(17.) a) AlF_3 b) K_2S c) Y_2O_3 d) Mg_3N_2

(21) (a) lattice energy is the energy required to transform 1 mole of an ionic compound from a solid to individual gas phase ions.

(or, it is the energy released when the gas phase ions form one mole of the ionic solid..)

For example, $MgCl_2$ has a lattice enthalpy of 2326 KJ/mole.

$$MgCl_{2(s)} + 2326\ KJ \longrightarrow Mg^{+2}_{(g)} + 2Cl^-_{(g)}$$

(lattice enthalpy isn't directly measured though)

(b) The stronger the attractive forces between ions, the higher the lattice enthalpy. These attractive forces can be estimated with Coulomb's Law:

$$F = \frac{k q_1 q_2}{r^2}$$

The higher the ion charges, the higher the lattice enthalpy.

The shorter the bond length, the higher the lattice enthalpy.

charges usually have more impact on the lattice energy than size differences, since ion radii are (mostly) with a factor of 2 of each other, for example: $Li^{+1} = .90$ Å $F^- = 1.19$Å

$Rb^{+1} = 1.66$Å $I^{-1} = 2.06$Å

but just doubling the charge of 1 ion can double the force, if all other factors are constant.

(c) Compounds with higher lattice energies tend to have higher melting points.

(23) Look up lattice enthalpies of KF, CaO, ScN (page 293), and explain

KF	808	KJ/mole	+1	and −1	ions	in	1:1 ratio
CaO	3414	"	+2	and −2	ions	in	1:1 ratio
ScN	7547	"	+3	and −3	ions	in	1:1 ratio

all ions are from period 4 elements, so the radii are fairly similar, but the higher the charge on the ions, the stronger the coulomb attractive forces, and the higher the lattice enthalpy.

(The numbers work pretty well: if the bond lengths, etc, were the same, then CaO would have 4× the attractive forces as KF, and ScN would have 9× the attractive forces.

808 : 3414 : 7547

= 1 : 4.2 : 9.3)

b) given: BaS, CaO, KCl, MgO, NaCl, SrO

have these mp, but not in order: 770, 993, 1200, 2351, 2613, and 2852 (all in °C)

which is which? justify choices.

MgO	2852°C
CaO	2613°C
SrO	2351°C
BaS	1200°C
NaCl	993°C
KCl	770°C

all of these have +2 / −2 charges. but Mg^{+2}, O^{-2} have the smallest radii of the group, and Ba^{+2}/S^{-2} have the largest radii; the list is ranked by radii. I assumed that charge was more important than size; it usually is. so these 4 cpds would have stronger attractive forces than those with only +1/−1 charges.

these two only have +1/−1 charges, so should have lower mp than the other 4 cpds, due to weaker coulomb forces. Na^{+1} has a smaller radius than K^{+1}, so NaCl has shorter bond lengths, and stronger attractive forces, than KCl, therefore so NaCl has a higher melting point (mp)

(37) (a) Electronegativity = an atom's ability to attract electrons to itself, or an elements ability/tendency to attract electrons to itself when bonded.

(b) The range in electronegativities is 0.7 to 4.0 (for the elements we have so far...)

(c) F has the greatest electronegativity : EN = 4.0 (F is the upper right most element, not including noble gases)

(d) Francium has the lowest electronegativity since it is the lower left most element.
(The book notes that Cesium is the element w/ the lowest EN if you don't include radioactive elements. but really, where's the fun in that?!)

(39) Which element is the most electronegative element in each of these sets?

(a) Na, Mg, K, Ca

(b) P, S, As, Se

(c) Be, B, C, Si

(d) Zn, Ge, Ga, As

EN increases ↗ on p. table

Na	Mg
K	Ca

← Mg! (a)

P	S
As	Se

← Sulfur! (b)

Be	B	C
		Si

← Carbon! (c)

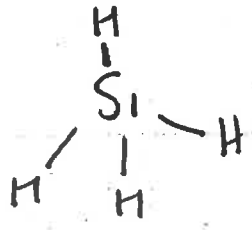
Zn	Ga	Ge	As

↑ Arsenic! (d)

(47) Lewis Dot Structures! (a) SiH_4 (b) CO (c) SF_2 (d) H_2SO_4 (e) ClO_2^{-1} (f) NH_2OH

(a) SiH_4

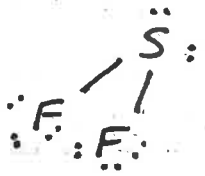
$4+4(1)=8$ v.e.



(b) CO

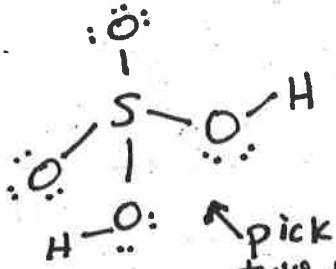
$4+6=10$ v.e.

(c) SF_2 $6+2(7)=20$ v.e.



(d) H_2SO_4

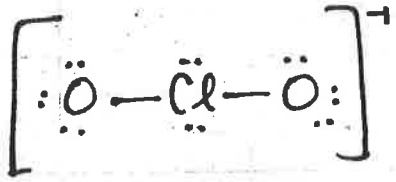
$2(1)+6+4(6)=32$ v.e.



Sulfuric acid! ↖ pick any two oxygens to attach H to.

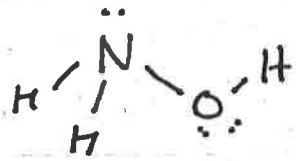
(e) ClO_2^{-1}

$7+2(6)+1=20$ v.e.



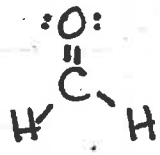
(f) NH_2OH

$5+2(1)+6+1=14$ v.e.



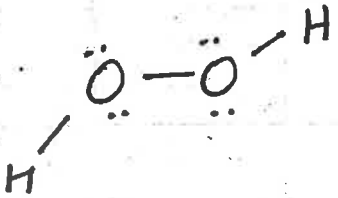
(48) (a) H_2CO

12 v.e.



(formaldehyde)

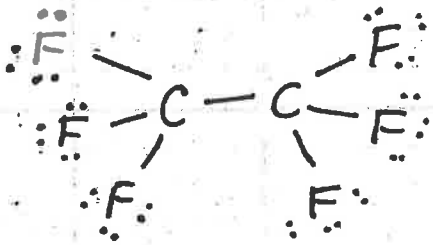
(b) H_2O_2



hydrogen peroxide

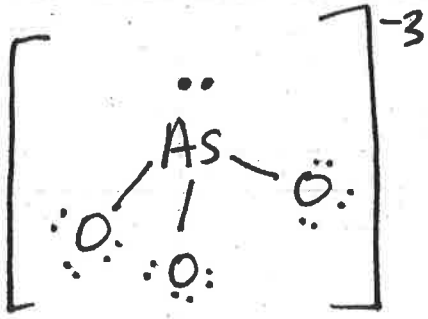
14 v.e.

(c) C_2F_6



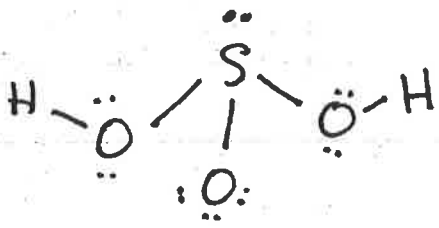
(d) AsO_3^{-3}

$5+3(6)+3=26$ v.e.



(e) H_2SO_3

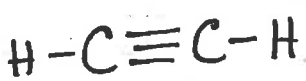
$2+6+3(6)=26$ v.e.



(pick any two oxygens to attach H to)

(f) C_2H_2

$2(4)+2(1)=10$ v.e.



(56) predict ordering of N–O bond lengths in NO^{+1}, NO_2^{-1}, NO_3^{-1}

NO^{+1}

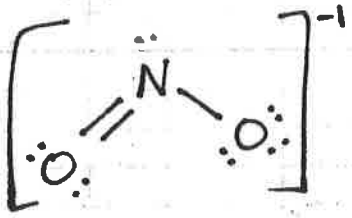
$5+6-1=10$ v.e.

$[:N\equiv O:]^{+1}$

Triple bond. This N–O bond will be the shortest

NO_2^{-1}

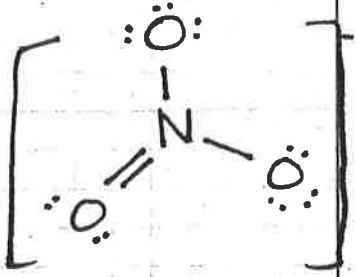
$5+2(6)+1=18$ v.e.



has resonance, so ≈ "1.5" bond on each side, between a single bond and double bond in length

NO_3^{-1}

$5+6(3)+1=24$ v.e.



has resonance overall 3 N–O bonds, so ≈ "1⅓" bond for each N–O bond. between a single and double bond in length, but close to a single bond.

(Single bonds are longest, Triple bonds are shortest)

(61) Chlorine oxides have one chlorine atom bonded to one or more oxygen atoms.

(for example, ClO, ClO_2, ...)

will have the longest N–O bonds.

No, none of the chlorine oxides will obey the octet rule, since we are only considering ones with one chlorine atom.

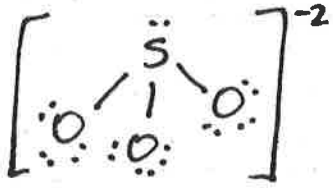
Cl has 7 valence e⁻, and oxygens each have 6 v.e. so the oxygen(s) will always provide an even number of valence electrons, but the single Chlorine atom will make the total number of valence electrons an odd number.

With an odd number of valence electrons, none of them can follow the octet rule for every atom in the molecule.

(63) Lewis Dots!

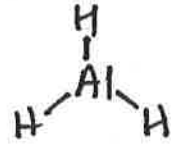
(a) SO_3^{-2}

$6+3(6)+2=26$ v.e.



(b) AlH_3

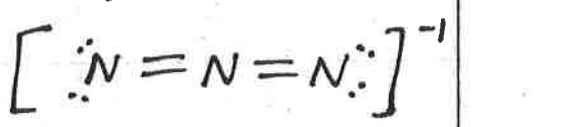
$3+3(1)=6$ v.e.



↖ AlH_3 doesn't obey the octet rule; it only has 6 v.e. total.

(c) N_3^{-1}

$3(5)+1=15$ v.e.

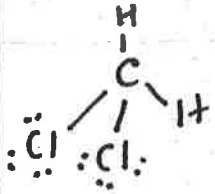


↖ could also do it with a triple-single or single-triple, but this one has formal charges closest to zero.

note: though Al is a metal, it has a high enough electronegativity to be able to bond covalently, as it is here.

(d) CH_2Cl_2

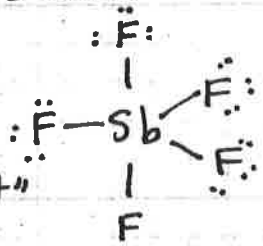
$4+2(1)+2(7)=20$ v.e.



(e) SbF_5 $5+5(7)=40$ v.e.

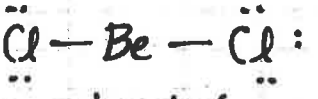
SbF_5 does not follow the octet rule.

Sb has an "expanded octet" with 10 v.e. around it, not 8.



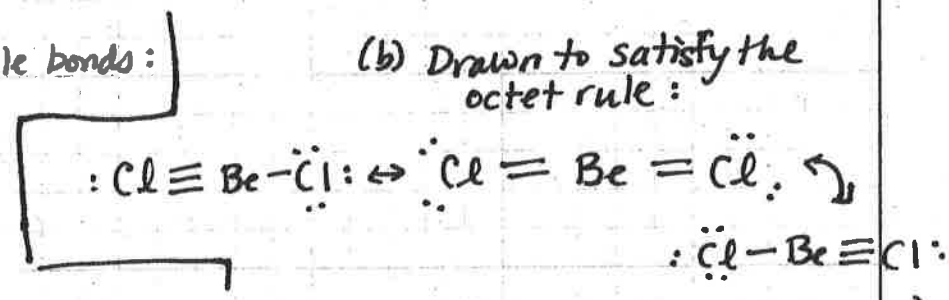
(65) $BeCl_2$ (g) 16 v.e.

(a) Drawn with single bonds:



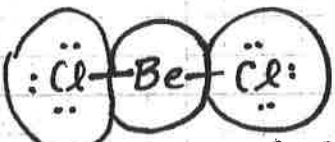
↖ does not satisfy octet rule.

(b) Drawn to satisfy the octet rule:



Formal charges:

(FC = # ve of element − # of e- "owned" by each atom in structure)

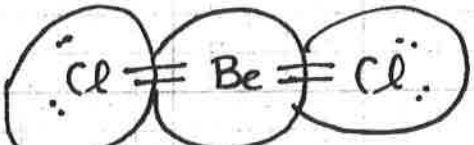


7-7=0 7-7=0

2-2=0

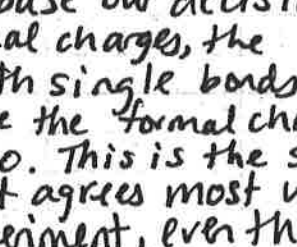
↖ all formal charges are zero

If we base our decision on formal charges, the $BeCl_2$ with single bonds is best since the formal charges are all zero. This is the structure that agrees most with experiment, even though it does not meet the octet rule.



7-6=+1 2-4=-2 7-6=+1

formal charges of +1, −2, +1



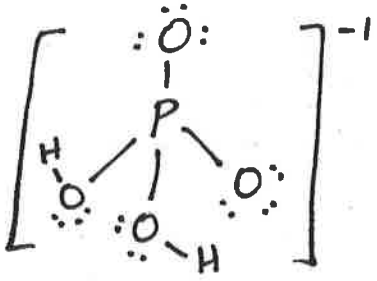
+2 −2 0

0 −2 +2

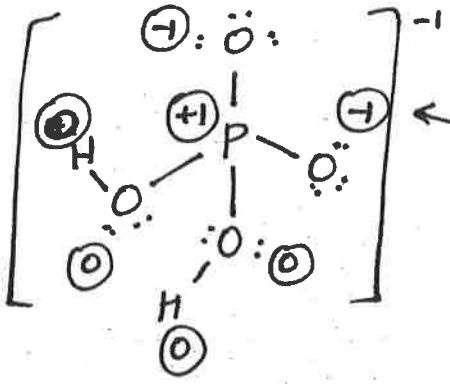
(68) $H_2PO_4^{-1}$

$2(1)+5+4(6)+1 = 32$ v.e.

(a) If we draw it to satisfy the octet rule:

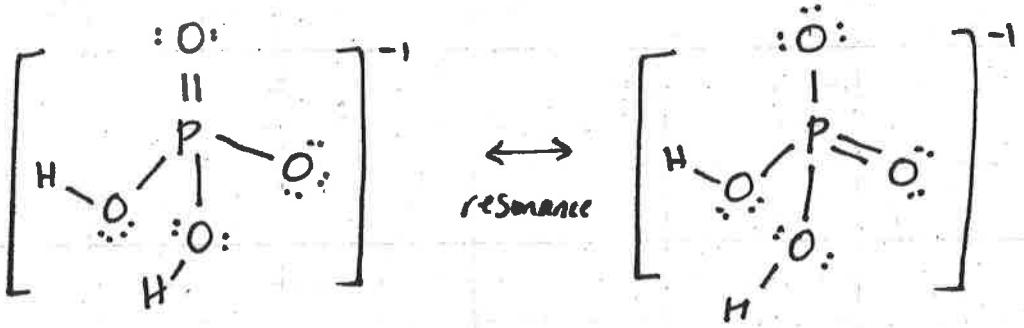


↑ pick any two oxygens for the H's.

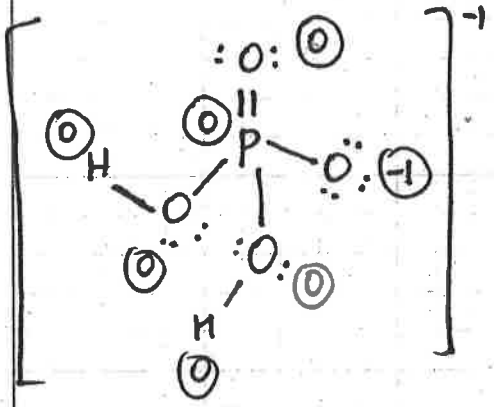


← Redrawn, showing formal charges

(b) If we draw it to achieve the best formal charges:



↑ Redrawing to show formal charges below:

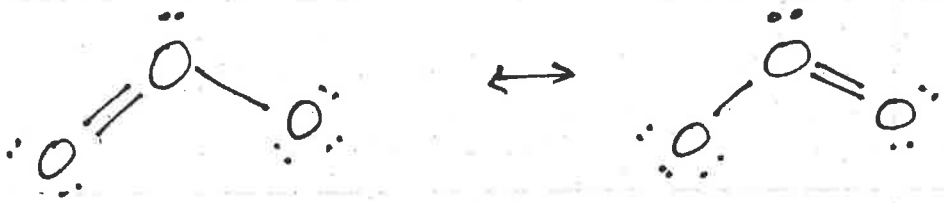


← or really, with resonance, it is more like a $-1/2$ on each oxygen, and still a zero on the phosphorus. This (b) has "better" formal charges than the structure in (a) since the charges are close to zero in (b)

The debate rages on....

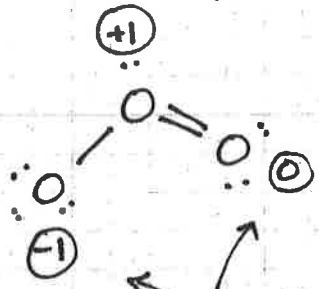
(90) Formal charge on central atom in O_3

a) $3(6) = 18$ v.e.



The central atom "owns" the whole lone pair (2 e-)
plus half of each bond $(\frac{1}{2})$(6 e- in bonds) = (3 e-)
So it "owns" 5 v.e.
but it "brought" 6 v.e. to the molecule since it is in column 6.

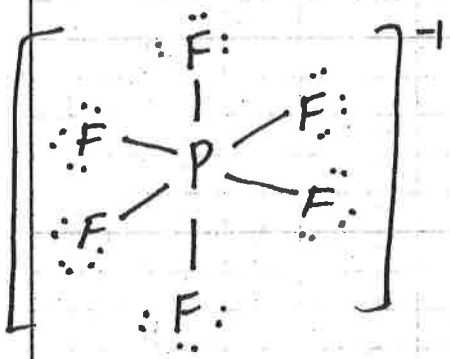
$$FC = 6 - 5 = +1$$



or with resonance, more like a $-\frac{1}{2}$ on each of these.

(b) Phosphorus in PF_6^{-1}

$5 + 6(7) + 1 = 48$ v.e.



Phosphorus contributes 5 v.e. since group 5.
and P "owns" half of each bond,
so $\frac{1}{2}$ (12 bonded e-) = 6 v.e.

$$5 - 6 = -1 = FC$$