Questions and Problems

Lewis Dot Symbols

Review Questions

9.1 What is a Lewis dot symbol? To what elements does the symbol mainly apply?

9.2 Use the second member of each group from Group 1A to Group 7A to show that the number of valence electrons on an atom of the element is the same as its group number.

9.3 Without referring to Figure 9.1, write Lewis dot symbols for atoms of the following elements: (a) Be, (b) K, (c) Ca, (d) Ga, (e) O, (f) Br, (g) N, (h) I, (i) As, (j) F.

9.4 Write Lewis dot symbols for the following ions: (a) Li^+, (b) Cl^-, (c) S^{2-}, (d) Sr^{2+}, (e) N^{3-}.

9.5 Write Lewis dot symbols for the following atoms and ions: (a) I, (b) I^-, (c) S, (d) S^{2-}, (e) P, (f) P^{3-}, (g) Na, (h) Na^+, (i) Mg, (j) Mg^{2+}, (k) Al, (l) Al^{3+}, (m) Pb, (n) Pb^{2+}.

The Ionic Bond

Review Questions

9.6 Explain what an ionic bond is.

9.7 Explain how ionization energy and electron affinity determine whether atoms of elements will combine to form ionic compounds.

9.8 Name five metals and five nonmetals that are very likely to form ionic compounds. Write formulas for compounds that might result from the combination of these metals and nonmetals. Name these compounds.

9.9 Name one ionic compound that contains only nonmetallic elements.

9.10 Name one ionic compound that contains a polyatomic cation and a polyatomic anion (see Table 2.3).

9.11 Explain why ions with charges greater than 3 are seldom found in ionic compounds.

9.12 The term "molar mass" was introduced in Chapter 3. What is the advantage of using the term "molar mass" when we discuss ionic compounds?

9.13 In which of the following states would NaCl be electrically conducting? (a) solid, (b) molten (that is, melted), (c) dissolved in water. Explain your answers.

9.14 Beryllium forms a compound with chlorine that has the empirical formula $BeCl_2$. How would you determine whether it is an ionic compound? (The compound is not soluble in water.)

Problems

9.15 An ionic bond is formed between a cation A^+ and an anion B^-. How would the energy of the ionic bond [see Equation (9.2)] be affected by the following changes? (a) doubling the radius of A^+, (b) tripling the charge on A^+, (c) doubling the charges on A^+ and B^-, (d) decreasing the radii of A^+ and B^- to half their original values.

9.16 Give the empirical formulas and names of the compounds formed from the following pairs of ions: (a) Rb^+ and I^-, (b) Cs^+ and SO_4^{2-}, (c) Sr^{2+} and N^{3-}, (d) Al^{3+} and S^{2-}.

9.17 Use Lewis dot symbols to show the transfer of electrons between the following atoms to form cations and anions: (a) Na and F, (b) K and S, (c) Ba and O, (d) Al and N.

9.18 Write the Lewis dot symbols of the reactants and products in the following reactions. (First balance the equations.)

(a) $Sr + Se \longrightarrow SrSe$

(b) $Ca + H_2 \longrightarrow CaH_2$

(c) $Li + N_2 \longrightarrow Li_3N$

(d) $Al + S \longrightarrow Al_2S_3$

9.19 For each of the following pairs of elements, state whether the binary compound they form is likely to be ionic or covalent. Write the empirical formula and name of the compound: (a) I and Cl, (b) Mg and F.

9.20 For each of the following pairs of elements, state whether the binary compound they form is likely to be ionic or covalent. Write the empirical formula and name of the compound: (a) B and F, (b) K and Br.

Both of these compounds have the same numbers and types of atoms. Explain the difference in their boiling points.

Which member of each of the following pairs of substances would you expect to have a higher boiling point? (a) O_2 and Cl_2, (b) SO_2 and CO_2, (c) HF and HI

Which substance in each of the following pairs would you expect to have the higher boiling point? Explain why. (a) Ne or Xe, (b) CO_2 or CS_2, (c) CH_4 or Cl_2, (d) F_2 or LiF, (e) NH_3 or PH_3

Explain in terms of intermolecular forces why (a) NH_3 has a higher boiling point than CH_4 and (b) KCl has a higher melting point than I_2.

What kind of attractive forces must be overcome in order to (a) melt ice, (b) boil molecular bromine, (c) melt solid iodine, and (d) dissociate F_2 into F atoms?

9.26 Calculate the lattice energy of calcium chloride given that the heat of sublimation of Ca is 121 kJ/mol and $\Delta H_f^\circ(CaCl_2) = -795$ kJ/mol. (See Tables 8.2 and 8.3 for other data.)

The Covalent Bond

Review Questions

9.27 What is Lewis's contribution to our understanding of the covalent bond?

9.28 Use an example to illustrate each of the following terms: lone pairs, Lewis structure, the octet rule, bond length.

9.29 What is the difference between a Lewis dot symbol and a Lewis structure?

9.30 How many lone pairs are on the underlined atoms in these compounds? HBr, H_2S, CH_4

9.31 Compare single, double, and triple bonds in a molecule, and give an example of each. For the same bonding atoms, how does the bond length change from single bond to triple bond?

9.32 Compare the properties of ionic compounds and covalent compounds.

Electronegativity and Bond Type

Review Questions

9.33 Define electronegativity, and explain the difference between electronegativity and electron affinity. Describe in general how the electronegativities of the elements change according to position in the periodic table.

9.34 What is a polar covalent bond? Name two compounds that contain one or more polar covalent bonds.

Problems

9.35 List the following bonds in order of increasing ionic character: the lithium-to-fluorine bond in LiF, the potassium-to-oxygen bond in K_2O, the nitrogen-to-nitrogen bond in N_2, the sulfur-to-oxygen bond in SO_2, the chlorine-to-fluorine bond in ClF_3.

9.36 Arrange the following bonds in order of increasing ionic character: carbon to hydrogen, fluorine to hydrogen, bromine to hydrogen, sodium to chlorine, potassium to fluorine, lithium to chlorine.

9.37 Four atoms are arbitrarily labeled D, E, F, and G. Their electronegativities are as follows: D = 3.8, E = 3.3, F = 2.8, and G = 1.3. If the atoms of these elements form the molecules DE, DG, EG, and DF, how would you arrange these molecules in order of increasing covalent bond character?

9.38 List the following bonds in order of increasing ionic character: cesium to fluorine, chlorine to chlorine, bromine to chlorine, silicon to carbon.

9.39 Classify the following bonds as ionic, polar covalent, or covalent, and give your reasons: (a) the CC bond in H_3CCH_3, (b) the KI bond in KI, (c) the NB bond in H_3NBCl_3, (d) the CF bond in CF_4.

9.40 Classify the following bonds as ionic, polar covalent, or covalent, and give your reasons: (a) the SiSi bond in $Cl_3SiSiCl_3$, (b) the SiCl bond in $Cl_3SiSiCl_3$, (c) the CaF bond in CaF_2, (d) the NH bond in NH_3.

Lewis Structure and the Octet Rule

Review Questions

9.41 Summarize the essential features of the Lewis octet rule. The octet rule applies mainly to the second-period elements. Explain.

9.42 Explain the concept of formal charge. Do formal charges represent actual separation of charges?

Problems

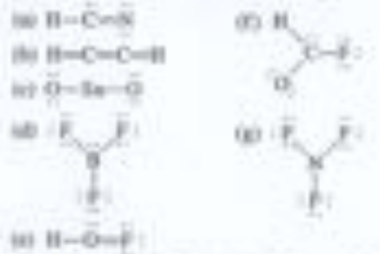
9.43 Write Lewis structures for the following molecules and ions: (a) NCl_3, (b) OCS, (c) H_2O_2, (d) CH_3COO^-, (e) CN^-, (f) $CH_3CH_2NH_3^+$.

9.44 Write Lewis structures for the following molecules and ions: (a) OF_2, (b) N_2F_2, (c) Si_2H_6, (d) OH^-, (e) CH_2ClCOO^-, (f) $CH_3NH_3^+$.

9.45 Write Lewis structures for the following molecules: (a) ICl, (b) PH_3, (c) P_4 (each P is bonded to three other P atoms), (d) H_2S, (e) N_2H_4, (f) $HClO_3$, (g) $COBr_2$ (C is bonded to O and Br atoms).

9.46 Write Lewis structures for the following ions: (a) O_2^{2-}, (b) C_2^{2-}, (c) NO^+, (d) NH_4^+. Show formal charges.

9.47 The following Lewis structures for (a) HCN, (b) C_2H_2, (c) SnO_2, (d) BF_3, (e) HOF, (f) HCOF, and (g) NF_3 are incorrect. Explain what is wrong with each one and give a correct structure for the molecule. (Relative positions of atoms are shown correctly.)



9.48 The skeletal structure of acetic acid shown below is correct, but some of the bonds are wrong. (a) Identify the incorrect bonds and explain what is wrong with them. (b) Write the correct Lewis structure for acetic acid.

The Concept of Resonance

Review Questions

9.47 Define bond length, resonance, and resonance structure. What are the rules for writing resonance structures?

9.48 Is it possible to "map" a resonance structure of a compound for study? Explain.

Problems

9.49 Write Lewis structures for the following species, including all resonance forms, and show formal charges: (a) HCO_2^-, (b) CH_3NO_2. Relative positions of the atoms are as follows:

9.50 Draw three resonance structures for the chlorate ion, ClO_3^-. Show formal charges.

9.51 Write three resonance structures for hydrazoic acid, HN_3. The atomic arrangement is HNNN. Show formal charges.

9.52 Draw two resonance structures for diazomethane, CH_2N_2. Show formal charges. The skeletal structure of the molecule is

9.53 Draw three resonance structures for the molecule N_2O. (atomic arrangement is ONN).) Show formal charges.

9.54 Draw three reasonable resonance structures for the OCN^- ion. Show formal charges.

Exceptions to the Octet Rule

Review Questions

9.55 Why does the octet rule not hold for many compounds containing elements in the third period of the periodic table and beyond?

9.56 Give three examples of compounds that do not satisfy the octet rule. Write a Lewis structure for each.

9.57 Because fluorine has seven valence electrons ($2s^2 2p^5$), seven covalent bonds in principle could form around the atom. Such a compound might be FH_7 or FCl_7. These compounds have never been prepared. Why?

9.58 What is a coordinate covalent bond? Is it different from a normal covalent bond?

Problems

9.59 The AlI_3 molecule has an incomplete octet around Al. Draw three resonance structures of the molecule in which the octet rule is satisfied for both the Al and the I atoms. Show formal charges.

9.60 In the vapor phase, beryllium chloride consists of discrete $BeCl_2$ molecules. Is the octet rule satisfied for Be in this compound? If not, can you form an octet around Be by drawing another resonance structure? How plausible is this structure?

9.61 Of the noble gases, only Rn, Xe, and Kr are known to form a few compounds with O and/or F. Write Lewis structures for the following molecules: (a) XeF_4, (b) XeF_6, (c) XeF_2, (d) $XeOF_4$, (e) XeO_2F_2. In each case Xe is the central atom.

9.62 Write a Lewis structure for $SbCl_5$. Does this molecule obey the octet rule?

9.63 Write Lewis structures for SeF_4 and SeF_6. Is the octet rule satisfied for Se?

9.64 Write Lewis structures for the reaction

$$AlCl_3 + Cl^- \longrightarrow AlCl_4^-$$

What kind of bond joins Al and Cl in the product?

9.65 Explain why an atom cannot have a permanent dipole moment.

9.66 The bonds in beryllium hydride (BeH_2) molecules are polar, and yet the dipole moment of the molecule is zero. Explain.

Problems

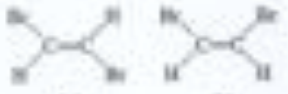
10.19 Referring to Table 10.3, arrange the following molecules in order of increasing dipole moment: H_2O, H_2S, H_2Te, H_2Se.

10.20 The dipole moments of the hydrogen halides decrease from HF to HI (see Table 10.3). Explain this trend.

10.21 List the following molecules in order of increasing dipole moment: H_2O, CBr_4, H_2S, HF, NH_3, CO_2.

10.22 Does the molecule OCS have a higher or lower dipole moment than CS_2?

10.23 Which of the following molecules has a higher dipole moment?



(a) (b)

Key Words

Antibonding molecular orbital, p. 440
Bond order, p. 444
Bonding molecular orbital, p. 440
Delocalized molecular orbital, p. 449
Dipole moment (μ), p. 420
Diamagnetic molecule, p. 445
Hybrid orbital, p. 428
Hybridization, p. 428
Molecular orbital, p. 440
Nonpolar molecule, p. 421
Pi bond (π bond), p. 437
Pi molecular orbital, p. 443
Polar molecule, p. 421
Sigma bond (σ bond), p. 437
Sigma molecular orbital, p. 441
Valence shell, p. 416
Valence-shell electron pair repulsion (VSEPR) model, p. 410

Electronic Homework Problems

The following problems are available at www.aris.mhhe.com (assigned by your instructor as electronic homework). Quantum Tutor problems are also available at the same site.

ARIS Problems. 10.7, 10.8, 10.9, 10.10, 10.12, 10.14, 10.21, 10.24, 10.33, 10.35, 10.36, 10.38, 10.41, 10.54, 10.55, 10.58, 10.60, 10.66, 10.69, 10.70, 10.73, 10.74, 10.76, 10.78, 10.81, 10.82, 10.85, 10.89, 10.99, 10.101, 10.104, 10.105, 10.109.

Quantum Tutor Problems. 10.7, 10.8, 10.9, 10.10, 10.11, 10.12, 10.14, 10.70, 10.73, 10.74, 10.75, 10.79, 10.81, 10.99, 10.108.

Questions and Problems

Molecular Geometry

Review Questions

10.1 How is the geometry of a molecule defined and why is the study of molecular geometry important?

10.2 Sketch the shape of a linear triatomic molecule, a trigonal planar molecule containing four atoms, a tetrahedral molecule, a trigonal bipyramidal molecule, and an octahedral molecule. Give the bond angles in each case.

10.3 How many atoms are directly bonded to the central atom in a tetrahedral molecule, a trigonal bipyramidal molecule, and an octahedral molecule?

10.4 Discuss the basic features of the VSEPR model. Explain why the magnitude of repulsion decreases in the following order: lone pair–lone pair > lone pair–bonding pair > bonding pair–bonding pair.

10.5 In the trigonal bipyramidal arrangement, why does a lone pair occupy an equatorial position rather than an axial position?

10.6 The geometry of CH_4 could be square planar, with the four H atoms at the corners of a square and the C atom at the center of the square. Sketch this geometry and compare its stability with that of a tetrahedral CH_4 molecule.

Problems

10.7 Predict the geometries of the following species using the VSEPR method: (a) PCl_3, (b) $CHCl_3$, (c) SiH_4, (d) $TeCl_4$.

10.8 Predict the geometries of the following species: (a) $AlCl_3$, (b) $ZnCl_2$, (c) $ZnCl_4^{2-}$.

10.9 Predict the geometry of the following molecules and ion using the VSEPR model: (a) CBr_4, (b) BCl_3, (c) NF_3, (d) H_2Se, (e) NO_2^-.

10.10 Predict the geometry of the following molecules and ion using the VSEPR model: (a) CH_2I_2, (b) ClF_3, (c) H_2S, (d) SO_3, (e) SO_4^{2-}.

10.11 Predict the geometry of the following molecules using the VSEPR method: (a) $HgBr_2$, (b) N_2O (arrangement of atoms is NNO), (c) SCN^- (arrangement of atoms is SCN).

10.12 Predict the geometries of the following ions: (a) NH_4^+, (b) NH_2^-, (c) CO_3^{2-}, (d) ICl_2^-, (e) ICl_4^-, (f) AlH_4^-, (g) $SnCl_3^-$, (h) H_3O^+, (i) BeF_4^{2-}.

10.13 Describe the geometry around each of the three central atoms in the CH_3COOH molecule.

10.14 Which of the following species are tetrahedral? $SiCl_4$, SnF_4, XeF_4, Cl_2, $CdCl_4^{2-}$

Dipole Moments

Review Questions

10.15 Define dipole moment. What are the units and symbol for dipole moment?

10.16 What is the relationship between the dipole moment and the bond moment? How is it possible for a molecule to have bond moments and yet be nonpolar?

Key Words

Adhesion, p. 469
Amorphous solid, p. 486
Boiling point, p. 493
Closest packing, p. 476
Cohesion, p. 469
Condensation, p. 490
Coordination number, p. 473
Critical pressure (P_c), p. 494
Critical temperature (T_c), p. 494
Crystalline solid, p. 472
Deposition, p. 497
Dipole-dipole forces, p. 463
Dispersion forces, p. 465
Dynamic equilibrium, p. 490
Equilibrium vapor pressure, p. 490
Evaporation, p. 489
Freezing point, p. 495
Glass, p. 486
Hydrogen bond, p. 467
Induced dipole, p. 464
Intermolecular forces, p. 463
Intramolecular forces, p. 463
Ion-dipole forces, p. 463
Melting point, p. 495
Molar heat of fusion (ΔH_{fus}), p. 496
Molar heat of sublimation (ΔH_{sub}), p. 497
Molar heat of vaporization (ΔH_{vap}), p. 490
Phase, p. 462
Phase changes, p. 489
Phase diagram, p. 498
Sublimation, p. 497
Supercooling, p. 497
Surface tension, p. 469
Triple point, p. 498
Unit cell, p. 473
van der Waals forces, p. 463
Vaporization, p. 489
Viscosity, p. 470
X-ray diffraction, p. 480

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Questions and Problems

Intermolecular Forces

Review Questions

11.1 Give an example for each type of intermolecular forces. (a) dipole-dipole interaction, (b) dipole-induced dipole interaction, (c) ion-dipole interaction, (d) dispersion forces, (e) van der Waals forces.

11.2 Explain the term "polarizability." What kind of molecules tend to have high polarizabilities? What is the relationship between polarizability and intermolecular forces?

11.3 Explain the difference between a temporary dipole moment and the permanent dipole moment.

11.4 Give some evidence that all atoms and molecules exert attractive forces on one another.

11.5 What physical properties should you consider in comparing the strength of intermolecular forces in solids and in liquids?

11.6 Which elements can take part in hydrogen bonding? Why is hydrogen unique in this kind of interaction?

Problems

11.7 The compounds Br_2 and ICl have the same number of electrons, yet Br_2 melts at −7.2°C and ICl melts at 27.2°C. Explain.

11.8 If you lived in Alaska, which of the following natural gases would you keep in an outdoor storage tank in winter? Explain why. methane (CH_4), propane (C_3H_8), or butane (C_4H_{10})

11.9 The binary hydrogen compounds of the Group 4A elements and their boiling points are: CH_4, −162°C; SiH_4, −112°C; GeH_4, −88°C; and SnH_4, −52°C. Explain the increase in boiling points from CH_4 to SnH_4.

11.10 List the types of intermolecular forces that exist between molecules (or basic units) in each of the following species: (a) benzene (C_6H_6), (b) CH_3Cl, (c) PF_3, (d) NaCl, (e) CS_2.

11.11 Ammonia is both a donor and an acceptor of hydrogen in hydrogen-bond formation. Draw a diagram showing the hydrogen bonding of an ammonia molecule with two other ammonia molecules.

11.12 Which of the following species are capable of hydrogen-bonding among themselves? (a) C_2H_6, (b) HI, (c) KF, (d) BeH_2, (e) CH_3COOH

11.13 Arrange the following in order of increasing boiling point: RbF, CO_2, CH_3OH, CH_3Br. Explain your reasoning.

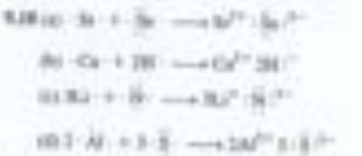
11.14 Diethyl ether has a boiling point of 34.5°C, and 1-butanol has a boiling point of 117°C:

diethyl ether 1-butanol

shielding and the greater nuclear attraction makes the second ionization of He greater than twice the ionization energy of H. **8.100** Z_{eff}: Li (1.28), Na (1.84), K (2.26). Z_{eff}/n: Li (0.640), Na (0.613), K (0.565). Z_{eff} increases as n increases. Thus, Z_{eff}/n remains fairly constant.

CHAPTER 9

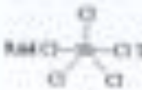
9.16 (a) RbI, rubidium iodide. (b) Cs_2SO_4, cesium sulfate. (c) Sr_3N_2, strontium nitride. (d) Al_2S_3, aluminum sulfide.



9.20 (a) BF_3, covalent. Boron trifluoride. (b) KBr, ionic. Potassium bromide. **9.26** 2195 kJ/mol. **9.36** C—H < Br—H < F—H < Li—Cl < Na—Cl < K—F. **9.38** Cl—Cl < Br—Cl < Si—C < Cs—F. **9.40** (a) Covalent. (b) Polar covalent. (c) Ionic. (d) Polar covalent.

9.62 Not plausible.

9.64 The octet rule is not obeyed.



9.66 Coordinate covalent bond.

9.70 303.0 kJ/mol. **9.72** (a) −2759 kJ/mol. (b) −2855 kJ/mol. **9.74** Ionic: RbCl, KO_2; covalent: PF_5, BrF_3, Cl_2. **9.76** Ionic: NaF, MgF_2, AlF_3; covalent: SiF_4, PF_5, SF_6, ClF_3. **9.78** KF: ionic, high melting point, soluble in water, its melt and solution conduct electricity. C_6H_6: covalent and discrete molecule, low melting point, insoluble in water, does not conduct electricity.

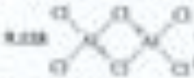
9.82 (a) $AlCl_4^-$. (b) AlF_6^{3-}. (c) $AlCl_3$. **9.84** CF_2: violates the octet rule; LiO_2: lattice energy too low; $CsCl_2$: second ionization energy too high to produce Cs^{2+}; PI_5: I atom too bulky to fit around P. **9.86** (a) False. (b) True. (c) False. (d) False. **9.88** −67 kJ/mol. **9.90** N_2. **9.92** NH_4^+ and CH_4; CO and N_2; $B_3N_3H_6$ and C_6H_6.

9.96 F_3^- violates the octet rule.

9.100 (c) No bond between C and O. (d) Large formal charges.

9.104 (a) −9.2 kJ/mol. (b) −9.2 kJ/mol. **9.106** (a) $^-$:C≡O:$^+$ (b) :N≡O:$^+$ (c) $^-$:C≡N: (d) :N≡N: **9.108** True. **9.110** (a) 114 kJ/mol. (b) Extra electron increases repulsion between F atoms. **9.112** Lone pair on C and negative formal charge on C.

9.116 Violates the octet rule.



The arrows indicate coordinate covalent bonds.

9.120 347 kJ/mol.

9.124 O: 3.16; F: 4.37; Cl: 3.48. **9.126** (1) The MgO solid containing Mg^+ and O^- ions would be paramagnetic. (2) The lattice energy would be like NaCl (too low). **9.128** −629 kJ/mol. **9.130** 259 nm. **9.132** (a) −1413.9 kJ/mol and −1937 kJ/mol. (b) 162 L. (c) 11.0 atm. **9.134** The repulsion between lone pairs on adjacent atoms weakens the bond. There are two lone pairs on each O atom in H_2O_2. The repulsion is the greatest; it has the smallest bond enthalpy (about 142 kJ/mol). There is one lone pair on each N atom in N_2H_4; it has the intermediate bond enthalpy (about 193 kJ/mol). There are no lone pairs on the C atoms in C_2H_6; it has the greatest bond enthalpy (about 347 kJ/mol).

CHAPTER 10

10.8 (a) Trigonal planar. (b) Linear. (c) Tetrahedral. **10.10** (a) Tetrahedral. (b) T-shaped. (c) Bent. (d) Trigonal planar. (e) Tetrahedral. **10.12** (a) Tetrahedral. (b) Bent. (c) Trigonal planar. (d) Linear. (e) Square planar. (f) Tetrahedral. (g) Trigonal bipyramidal. (h) Trigonal pyramidal. (i) Tetrahedral. **10.14** $SiCl_4$, CI_4, $CdCl_4^{2-}$. **10.20** Electronegativity decreases from F to I. **10.22** Higher.

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CHAPTER 11

[illegible]

CHAPTER 12

[illegible]