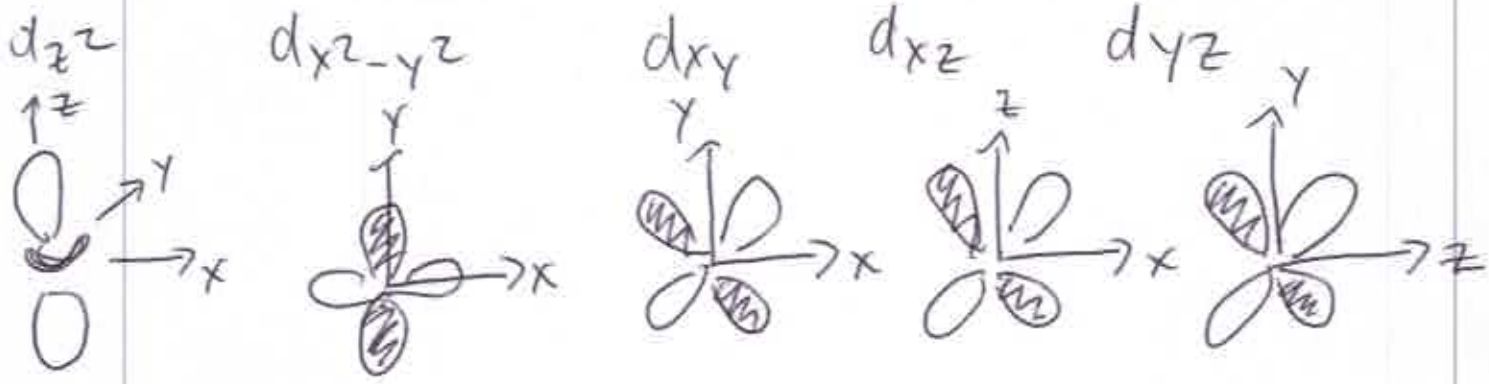
CH431 Lecture 03 9/1/16



Cu^{2+} vs Cu^{+}

↑ smaller than

↑ harder than

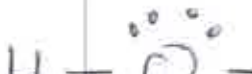
sidenote for future course topics

H_2O

O 2H

6 2x1 = 8 electrons

1) ~~dots~~ can't $\bar{e}$s in molecule

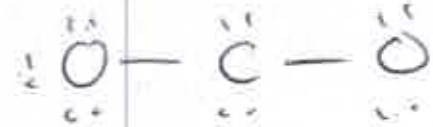


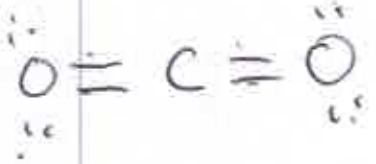
2) Draw single bonds between atoms

3) Fill remaining electrons in

CO_2 C 20

4 2x6 = 16

 20 electrons ✗ not the structure

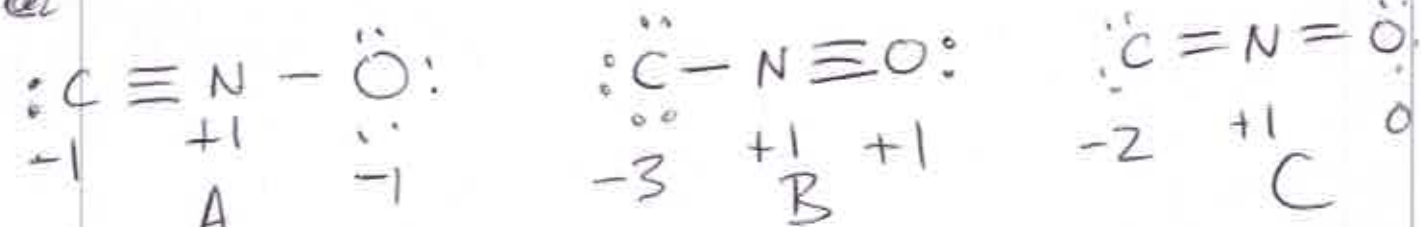
 16 electrons ✓ good structure

---

Formal charge example

$CNO^{\ominus}$ C N O + è

4 + 5 + 6 + 1 = 16 electrons



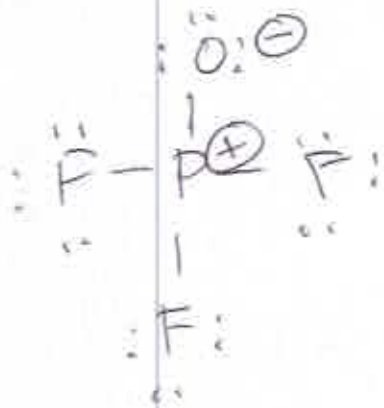
A ✓ negative charge on most electronegative atom

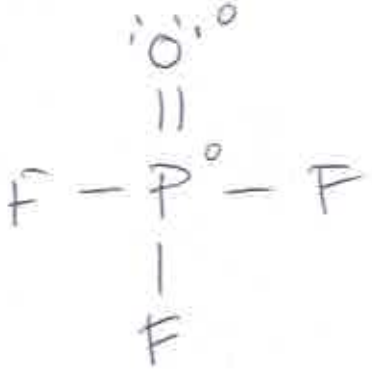
B ✗ Δ formal charge too big

C ✗ Δ formal charge too big

POF_3 P O 3F

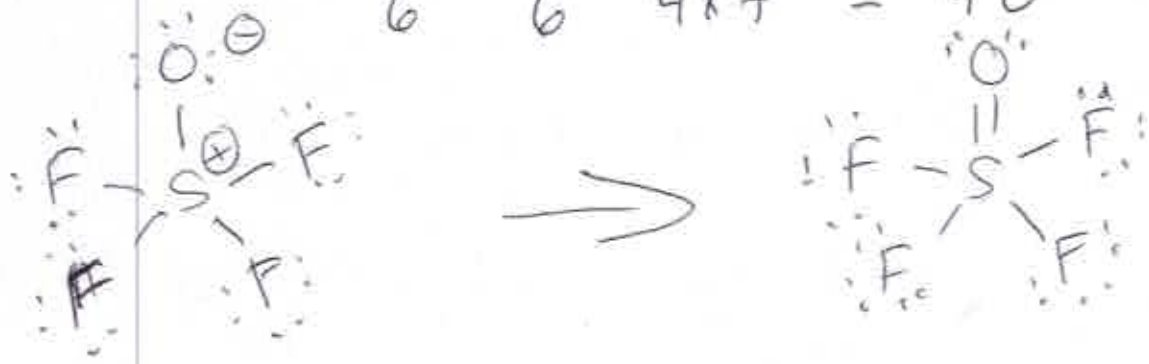
5 6 $3 \times 7 = 32$





SOF_4 S O 4F

6 6 $4 \times 7 = 40$

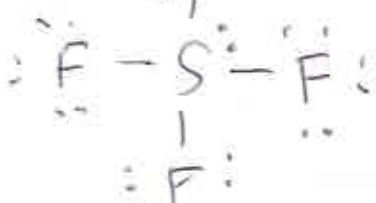


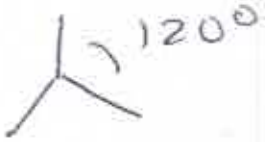
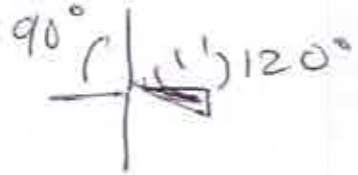
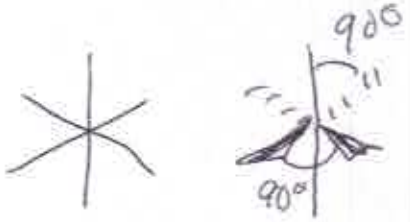
CeF_3 Ce 3F

$7 + 3 \times 7 = 28$

SF_4 S 4F

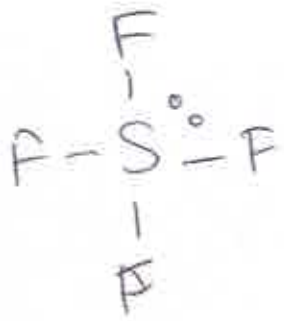
$6 + 4 \times 7 = 34$

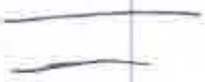


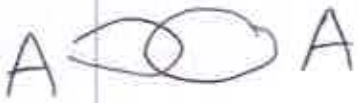
2	ē pairs	linear	:—: 180°
3		trigonal planar	
4		tetrahedral	
5		trigonal bipyramidal	
6		~~octa~~ octahedral	

trigonal bipyramidal

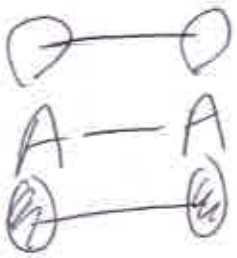
SF_4



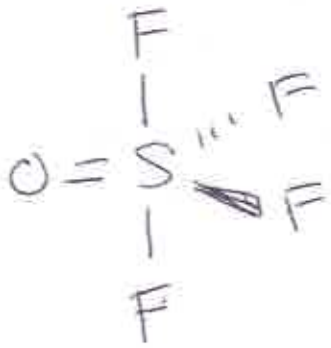




Single bond



double bond



putting S=O bond in equatorial plane minimizes 90° interactions with other ~~bond~~ ē pairs

(similar approach to lone pairs)