

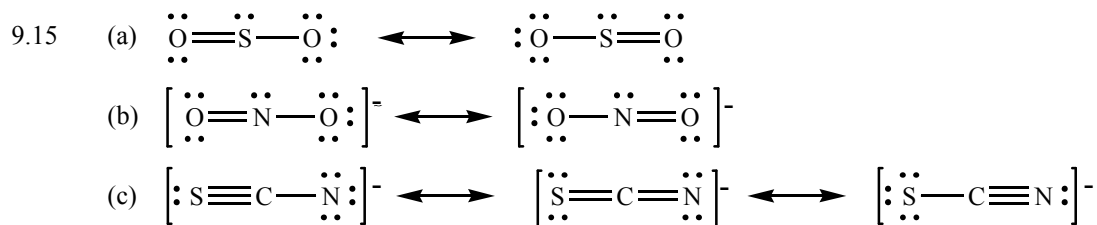
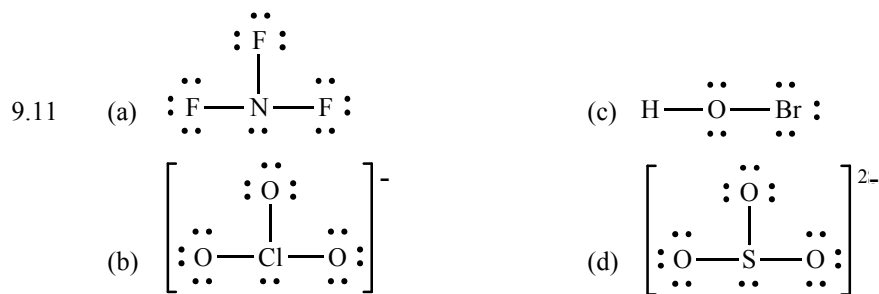
AP Chemistry Chapter 9 Answers - Kotz

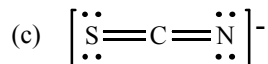
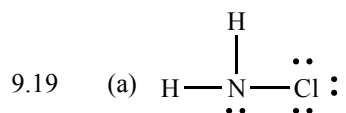
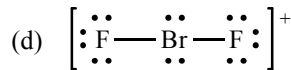
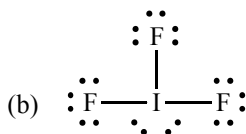
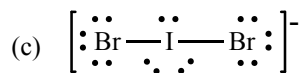
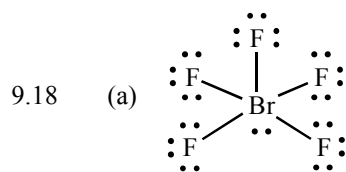
Text Problems: 3, 7, 8, 11, 15, 18, 19, 24, 27, 31, 35, 39, 43, 47, 51, 55, 83, 90

9.3	Group 3A	3 bonds
	Group 4A	4 bonds
	Group 5A	3 bonds
	Group 6A	2 bonds
	Group 7A	1 bond

9.7 RbI < LiI < LiF < CaO

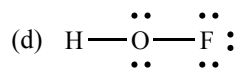
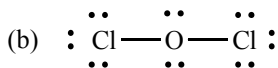
9.8	Li(s) → Li(g)	$\Delta H_f^\circ = +159.37 \text{ kJ/mol}$
	Li(g) → Li ⁺ (g) + e ⁻	IE = +520. kJ/mol
	$\frac{1}{2} \text{F}_2(\text{g}) \rightarrow \text{F}(\text{g})$	$\Delta H_f^\circ = +78.99 \text{ kJ/mol}$
	F(g) + e ⁻ → F ⁻ (g)	EA = -328.0 kJ/mol
	Li ⁺ (g) + F ⁻ (g) → LiF(s)	$\Delta H_{\text{lattice}} = -1037 \text{ kJ/mol}$





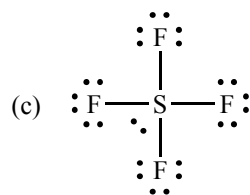
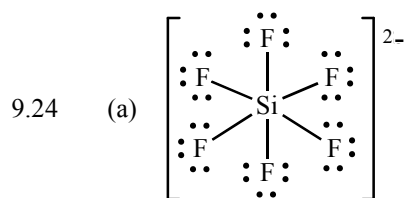
electron pair geometry, tetrahedral
molecular geometry, trigonal pyramidal

electron pair geometry, linear
molecular geometry, linear



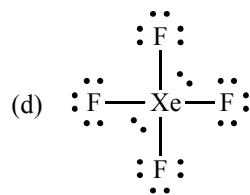
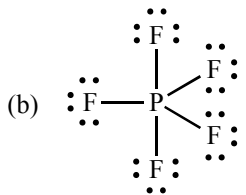
electron pair geometry, tetrahedral
molecular geometry, bent

electron pair geometry, tetrahedral
molecular geometry, bent



electron-pair geometry, octahedral
molecular geometry, octahedral

electron-pair geometry, trigonal bipyramid
molecular geometry, seesaw

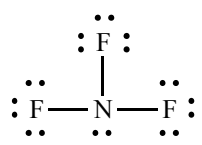
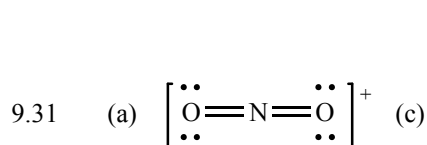


electron-pair geometry, trigonal bipyramid
molecular geometry, trigonal bipyramid

electron-pair geometry, octahedral
molecular geometry, square planar

9.27 angle 1 = 120°; angle 2 = 109°; angle 3 = 120°; angle 4 = 109°; angle 5 = 109°

The —CH₂—CH(NH₂)—CO₂H chain cannot be linear because the first two carbon atoms have bond angles of 109° (with their connecting atoms) and the third carbon has a 120° bond angle with the C and O on either side.



$O = 6 - 4 - \frac{1}{2}(4) = 0$

$F = 7 - 6 - \frac{1}{2}(2) = 0$

$$N = 5 - 0 - \frac{1}{2}(8) = 1$$

$$N = 5 - 2 - \frac{1}{2}(6) = 0$$

- 9.35 (a) The C—H and C=O bonds are polar; the C—C and C=C bonds are nonpolar
 (b) The C=O bond is the most polar bond; the O atom is the more negative atom



$N = 5 - 2 - \frac{1}{2}(6) = 0$	$N = 5 - 4 - \frac{1}{2}(4) = -1$	$N = 5 - 6 - \frac{1}{2}(2) = -2$
$N = 5 - 0 - \frac{1}{2}(8) = 1$	$N = 5 - 0 - \frac{1}{2}(8) = 1$	$N = 5 - 0 - \frac{1}{2}(8) = 1$
$O = 6 - 6 - \frac{1}{2}(2) = -1$	$O = 6 - 4 - \frac{1}{2}(4) = 0$	$O = 6 - 2 - \frac{1}{2}(6) = 1$

- (c) The first resonance structure is most reasonable (the most electronegative element, oxygen, has a negative formal charge).

9.43 Molecule	$\Delta\chi$ for bond
H ₂ O	O—H = 3.5 – 2.1 = 1.4
NH ₃	N—H = 3.0 – 2.1 = 0.9
CO ₂	O—C = 3.5 – 2.5 = 1.0
ClF	F—Cl = 4.0 – 3.0 = 1.0
CCl ₄	Cl—C = 3.0 – 2.5 = 0.5

- (i) The bonds are most polar in H₂O.
 (ii) CO₂ and CCl₄ are nonpolar molecules.
 (iii) The F atom in ClF is more negatively charged.

9.47	(a) H ₂ CO	two carbon-hydrogen single bonds	bond order = 1
		one carbon-oxygen double bond	bond order = 2
	(b) SO ₃ ²⁻	three sulfur-oxygen single bonds	bond order = 1
	(c) NO ₂ ⁺	two nitrogen-oxygen double bonds	bond order = 2
	(d) NOCl	(assume atoms are connected N—O—Cl)	
		one nitrogen-oxygen double bond	bond order = 2
		one oxygen-chlorine single bond	bond order = 1

9.51	NO ₂ ⁺	two NO double bonds	NO bond order = 2
	NO ₂ ⁻	one NO single bond, one NO double bond	
		two resonance structures	NO bond order = $\frac{3}{2}$
	NO ₃ ⁻	two NO single bonds, one NO triple bond	
		three resonance structures	NO bond order = $\frac{4}{3}$

NO₃⁻ has the longest NO bonds (lowest bond order)

NO₂⁺ has the shortest NO bonds (highest bond order)

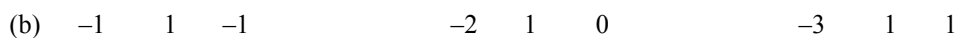
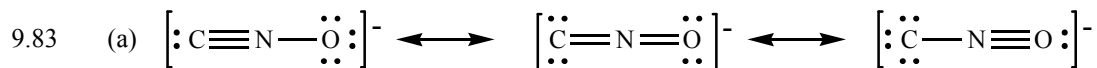
9.55 One C=C double bond and one H—H single bond are broken in the reaction

One C—C single bond and two C—H single bonds are formed in the reaction

$$\Delta H^{\circ}_{\text{rxn}} = (1 \text{ mol} \cdot D_{\text{C}=\text{C}} + 1 \text{ mol} \cdot D_{\text{H}-\text{H}}) - (1 \text{ mol} \cdot D_{\text{C}-\text{C}} + 2 \text{ mol} \cdot D_{\text{C}-\text{H}})$$

$$\Delta H^{\circ}_{\text{rxn}} = 1 \text{ mol} (610 \text{ kJ/mol}) + 1 \text{ mol} (436 \text{ kJ/mol}) - [1 \text{ mol} (346 \text{ kJ/mol}) + 2 \text{ mol} (413 \text{ kJ/mol})]$$

$$\Delta H^{\circ}_{\text{rxn}} = -126 \text{ kJ}$$



The first resonance structure is most reasonable because the negative formal charge is on the most electronegative atom.

(c) Carbon, the least electronegative element in the ion, has a negative formal charge. In addition, all three resonance structures have an unfavorable charge distribution.

9.90 (a) S: $6 - 4 - \frac{1}{2}(4) = 0$ O: $6 - 4 - \frac{1}{2}(4) = 0$

(b) angle 1 = 109°; angle 2 = 109°; angle 3 = 120°

(c) The C=C bonds are shorter than the C—C bonds

(d) The C—O bond is most polar

(e) The molecule is polar

(f) The four C atoms are planar and trigonal, so the ring as a whole is planar.