CO_2(s) has only weak London forces between its molecules, making this a gas under normal conditions.

UNIT 1 REVIEW

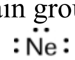
Part 1

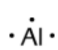
(Page 137)

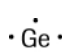
1. 2, 1, 4, 3
2. C
3. D
4. A
5. C
6. C
7. 2, 1, 4, 3
8. A
9. B
10. D
11. C
12. B
13. B
14. A
15. D

Part 2

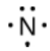
(Pages 138–141)

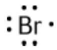
16. PCl_5. Instead of one lone pair and three bonding electrons, each of the five electrons in phosphorus bonds with a chlorine atom.
17. PCl_5 is a solid because the London forces are stronger due to the increased number of electrons on each molecule.
18. HCl(aq)
19. PCl_3 has a trigonal pyramidal shape.
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) 

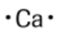
4 lone pairs, 0 bonding electrons
    (b) 

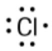
0 lone pairs, 3 bonding electrons
    (c) 

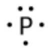
0 lone pairs, 4 bonding electrons

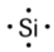
(d)  1 lone pair, 3 bonding electrons

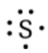
(e)  3 lone pairs, 1 bonding electron

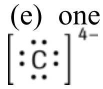
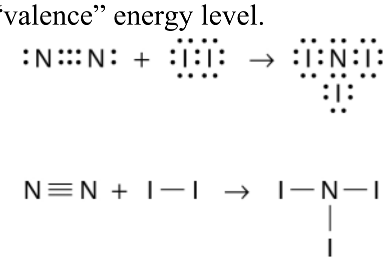
26. (a)  (2)

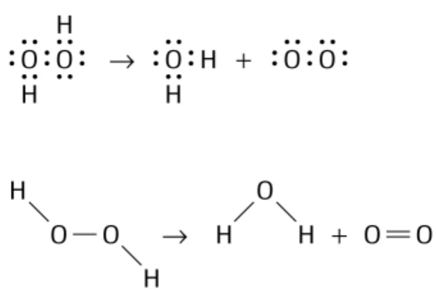
(b)  (1)

(c)  (3)

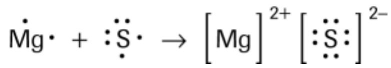
(d)  (4)

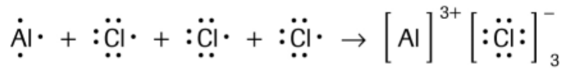
(e)  (2)

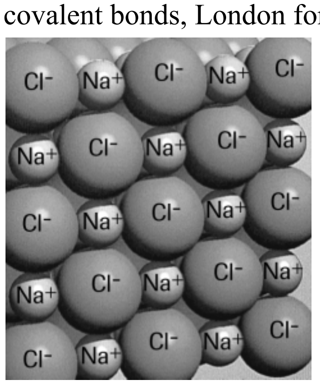
27. For a covalent bond to form between two approaching atoms, both atoms must have a valence orbital occupied by a single electron and the orbitals must be able to overlap in space.
28. (a) two lone pairs
    (b) one lone pair
    (c) two lone pairs
    (d) no lone pairs
    (e) one lone pair
29. 
30. The electron arrangement that gives an atom maximum stability is one with eight electrons in the "valence" energy level.
31. (a) 

(b) 

32. Empirically, molecules of gases such as oxygen and nitrogen were known to be diatomic. Theoretically, an oxygen atom has only six valence electrons and a nitrogen atom has only five. Since theory must be modified to explain observation, the concept of double and triple bonds was introduced to explain the stability of molecular structures of substances such as these, that did not have atoms with full valence shells if single bonding was assumed.
33. Empirical evidence for double and triple bonds includes reaction rate—molecules with multiple bonds react more rapidly than molecules with single bonds. Other evidence indicates that multiple bonds are shorter and stronger than single bonds between the same kinds of atoms.
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).

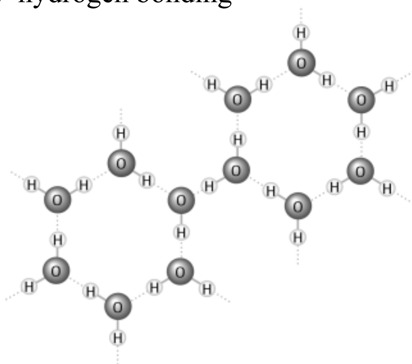
35. Three types of intermolecular forces are hydrogen bonds (e.g., in water), dipole–dipole forces (e.g., in hydrogen sulfide), and London forces (e.g., in hexane).
36. Ionic compounds are relatively hard, brittle crystals of varying solubility that have relatively high melting points, and conduct electricity 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) 

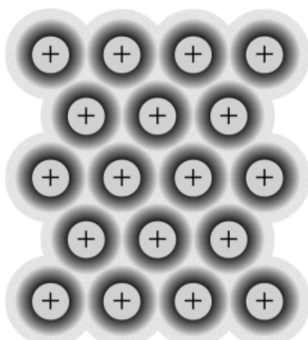
39. (a) Valence shell refers to the outer or highest energy level of electrons.
    (b) Bonding pair refers to a pair of electrons shared between atoms in a valence orbital.
    (c) Lone pair refers to a pair of electrons in a valence orbital that are not shared with another atom.
    (d) Electron pair 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) 

(b) dipole-dipole forces

(c) hydrogen bonding



(d) metallic bonding



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) $:\ddot{O}::C::\ddot{O}:$

(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 bond is not very polar, with an electronegativity difference of 0.2, whereas a C—Cl bond is significantly more polar, with an electronegativity difference of 0.6.

(b) A molecule of NCl_3(l) should be only slightly polar because the bonds are only slightly polar. A molecule of CCl_4(l) should be nonpolar because it is symmetrical, so the bond dipoles total to a zero molecular dipole, cancelling any molecular polarity.

48. (a) BeH_2 has no lone pairs and two bonding pairs. Its shape is linear and the bonds are polar; therefore, it is nonpolar because the bond dipoles cancel. H_2S has two lone pairs and two

bonding pairs. It is angular with polar bonds, and therefore it is polar because the bond dipoles do not cancel.

(b) Boron in BH_3 has no lone pairs of electrons to influence its shape, whereas nitrogen in NH_3 does. Therefore three groups of electrons in BH_3 repel each other to produce a trigonal planar shape. The extra lone pair on the N atom means that four groups of electrons repel each other to produce a tetrahedral shape of electrons and a trigonal pyramidal shape of the three bonds.

(c) LiH is an ionic crystal given its high melting point, whereas HF is a molecular compound given its low melting point. The London forces between individual HF molecules are more easily broken than all the ionic bonds in the LiH crystal lattice.

49. All three compounds have London forces and dipole–dipole forces acting on them. Since Br is found in all three compounds, they all have similar polarity. The strength of their London forces varies. $CH_3Br(g)$ has 50 electrons, $C_2H_5Br(l)$ has 64 electrons, and $C_3H_7Br(l)$ has 80 electrons. Consequently, $C_3H_7Br(l)$ has the strongest London forces and the highest boiling point, whereas $CH_3Br(g)$ has the weakest London forces and therefore the lowest boiling point.

50. $CH_4(g)$: London forces (10 e^-)
$NH_3(g)$: London forces (10 e^-), dipole-dipole forces, hydrogen bonding
All three molecules have London forces. $BF_3(g)$: London forces (32 e^-)
$CH_4(g)$ is nonpolar and has the weakest London forces; therefore, it has the lowest boiling point. $BF_3(g)$ is also nonpolar (bond dipoles cancel) but it has stronger London forces than $CH_4(g)$; therefore it has the next highest boiling point. $NH_3(g)$ has all three intermolecular forces present, in particular, hydrogen bonding which is known to have a significant effect on boiling points. Therefore, $NH_3(g)$ has the highest boiling point.

51. (a) protons and electrons in the same molecule
(b) protons and electrons in different molecules
(c) oppositely charged ends of polar molecules
(d) proton of a hydrogen atom bonded to a nitrogen, oxygen, of fluorine and a lone pair of electrons in another molecule
(e) positive and negative ions

52. (a) Nickel has a much higher melting point than sodium chloride because the metallic bonding holding nickel atoms together is stronger than the ionic bonding holding sodium and chloride ions together.
(b) Solid nickel will conduct well because the valence electrons of the atoms are free to move. Solid sodium chloride will not conduct because the charges (ions) are not free to move.
(c) Solid nickel will not dissolve because the atoms attract each other much more than water molecules can attract them. Solid sodium chloride will dissolve because the charged entities (ions) are very strongly attracted by polar water molecules.

53. (a) covalent bonds, London forces
(b) covalent bonds, London forces, dipole–dipole forces, hydrogen bonding
(c) covalent bonds, London forces, dipole–dipole forces, hydrogen bonding
(d) covalent bonds, London forces, dipole–dipole forces
(e) ionic bonds
(f) covalent bonds (in a network)

54. (a) Metallic bonds form between metal atoms of the same element in a crystal: the loosely bound valence electrons are attracted to the nuclei of surrounding atoms.
Covalent network bonds form between nonmetal atoms of the same or different elements in a crystal. The atoms form covalent bonds with surrounding atoms in the crystal, rather than within a molecule.

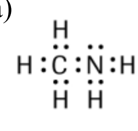
(b) Covalent bonds form between nonmetal atoms that share electrons with a set number of atoms in a molecule, which has a fixed chemical formula.

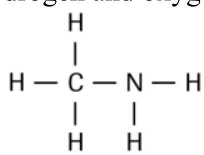
Ionic bonds form between metal and nonmetal ions, through the transfer of electrons. The ions are arranged in a crystal lattice that is defined by a formula unit.

55. Metallic bonding can vary in strength. Weak metallic bonding in a metal results in a low boiling point, whereas strong metallic bonding in a metal results in a high boiling point. The strength of a metallic bond depends on the electronegativity of the metal, and on how closely the atoms can pack together in solid form.

56. Hydrogen (intermolecular) bonds are usually much weaker than most covalent (intramolecular) bonds. Both types of bonds are present in water. When water is heated, hydrogen bonds are broken, which causes the water to vaporize. A significantly larger amount of heat would need to be added in order to break the covalent bonds of the water molecule to form hydrogen and oxygen gases.

57. (a)

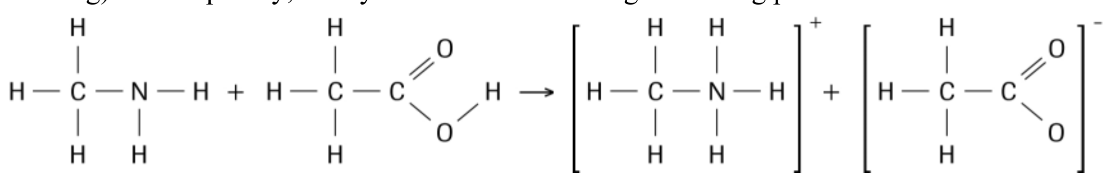




(b) The shape of methylamine around the carbon atom is tetrahedral. Its shape around the nitrogen atom is trigonal pyramidal.

(c) Ethane is nonpolar, with only London forces acting on it, whereas methylamine is polar, with all three intermolecular forces acting on it (London, dipole–dipole, and hydrogen bonding). Consequently, methylamine has a much higher boiling point.

(d)



(e) Lemon juice and vinegar reduce the odour of fish because they are acidic, and react (as shown in (d)) with the basic amines that cause “fishy” odours.

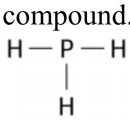
58. (a) Na_2O, MgO, and Al_2O_3 are ionic whereas P_2O_5, SO_2, and ClO_2 are molecular.

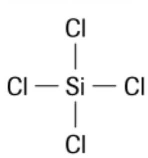
(b) SiO_2 is neither ionic nor molecular. It is a network covalent crystal.

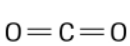
(c)

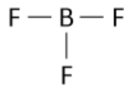
Na_2O	2.5
MgO	2.1
Al_2O_3	1.8
SiO_2	1.5
P_2O_5	1.2
SO_2	0.8
ClO_2	0.2

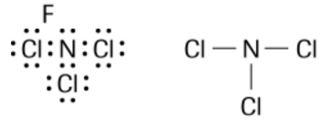
(d) The greater the difference in electronegativity, the more ionic are the properties of the compound.

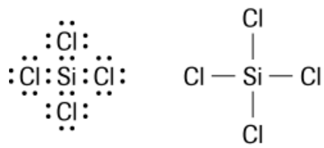
59. (a)  PH_3, phosphorus trihydride

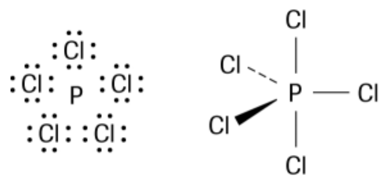
(b)  $SiCl_4$, silicon tetrachloride

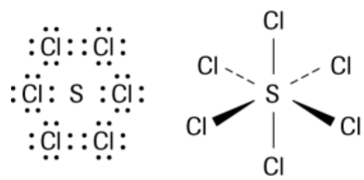
(c)  CO_2, carbon dioxide

(d)  BF_3, boron trifluoride (does not follow the octet rule)

60. (a)  polar

(b)  nonpolar

(c)  (does not follow the octet rule)

(d)  (does not follow the octet rule)

61. Some substances dissolve more easily than others because some entities have greater affinity for solvents than others. For instance, salt dissolves in water because water molecules are polar, and strongly attract both the positive and the negative ions in sodium chloride. When each ion is surrounded by water molecules, it separates from the solid and enters into solution.
    - Different substances have widely different melting points because the bonds holding the entities together are of differing strengths.
    - Models can increase our understanding by giving us a mental picture of the concept, helping us to visualize it in three dimensions.
62. The bond lengths given in Table 2 indicate that as the type of covalent bond between two atoms involves more electrons (single, double, triple), the bond length decreases.
63. Saturated fats have only single bonds between carbon atoms in the fatty acid chain—they are fully saturated with hydrogen atoms. Unsaturated fats have one or more double bonds between carbon atoms in the fatty acid chain, which reduces the number of hydrogen atoms that bond to the carbon chain. Trans fats are formed when liquid oils (unsaturated fats) are made into solid (saturated) fats through the process of hydrogenation (addition of H atoms). Hydrogenation adds H atoms to opposite sides of the C=C bond, hence the term “trans.”

Fat and Caloric Comparison of Four Food Products

	Old cheddar cheese	Margarine	Potato chips	Ice cream
Serving size	30 g	2 tsp or 10 g	21 chips or 50 g	½ cup or 125 mL
Calories	120 cal	70 cal	200 cal	190 cal
Total fat	10 g 15%	8 g 12%	18 g 27%	11 g 17%
Saturated fat + Trans fat	6 g 0.3 g 32%	1.5 g 1 g 13%	2 g 0 g 10%	6g 0.4g 32%
Cholesterol	2.5 mg	0 g	0 g	15 mg

Note: The percentages provided are the percentage of the daily recommended intake.

Many medical studies show that consumption of trans fats raises the concentration of low-density lipoproteins (LDL) in the bloodstream, while lowering high-density lipoproteins

(HDL). LDL moves cholesterol from the liver and releases it in the blood vessels, while HDL carries cholesterol from the blood back to the liver. There have been a number of health concerns raised recently regarding trans fat consumption. Consumption of trans fats appears to be linked to cholesterol deposits on artery walls. These deposits can impede blood flow and eventually cause a complete blockage, resulting in a stroke or heart attack and leading to the development of insulin resistance and type 2 diabetes.