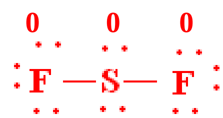
REVIEW QUESTIONS
Chapter 9

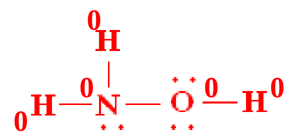
1. Draw Lewis structures for each of the following structures and assign formal charges to each atom:

   a) SF_2 **20 electrons**

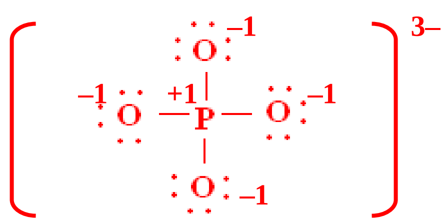
   b) NH_2OH (N and O are bonded to one another)

   14 electrons

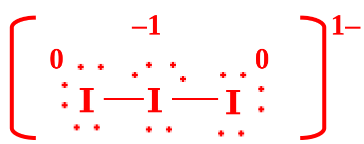
   c) PO_4^{3-}

   32 electrons

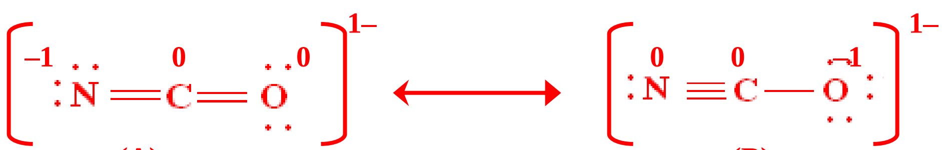
   d) I_3^-

   22 electrons

2. Draw two possible resonance structures for the isocyanate ion (NCO^-) and using formal charges determine which structure has greater contribution to the resonance hybrid.

16 electrons



(A) **(B)**

Both structures possess the same magnitude of formal charges. However, structure B has greater contribution to the resonance hybrid since the negative charge is carried by oxygen which is more electronegative than nitrogen.

3. Arrange the bonds in each of the following sets in order if increasing polarityCl:

a)	C — S	B — F	N — O	**C — S < N — O < B — F**
ΔEN	**0.03**	**1.94**	**0.40**	
b)	O — Cl	S — Br	C — P	**O — Cl < C — P < S — Br**
ΔEN	**0.28**	**0.38**	**0.36**	

4. Classify each of the following bonds as ionic, polar covalent or non-polar covalent:

a) B — Cl **polar covalent** **ΔEN= 1.12**

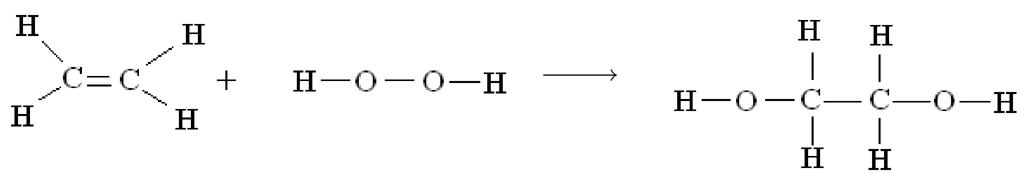
b) Mg — Br **polar covalent** **ΔEN= 1.65**

c) Cl — Cl **non-polar covalent** **ΔEN= 0**

d) Na — Br **ionic** **ΔEN= 2.03**

5. Use bond energies listed in Table 9.5 in your textbook to find ΔH for the reactions shown below:

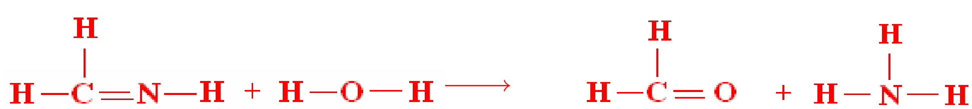
a)



Bond Breaking (BB)		**Bond Forming (BF)**	
C=C	**1 x 602 = 602 kJ**	**C—C**	**1 x 346 = 346 kJ**
O—O	**1 x 142 = 142 kJ**	**C—O**	**2 x 358 = 716 kJ**

ΔH = Σ BB – Σ BF = (602+142) – (346+716) = –318 kJ

b) CH_2NH + H_2O $\longrightarrow$ CH_2O + NH_3



Bond Breaking (BB)		**Bond Forming (BF)**	
C=N	**1 x 615 = 615 kJ**	**C=O**	**1 x 745 = 745 kJ**
O—H	**2 x 459 = 918 kJ**	**N—H**	**2 x 386 = 772 kJ**

ΔH = Σ BB – Σ BF = (615+918) – (745+772) = +16 kJ

6. Use the data provided below to calculate the lattice energy of RbCl. Is this value greater or less than the lattice energy of NaCl? Explain.

   Electron affinity of Cl = –349 kJ/mol
   1^{st} ionization energy of Rb = 403 kJ/mol
   Bond energy of Cl_2 = 242 kJ/mol
   Sublimation energy of Rb= 86.5 kJ/mol
   ΔH_f [RbCl (s)] = –430.5 kJ/mol

   Rb (s) + ½ Cl_2 (g) → RbCl (s)

   This equation can be written as the sum of the following:

Rb (s) → Rb (g)	**sublimation**	**ΔH_1= +86.5 kJ/mol**
Rb (g) → Rb^+ + e^-	**1^{st} ionization energy**	**ΔH_2= +403 kJ/mol**
½ Cl_2 (g) → Cl	**bond energy of Cl_2**	**ΔH_3= +121 kJ/mol**
Cl + e^- → Cl^-	**electron affinity of Cl**	**ΔH_4= –349 kJ/mol**
Rb^+ + Cl^- → RbCl	**lattice energy**	**ΔH_5= ???**
Rb (s) + ½ Cl_2 (g) → RbCl (s)		**ΔH_f = –430.5 kJ/mol**

   $\Delta H_f = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4 + \Delta H_5$

   Lattice energy = $\Delta H_5 = \Delta H_f - (\Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4)$
   = –430.5 – (86.5 + 403 + 121 – 349)
   = –692 kJ

   This value would be expected to be smaller than NaCl. This is because Rb is a larger ion than Na and would be further apart from the anion. Lattice energy is inversely proportional to the distance between the ions.

7. Arrange the following compounds in order of increasing lattice energy:

   NaF CaO CsI

   CsI < NaF < CaO

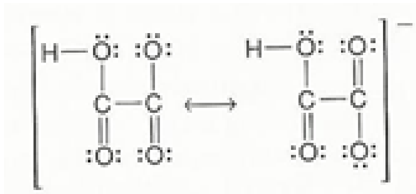
   Calcium has a +2 ion and oxygen has –2 ion, while both NaF and CsI possess +1 and –1 charges. Since lattice energy is directly proportional to the charges, CaO would have the largest value.

   Sodium ion and fluoride ions are smaller than cesium and iodide ions. Since lattice energy is inversely proportional to the size of the ions, CsI would have the lowest value.

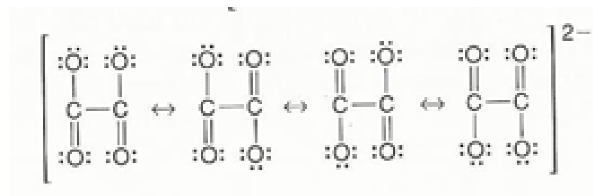
8. Oxalic acid ($H_2C_2O_4$) is a weak acid that can lose two hydrogens to form the following anions:

$HC_2O_4^{-}$ and $C_2O_4^{2-}$

Draw Lewis structures for the two anions above, and comment on the relative strength and length of their C–O bonds.



$HC_2O_4^{-}$



$C_2O_4^{2-}$

In $HC_2O_4^{-}$, since the resonance structures are not equivalent, the C-O bonds are of different length (C=O bond shorter than C–O bond) and different strengths (C=O bond stronger than C–O bond).

In $C_2O_4^{2-}$, since the resonance structures are equivalent, all C-O bonds are of the same length and strength, and have values intermediate to those in $HC_2O_4^{-}$.

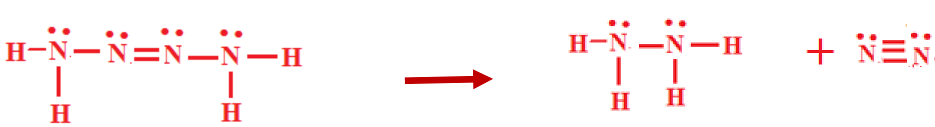
9. Bond energies can be combined with values for other atomic properties to obtain ΔH values that cannot be measured directly. Use bond energy and other data found in your textbook to calculate $\Delta H°_{rxn}$ for the ionic dissociation of chlorine gas:

$$Cl_2\,(g) \rightarrow Cl^{+}\,(g) + Cl^{-}\,(g)$$

Bond dissociation	**$Cl_2\,(g) \rightarrow 2\,Cl\,(g)$**	**ΔH = +243 kJ**
1st ionization energy	**$Cl\,(g) \rightarrow Cl^{+}\,(g) + e^{-}$**	**ΔH = +1251 kJ**
Electron affinity	**$Cl\,(g) + e^{-} \rightarrow Cl^{-}\,(g)$**	**ΔH = –349 kJ**
	$Cl_2\,(g) \rightarrow Cl^{+}\,(g) + Cl^{-}\,(g)$	**ΔH = + 1145 kJ**

10. Tetrazene (N_4H_4) is a thermally unstable nitrogen hydride with the atom sequence shown below. It decomposes above 0°C to form hydrazine (N_2H_4) and nitrogen gas. Draw a Lewis structure for tetrazene and calculate the $\Delta H°_{rxn}$ for its decomposition.

H_2NNNNH_2

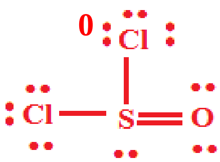


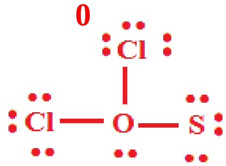
Bond Breaking (BB)			**Bond Forming (BF)**	
1	**N–N**	**163 kJ**	**1 N≡N**	**946 kJ**
1	**N=N**	**418 kJ**		

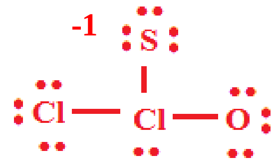
ΔH = Σ BB – Σ BF = 581–946 = –365 kJ

11. Thionyl chloride ($SOCl_2$) can have the 3 skeletal structures shown below. Complete the Lewis structure for each, assign formal charges and determine which structure is the most plausible for this compound.

Cl	Cl	S
Cl S O	Cl O S	Cl Cl O

0  0 0 0

0  0 +1 -1

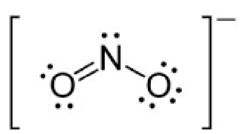
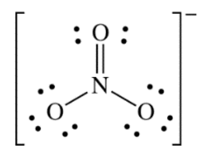
-1  0 +2 -1

Most plausible structure (lowest formal charges)

Not acceptable FC's (S has negative FC)

Not acceptable FC's (very large values)

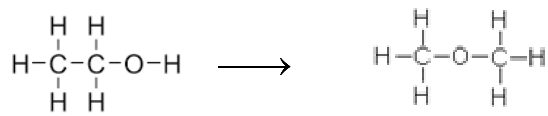
12. Rank the length of the N–O bond length in the following ions:

NO^+	NO_2^-	NO_3^-
		
	(2 resonance structures)	**(3 resonance structures)**

	NO^+	NO_2^-	NO_3^-
Bond order:	**3**	**1.5 (3/2)**	**1.33 (4/3)**
Bond length:	**shortest**	**intermediate**	**longest**

13. Two compounds are isomers if they have the same chemical formula but a different arrangement of atoms. Use bond energies available in Table 9.3 in your test to estimate ΔH_{rxn} for each of the following isomerization reactions and indicate which isomer is more stable.

a)



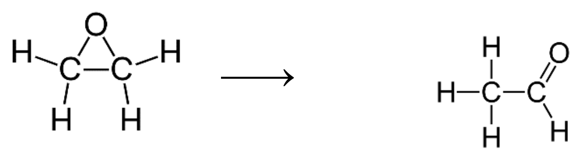
Ethanol Dimethyl ether

Bond Breaking (BB)			**Bond Forming (BF)**		
1	**O–H**	**464 kJ**	**1**	**C–O**	**360 kJ**
1	**C–C**	**347 kJ**	**1**	**C–H**	**414 kJ**

$\Delta H = \Sigma$ BB – Σ BF = 811–774 = 37 kJ

Ethanol is the more stable isomer since it has stronger bonds (larger bond energy values)

b)



Ethylene oxide Acetaldehyde

Bond Breaking (BB)			**Bond Forming (BF)**		
2	**C–O**	**360 kJ**	**1**	**C=O**	**736 kJ**

$\Delta H = \Sigma$ BB – Σ BF = 720–736 = –16 kJ

Acetaldehyde is the more stable isomer since it has stronger bonds (larger bond energy values)