-$ donors. $4e^-$ can be more easily donated by I than by any other halide, even so exceptions are possible:



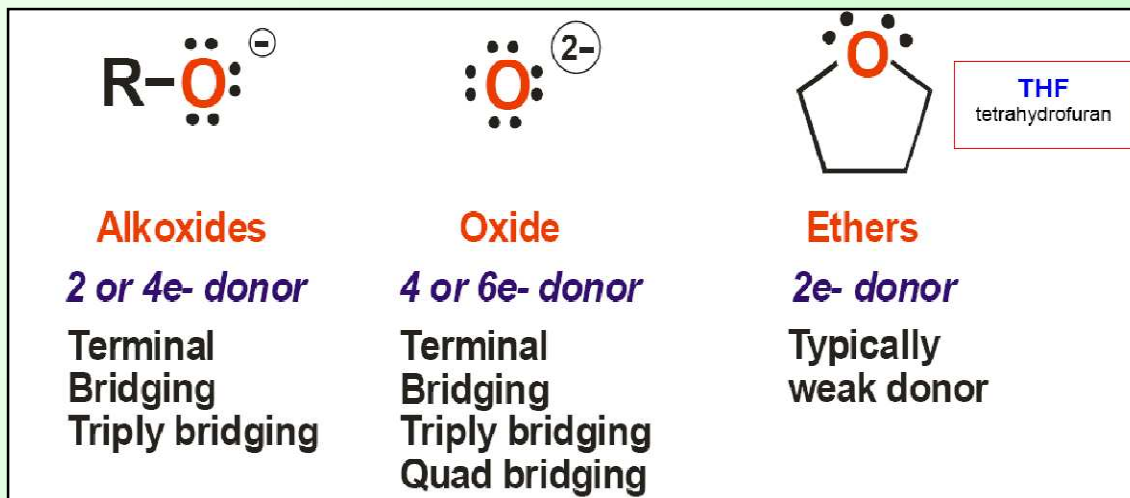
Bridging:



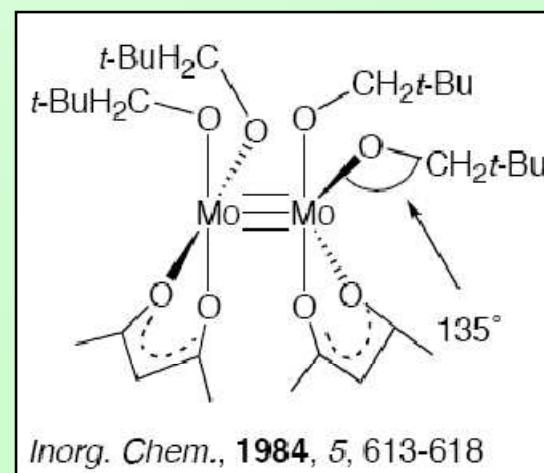
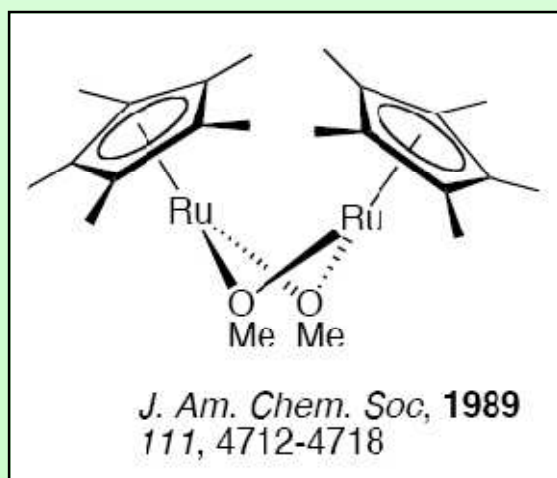
Coordinationally unsaturated metal halides often exist as insoluble polymers connected by bridging halides. PdCl_2 is a classic example. Ligands such as CH_3CN can displace the bridging halide to give the soluble square planar complex $(\text{CH}_3\text{CN})_2\text{PdCl}_2$.



Oxygen donors:



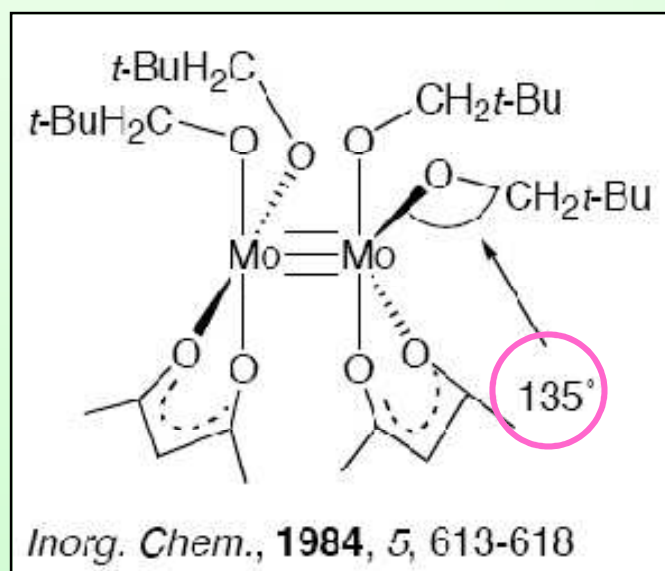
Alkoxides commonly act as bridging ligands for electropositive metals. To avoid formation of bridges, bulky alkoxides must be used. With high valent transition metals, π -donation becomes important.



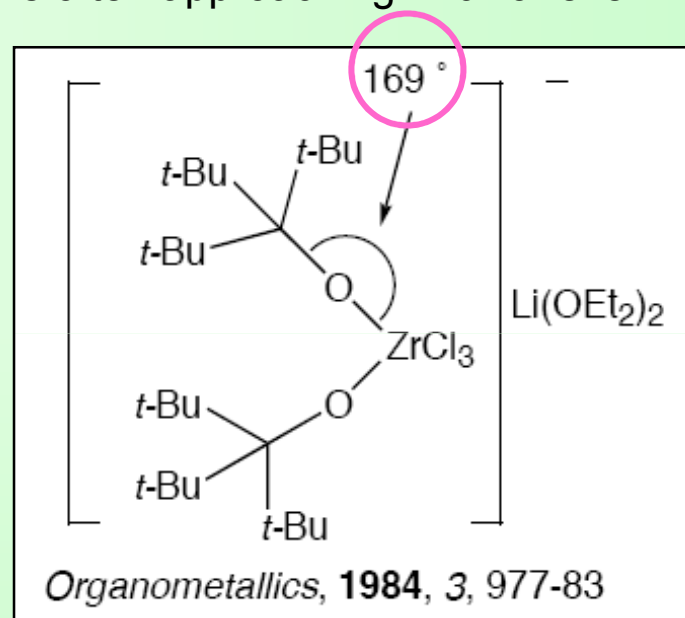
Important information about bonding type gives **M-O-C** angle:

Usually in ethers (like Et₂O) the C-O-C angle is about 110°

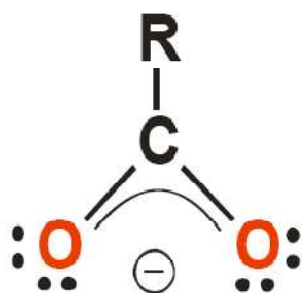
But in the metal-alkoxyde systems M-O-C angle is often approaching 120° or even 180°



An angle near 120° suggests that the O is sp² hybridized and acting as a 4 electron donor.



An angle near 180° suggests that the O is sp hybridized and acting as a 6 electron donor. In this case the alkoxide can be considered to be isoelectronic with the Cp ligand.



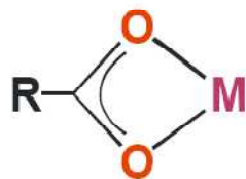
Carboxylates

2 or 4e- donor

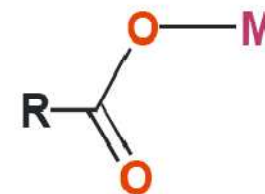
η^1 -terminal

η^2 -terminal

Bridging



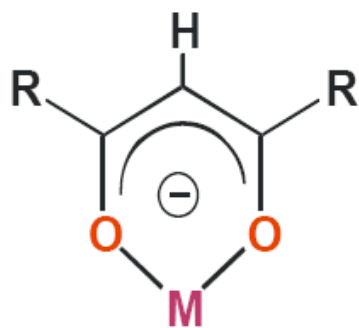
η^2 -chelating 4e-



η^1 2e- mode



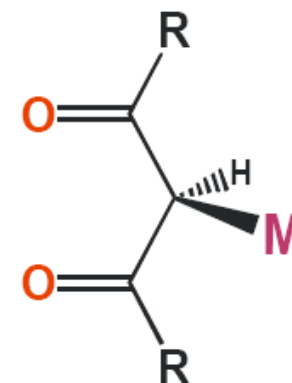
μ -bridging



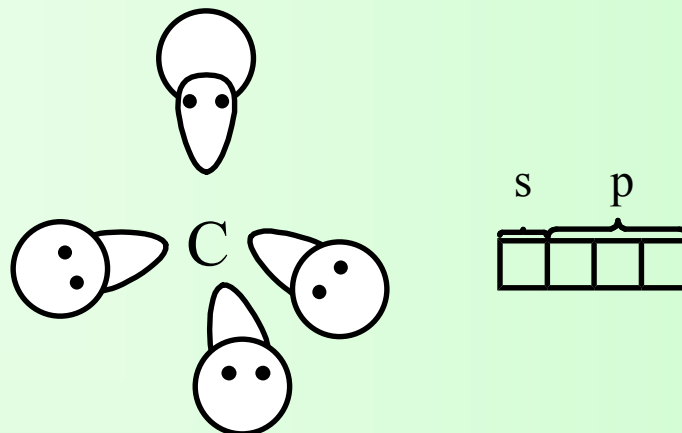
Acetoacetonates (acac)

2 or 4e- donor

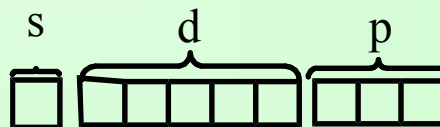
**Unusual C-bonded
acac**



For organic compounds: maximum 8 electrons on C:



For metal complexes: maximum 18 electrons on M:



Complexes with 18 e⁻ counts are referred to as saturated, because there are no empty low-lying orbitals to which another incoming ligand can coordinate.

Complexes with less than 18e⁻ are called unsaturated and can electronically bind additional ligands



Cationic 2e⁻ donor: NO⁺ (nitrosyl)

Neutral 2e⁻ donors: PR₃ (phosphines), CO (carbonyl), R₂C=CR₂ (alkenes), RC≡CR (alkynes, can also donate 4 e⁻), N≡CR (nitriles)

Anionic 2e⁻ donors: Cl⁻ (chloride), Br⁻ (bromide), I⁻ (iodide), CH₃⁻ (methyl), CR₃ (alkyl), Ph⁻ (phenyl), H⁻ (hydride)

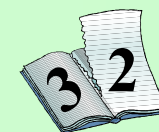
The following can also donate 4 e⁻ if needed, but initially count them as 2e⁻ donors (unless they are acting as bridging ligands):

OR⁻ (alkoxide), SR⁻ (thiolate), NR₂⁻ (amide), PR₂⁻ (phosphide)

Anionic 4e⁻ donors: C₃H₅⁻ (allyl), O²⁻ (oxide), S₂⁻ (sulfide), NR₂⁻ (imide), CR₂⁻ (alkylidene)

and from the previous list: OR⁻ (alkoxide), SR⁻ (thiolate), NR₂⁻ (amide), PR₂⁻

Anionic 6e⁻ donors: Cp⁻ (cyclopentadienyl), O₂⁻ (oxide)



Electron Count Oxidation State Coordination Number

- Basic tools for understanding structure and reactivity.
- Doing them should be “automatic”.
- Not always unambiguous \Rightarrow don't just *follow* the rules, *understand* them!

The basis of counting electrons

- Every element has a certain number of valence orbitals:

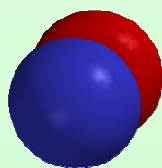
1 (1s) for H

4 (ns , $3 \times np$) for main group elements

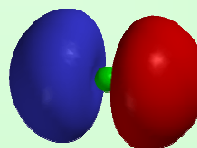
9 (ns , $3 \times np$, $5 \times (n-1)d$) for transition metals



s



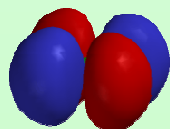
p_x



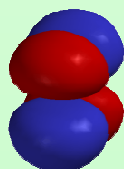
p_y



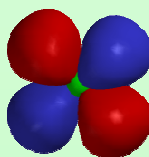
p_z



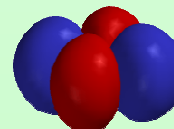
d_{xy}



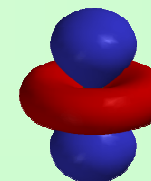
d_{xz}



d_{yz}



$d_{x^2-y^2}$



d_{z^2}



The basis of counting electrons

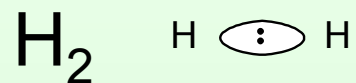
- Every orbital wants to be “used”, i.e. contribute to binding an electron pair.
- Therefore, every element wants to be surrounded by 2/8/18 electrons.
- The strength of the preference for *electron-precise* structures depends on the position of the element in the periodic table.



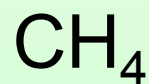
The basis of counting electrons

- **Too few electrons:**
An empty orbital makes the compound very electrophilic, i.e. susceptible to attack by nucleophiles.
- **Too many electrons:**
There are fewer covalent bonds than one would think (not enough orbitals available). An ionic model is required to explain part of the bonding. The "extra" bonds are relatively weak.
- **Metal-centered (unshared) electron pairs:**
Metal orbitals are fairly high in energy. A metal atom with a lone pair is a strong σ -donor (nucleophile) and susceptible to electrophilic attack.

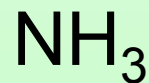
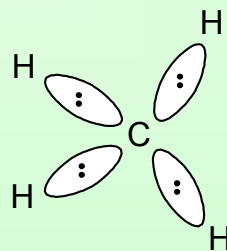
Use a localized (valence-bond) model
to count electrons



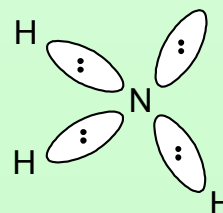
Every H has 2 e. *OK*

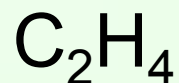


H has 2 e, C 8. *OK*

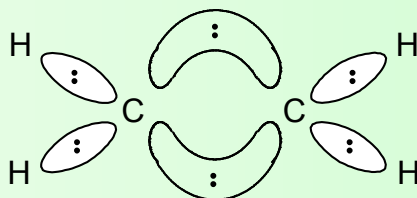


N has 8 e. Nucleophile! *OK*

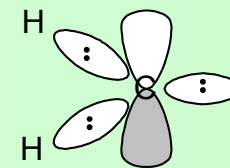




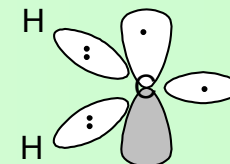
C has 8 e. *OK*

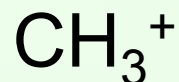


C has only 6 e, and an empty p_z orbital:
extremely reactive ("singlet carbene").
Unstable. Sensitive to nucleophiles and
electrophiles.

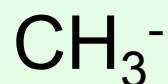
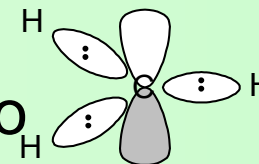


C has only 6 e, is a "biradical" and
extremely reactive ("triplet carbene"), but
not especially for nucleophiles or
electrophiles.

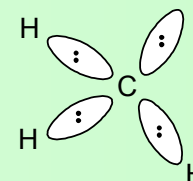




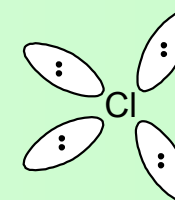
C has only 6 e, and an empty p_z orbital:
extremely reactive. Unstable. Sensitive to
nucleophiles.

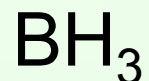


C has 8 e, but a lone pair. Sensitive to
electrophiles.

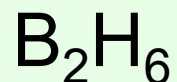
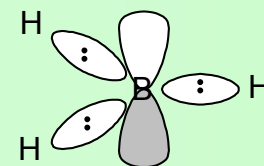


Cl has 8 e, 4 lone pairs. **OK** Somewhat
sensitive to electrophiles.

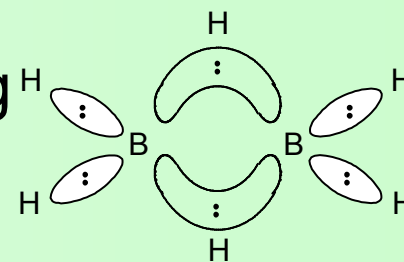




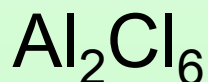
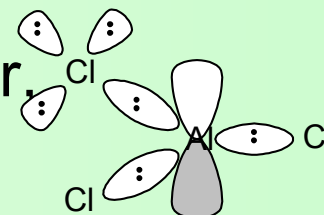
B has only 6 e, not stable as monomer,
forms B_2H_6 :



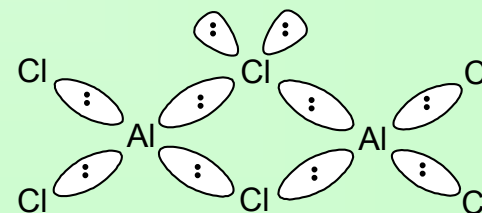
B has 8 e, all H's 2 (including the bridging H!). 2-electron-3-center bonds! *OK*

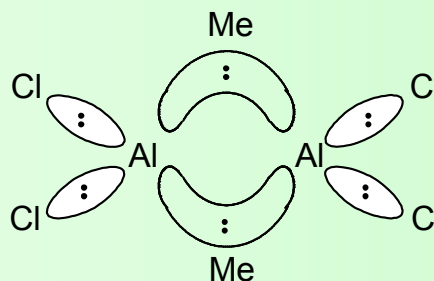
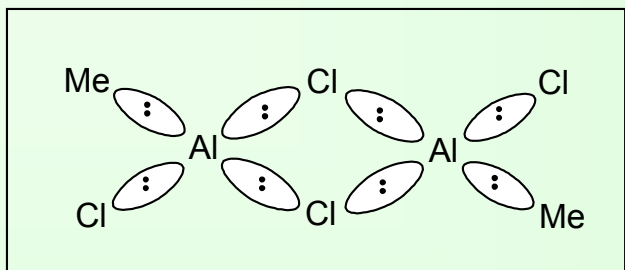
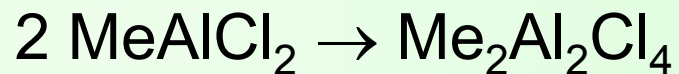


Al has only 6 e, not stable as monomer,
forms Al_2Cl_6 :

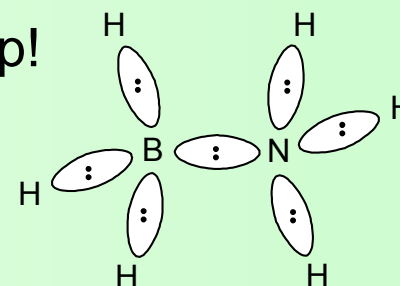
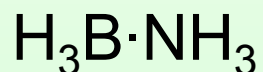


Al has 8 e, all Cl's too (including the bridging Cl!). Regular
2-electron-2-center bonds! *OK*



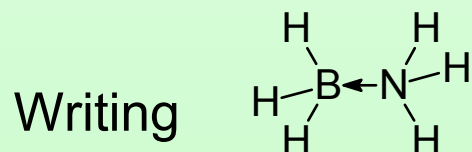
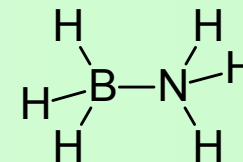


2-electron-3-center bonds are a stopgap!

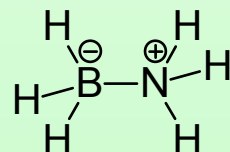


N-B: donor-acceptor bond (nucleophile NH_3 has attacked electrophile BH_3).

Organometallic chemists are "sloppy" and write:



or



would be more

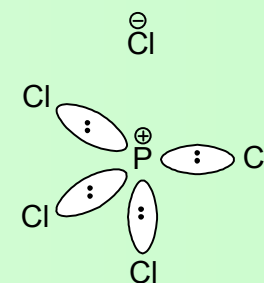
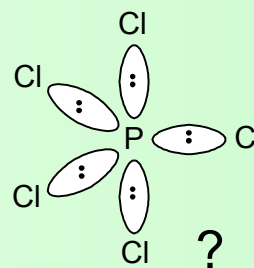
correct (although the latter does not reflect the "real" charge distribution).

PCl₅

P would have 10 e, but only has 4 valence orbitals, so it cannot form more than 4 "net" P-Cl bonds.

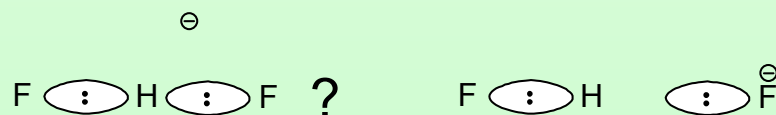
You can describe the bonding using ionic structures (hyperconjugation).

Easy dissociation in PCl₃ + Cl₂.



HF₂⁻

Write as FH·F⁻, mainly ion-dipole interaction.



Examples of electron counting:



1. Nature of central atom (oxidation state, number of electrons)

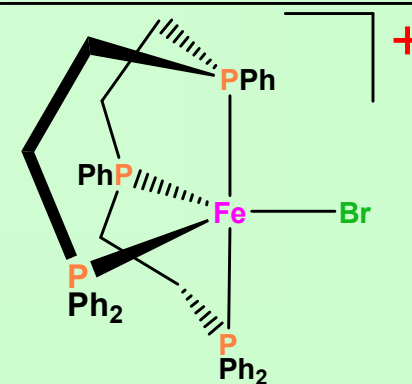
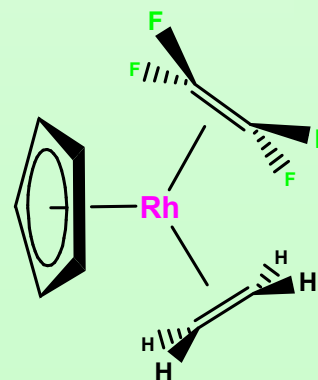
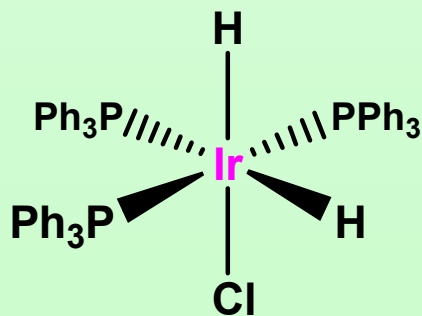
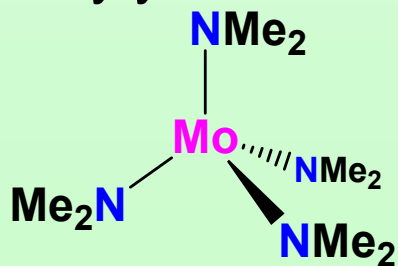
Charge of the complex: **0** Charge of the non-neutral ligands: $2 * (-1) = -2$ Metal oxidation state: **+2**

Number of electrons for Pt^{2+} : for Pt^0 according to the periodic table 10,
For Pt^{2+} : $8e^-$

2. Nature of ligands (number of electrons donating to the metal)

$$4 * (2e^- \text{ - ligands}) = 8e^- \quad \text{Final result: } 8 + 8 = 16e^-$$

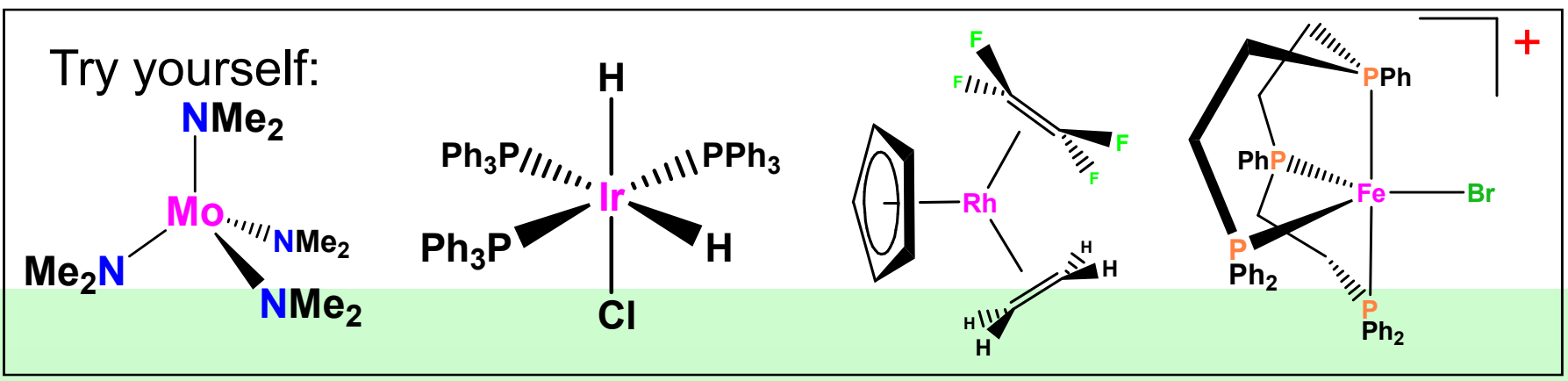
Try yourself:

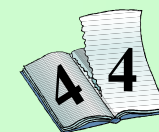


Periodic Table of the Elements

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
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	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Unq	105 Unp	106 Unh	107 Uns	108 Uno	109 Une	110 Unn								

- hydrogen
- alkali metals
- alkali earth metals
- transition metals
- poor metals
- nonmetals
- noble gases
- rare earth metals





Exercises

Give electron count and oxidation state for the following compounds.
Draw conclusions about their (in)stability.

