CHAPTERS 7–9 BRINGING IT TOGETHER

Once again you have an opportunity to test your understanding of concepts, your knowledge of scientific terms, and your problem-solving skills. Read through the following questions carefully, and answer each as fully as possible. When necessary, review topics that give you difficulty. When you are able to answer these questions correctly, you are ready to study the next group of chapters.

1. What are the three principal particles that make up the atom? On the atomic mass scale, what are their approximate masses? What are their electrical charges?

2. A beam of green light has a wavelength of 500 nm. What is the frequency of this light? What is the energy, in joules, of one photon of this light? What is the energy, in joules, of one mole of photons of this light? Would blue light have more or less energy per photon than this light?

3. Arrange the following kinds of electromagnetic radiation in order of increasing frequency: X rays, blue light, radio waves, gamma rays, microwaves, red light, infrared light, ultraviolet light.

4. What is a *continuous spectrum*? How does it differ from an *atomic spectrum*?

5. What experimental evidence is there that matter has wavelike properties?

6. What is the difference between a *traveling wave* and a *standing wave*? What is a *node*?

7. How is the energy of an electron related to the number of nodes in its electron wave?

8. What is a *wave function*? What Greek letter is usually used to represent a wave function? What word do we use to refer to an electron wave in an atom?

9. What are the quantum numbers of the electrons in the valence shells of (a) sulfur, (b) strontium, (c) lead, (d) bromine, and (e) boron?

10. If a given shell has $n = 4$, which kinds of subshells (*s, p,* etc.) does it have? What is the maximum number of electrons that could populate this shell?

11. Use the periodic table to predict the electron configurations of (a) tin, (b) germanium, (c) silicon, (d) lead, and (e) nickel.

12. Give the electron configurations of the ions (a) Pb^{2+}, (b) Pb^{4+}, (c) S^{2-}, (d) Fe^{3+}, and (e) Zn^{2+}.

13. What causes an atom, molecule, or ion to be paramagnetic? Which of the ions in the preceding question are paramagnetic? What term describes the magnetic properties of the others?

14. Give the shorthand electron configurations of (a) Ni, (b) Cr, (c) Sr, (d) Sb, and (e) Po.

15. Define *ionization energy* and *electron affinity.* In terms of these properties, which kinds of elements tend to react to form ionic compounds?

16. In general, the second ionization energy of an atom is larger than the first, the third is larger than the second, and so on. Why?

17. Which of the following elements has the largest difference between its second and third ionization energy? Explain your choice.

(a) Li (b) Be (c) B (d) C

18. Which of the following processes are endothermic?

(a) $P^-(g) + e^- \longrightarrow P^{2-}(g)$

(b) $Fe^{3+}(g) + e^- \longrightarrow Fe^{2+}(g)$

(c) $Cl(g) + e^- \longrightarrow Cl^-(g)$

(d) $S(g) + 2e^- \longrightarrow S^{2-}(g)$

19. Sketch the shape of (a) an *s* orbital, (b) a *p* orbital, and (c) the $3d_{xz}$ orbital.

20. What is meant by the term *electron density*?

21. Give orbital diagrams for the valence shells of selenium and thallium.

22. Which ion would be larger? (a) Fe^{2+} or Fe^{3+}, (b) O^- or O^{2-}

23. Which of the following pairs of elements would be expected to form ionic compounds? (a) Br and F, (b) H and P, (c) Ca and F.

24. Use Lewis symbols to diagram the reaction of calcium with sulfur to form CaS.

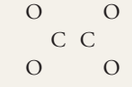
25. Draw Lewis structures for (a) SbH_3, (b) IF_3, (c) $HClO_2$, (d) C_2^{2-}, (e) AsF_5, (f) O_2^{2-}, (g) HCO_3^-, (h) TeF_6, and (i) HNO_3.

26. Use the VSEPR theory to predict the shapes of (a) $SbCl_3$, (b) IF_5, (c) AsH_3, (d) BrF_2, and (e) OF_2.

27. What kinds of hybrid orbitals are used by the central atom in each of the species in the preceding question?

28. Referring to your answers to Questions 25 and 26, which of the following molecules would be nonpolar? SbH_3, IF_3, AsF_5, $SbCl_3$, OF_2

29. The oxalate ion has the following arrangement of atoms.

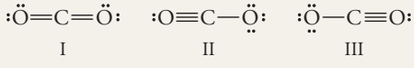


Draw all of its resonance structures.

30. What is meant by the term *overlap of orbitals?*

31. What are *sigma bonds*? What are *pi bonds*? How are sigma and pi bonds used to explain the formation of double and triple bonds?

32. Some resonance structures that can be drawn for carbon dioxide are shown below.



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Assign formal charges to the atoms in these structures. Explain why Structure I is the preferred structure.

33. Ozone, O_3, consists of a chain of three oxygen atoms.

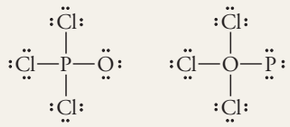
(a) Draw the two resonance structures for ozone that obey the octet rule.

(b) Based on your answer to (a), is the molecule linear or nonlinear?

(c) Assign formal charges to the atoms in the resonance structures you have drawn in part (a).

(d) On the basis of your answers to (b) and (c), explain why ozone is a polar molecule even though it is composed of three atoms that have identical electronegativities.

34. Why, on the basis of formal charges and relative electronegativities, is it more reasonable to expect the structure of $POCl_3$ to be the one on the left rather than the one on the right?



Is either of these the "best" Lewis structure that can be drawn for this molecule?

35. A certain element *X* was found to form three compounds with chlorine having the formulas XCl_2, XCl_4, and XCl_6. One of its oxides has the formula XO_3, and *X* reacts with sodium to form the compound Na_2X.
(a) Is *X* a metal or a nonmetal?
(b) In which group in the periodic table is *X* located?
(c) In which periods in the periodic table could *X* possibly be located?
(d) Draw Lewis structures for XCl_2, XCl_4, XCl_6, and XO_3. (Where possible, follow the octet rule.) Which has multiple bonding?
(e) What do we expect the molecular structures of XCl_2, XCl_4, XCl_6, and XO_3 to be? Which are polar molecules?
(f) The element *X* also forms the oxide XO_2. Draw a Lewis structure for XO_2 that obeys the octet rule.
(g) Assign formal charges to the atoms in the Lewis structures for XO_2 and XO_3 drawn for parts (d) and (f).
(h) What kinds of hybrid orbitals would *X* use for bonding in XCl_4 and XCl_6?
(i) If *X* were to form a compound with aluminum, what would be its formula?
(j) Which compound of *X* would have the more ionic bonds, Na_2X or MgX?
(k) If *X* were in Period 5, what would be the electron configuration of its valence shell?

36. Where in the periodic table are the very reactive metals located? Where are the least reactive ones located?

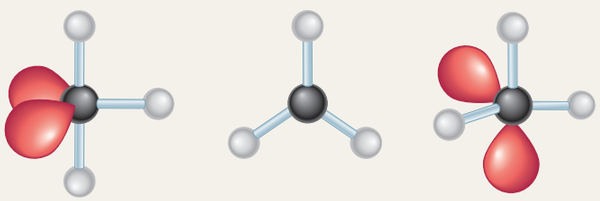
37. What are bonding and antibonding molecular orbitals? How do they differ in shape and energy?

38. What is a *delocalized molecular orbital*? How does molecular orbital theory avoid the concept of resonance?

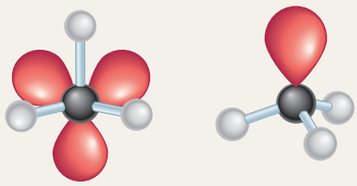
39. Predict the shapes of the following molecules and ions:
(a) PF_3 (c) PF_6^-
(b) PF_4^+ (d) PF_5

40. Which of the substances in the preceding question have a net dipole moment?

41. According to the VSEPR model, which of the following best illustrates the structure of the $AsCl_3^{2-}$ ion?



(a) (b) (c)



(d) (e)

42. Describe how molecular orbital theory explains the bonding in the oxygen molecule.

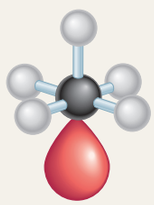
43. Consider the following statements: (1) Fe^{2+} is easily oxidized to Fe^{3+}, and (2) Mn^{2+} is difficult to oxidize to Mn^{3+}. On the basis of the electron configurations of the ions, explain the difference in ease of oxidation.

44. For each of the following pairs of compounds, which has the larger lattice energy: (a) MgO or NaCl, (b) MgO or BeO, (c) NaI or NaF, and (d) MgO or CaS? Explain your choices.

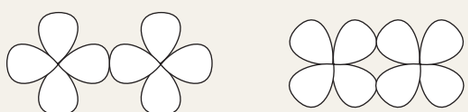
45. The melting point of Al_2O_3 is much higher than the melting point of NaCl. On the basis of lattice energies, explain why this is so.

46. Why is the change in atomic size, going from one element to the next in a period, smaller among the transition elements than among the representative elements?

47. The VSEPR model predicts the structure below for a certain molecule. Which kind of hybrid orbitals does the central atom in the molecule use to form its covalent bonds?

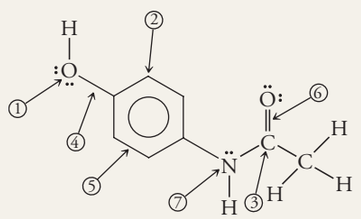


48. Which kind of bond, σ or π, is produced by the overlap of *d* orbitals pictured below?



(a) (b)

49. The following is the chemical structure of acetaminophen, the painkiller in Tylenol.



What kinds of hybrid orbitals are used by atoms 1, 2, 3, and 7? How many σ and π bonds are in bonds 4 and 6? What is the average bond order of bond 5?