BONDING Unit Review (pg. 137 Chapter 3 review)

Part 1

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1. 2, 1, 4, 3	2. C	3. D	4. A	5. C	6. C	7. 2, 1, 4, 3	8. A	9. B	10. D
11. C	12. B	13. B	14. A	15. D					

Part 2

16. PCl5. Instead of one lone pair and three bonding electrons, each of the five electrons in phosphorus bonds with a chlorine atom.

17. PCl5 is a solid because the London forces are stronger due to the increased number of electrons on each molecule.

18. HCl(aq)

19. PCl3 has a trigonal pyramidal shape.

20. (a) Chemical reactivity increases with increasing atomic sizes in Groups 1 and 2.

(b) Chemical reactivity decreases with increasing atomic sizes in Groups 16 and 17.

(c) Chemical reactivity decreases and then increases from left to right along Period 3.

(d) Chemical reactivity does not vary significantly within Group 18.

21. (i) Electronegativity generally decreases as you go down a group.

(ii) Electronegativity generally increases from left to right along periods.

22. Metals react with nonmetals to form ionic compounds. Nonmetals react with other nonmetals to form molecular compounds.

23. eight

24. The electronegativities of the main group metals are lower than the electronegativities of the main group nonmetals.

25.

(a) 4 lone pairs, 0 bonding electrons

(b) 0 lone pairs, 3 bonding electrons

(c) 0 lone pairs, 4 bonding electrons

(d) 1 lone pair, 3 bonding electrons

(e) 3 lone pairs, 1 bonding electron

26.

(a) (2)

(b) (1)

(c) (3)

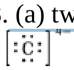
(d) (4)

(e) (2)

27. For a covalent bond to form between two approaching atoms, both atoms must have a valence orbital occupied by a single electron and the orbitals must be able to overlap in space.

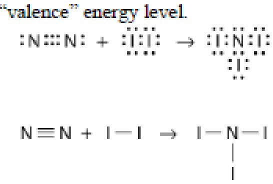
28. (a) two lone pairs (b) one lone pair (c) two lone pairs (d) no lone pairs (e) one lone pair

29.

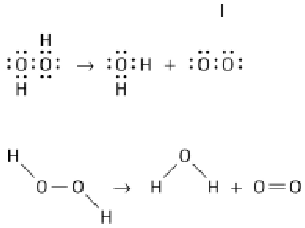


30. The electron arrangement that gives an atom maximum stability is one with eight electrons in the "valence" energy level.

31. (a)

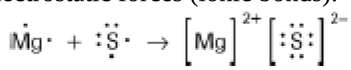


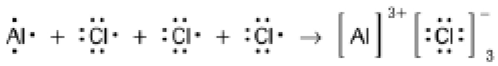
(b)



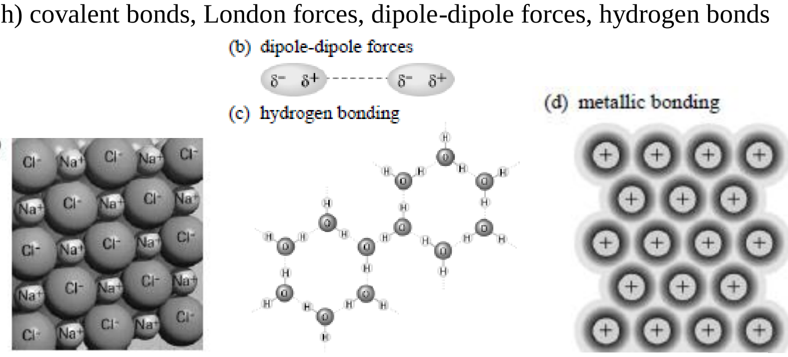
34. Boiling point indicates the number and strength of intermolecular forces. A compound with many strong intermolecular forces (e.g., hydrogen bonds) has a higher boiling point than a

compound with fewer and weaker intermolecular forces (e.g., London forces).
35. Three types of intermolecular forces are hydrogen bonds (e.g., in water), dipole–dipole forces (e.g., in hydrogen sulfide), and London forces (e.g., in hexane).
36. Ionic compounds are relatively hard, brittle crystals of varying solubility that have relatively high melting points, and conduct electricity in liquid and aqueous solution phases.
37. Ionic compounds are composed of cations and anions, formed by electron transfer, that are arranged in crystal lattice structures in simple whole-number ratios and held together by strong electrostatic forces (ionic bonds).

38. (a) 

(b) 

39. (a) Valence shell refers to the outer or highest energy level of electrons.
(b) Bonding pair refers to a pair of electrons shared between atoms in a valence orbital.
(c) Lone pair refers to a pair of electrons in a valence orbital that are not shared with another atom.
(d) Electron pair repulsion assumes that valence orbitals occupied by a pair of electrons are "full" and will repel any other full valence orbital.
40. (i) Draw the Lewis formula for the molecule, including the electron pairs around the central atom.
(ii) Count the number of bonding pairs (bonded atoms) and lone pairs of electrons around the central atom.
(iii) Refer to Table 7, page 24 of the Student Book and use the number of pairs of electrons to predict the shape of the molecule.
41. (a) angular (V-shaped) (b) trigonal planar (c) tetrahedral (d) tetrahedral (e) linear
(f) angular (V-shaped) (g) trigonal pyramidal (h) angular (V-shaped)
42. The numbers in a molecular formula indicate the actual number of atoms of each element in the molecule. The numbers in an ionic formula indicate the ratio of the ions in the ionic crystal.
43. (a) covalent bonds, London forces (b) metallic bonds (c) ionic bonds
(d) covalent bonds, London forces, dipole-dipole forces (e) covalent bonds (in a network)
(f) ionic bonds (g) covalent bonds, London forces, dipole-dipole forces, hydrogen bonds
(h) covalent bonds, London forces, dipole-dipole forces, hydrogen bonds

44. (a) 

45. (a)

2 e−	8 e−		3 e−
8 e−	8 e−	1 e−	8 e−
2 e−	2 e−	2 e−	2 e−
12 p+	19 p+	3 p+	13 p+
Mg	K	Li	Al
magnesium atom	potassium atom	lithium atom	aluminium atom

(b) Since metals have a low electronegativity, their valence electrons are not held tightly by the nucleus, and are free to move through vacant valence orbital spaces, so metals have high conductivity. The malleability of metals is explained by the non-localized nature of the metallic bond. In a metallic bond, atoms can slide over each other, causing the metal to be bendable.
46. (a) :Ö::C::Ö:
(b) A carbon dioxide molecule is linear, with a bond angle of 180°.
(c) Carbon dioxide has two double bonds, each of which is strongly polar. The two bond polarities are exactly opposite and so the resultant molecular dipole is zero, and the molecule is nonpolar.
47. a) An N-Cl bondis not very polar EN difference of 0.2 compared to C-Cl with EN diff. of 0.6
b) A NCl_3 should only be slightly polar because bonds are only slightly polar, CCl_4 is nonpolar
53. a) covalent, London forces b) covalent bonds, L, DD, HB, c) covalent, L, DD, HB d) covalent, L, DD e) ionic f) covalent (in a network)