

Chapter 3 Review Answers

Student Textbook pages 159–161

Answers to Knowledge/
Understanding Questions

- Rutherford noticed, in his gold foil experiment, that most of the alpha particles passed through the gold foil with almost no deflection. About one in every 8000, however, bounced back (scattered). Based on this evidence, Rutherford proposed that the atom is mostly empty space. The bouncing back of the alpha particles suggested that most of the mass of the atom is in a very tiny and dense nucleus of the atom.
- Rutherford's nuclear model successfully accounted for the observations of alpha particle scattering, but it raised questions. If the electron were circling the nucleus, then classical physics predicted that it would give off energy, resulting in the electron spiralling into the nucleus. However, atoms did not collapse, and neither did they give off a continuous spectrum. For example, the atomic spectrum of hydrogen appeared as lines instead of a continuous smear of colour, like the spectrum of the sun. These observations and inferences led Bohr to place restrictions on electrons. These restrictions allowed him to account for the atomic spectra of hydrogen. He and Max Planck introduced the idea that the electrons in an atom can only have certain energies: in other words, the energy of an electron in an atom is quantized.
- Both Rutherford and Bohr proposed atomic models that had a nucleus, with electrons orbiting around the nucleus. In Bohr's model, the paths of the electrons were fixed—they could only be at certain distances from the nucleus. In Rutherford's model, however, electrons were free to spin around the nucleus at different distances from the nucleus. There were no fixed paths. Bohr's description of an atom was closer to a planetary model than Rutherford's, since Earth does follow a predictable path around the sun.
- (a) Planck proposed that matter, at the atomic level, could absorb or emit only discrete quantities of energy. The energy of an atom was quantized.
(b) De Broglie proposed that matter had wave-like properties.
(c) Einstein proposed that light was also quantized and had particle-like properties. The quanta of light were later called photons.
(d) Heisenberg showed that it was not possible to know both a particle's position and its momentum precisely. He and Schrödinger are credited with introducing quantum mechanics.
(e) Schrödinger used mathematics and statistics to combine de Broglie's idea of matter waves and Einstein's idea of quantized energy particles (photons). Schrödinger's ideas, together with the Heisenberg's uncertainty principle, led to a description of atomic particles in terms of wave equations.
- The *s* block includes hydrogen, helium, and the elements of group 1 (IA) and group 2 (IIA). Valence electrons occupy only the *ns* orbital in this block. The *p* block includes elements from group 13 (IIIA) to group 18 (VIIIA). Electron configurations of the *p* block atoms take the general form ns^2np^a , where *a* represents a number from 1 to 6. The *d* block includes all the transition elements from group 3 (IIIB) to group 12 (IIB). In general, *d* block elements have filled *ns* orbitals, as well as filled or partially filled *d* orbitals. The *f* block includes all the inner transition elements. Atoms of *f* block elements have filled orbitals in the outer energy levels, as well as filled or partially filled *4f* or *5f* orbitals.
- Pauli exclusion principle restricted the number of electrons in an orbital to be no more than two. Thus, the maximum number of electrons is 2 for an *s* type orbital, 6 for a *p* type, 10 for a *d* type, and 14 for an *f* type. According to Hund's rule, when electrons are added to a *p*, *d*, or *f* type orbital, electrons of the same spin are added separately to orbitals with the same energy first, before any pairing of electrons occurs. Thus, when writing the electron configuration of the first three electrons with $l = 1$ and $m_l = 0, +1$ and -1 , all three electrons have $m_s = +\frac{1}{2}$.
- The electron is in the *p* orbital ($l = 1$) of the third energy level ($n = 3$): that is, the *3p* orbital.
- An excited state sulfur atom is a sulfur atom that has absorbed a quantity of energy so that one or more of its ground state electrons are moved to an orbital of higher energy. The ground state electron configuration for sulfur is $1s^2 2s^2 2p^6 3s^2 3p^4$. Examples of excited states can be $1s^2 2s^2 2p^6 3s^2 3p^3 4s^1$ and $1s^2 2s^2 2p^6 3s^2 3p^3 5s^1$.
- The second member of each of the following pairs has a larger radius because it is farther down a group in the periodic table. Each period of the period table begins with the first member of a new and higher energy level, resulting in a larger radius of the atom.
(c) Ca, Sr
(d) Al, Ga
(f) Cl, Br
For the two pairs in (a) and (e), Na and O have the larger atomic radius since atomic radius decreases across a period. The effective nuclear charge increases across a period, pulling the electrons closer to the nucleus. For (b), K has the larger atomic radius because Cl is near the end of period 3 and K is at the beginning of the next period, where a new and higher energy level begins.
- U, *f* block, inner transition; Zr, *d* block, transition; Se, *p* block, main group; Rb, *s* block, main group; Re, *d* block, transition; Sr, *s* block, main group; Dy, *f* block, inner transition; Kr, *p* block, main group

11. The formulas for the oxides are Na_2O , K_2O , MgO , CaO , Al_2O_3 , and Ga_2O_3 . Each member of a group has the same outermost, or highest energy, electrons. Knowing the formula of one member of a group enables you to know the formulas of the other members of the same group.
12. (a) The ground state electron configuration for arsenic, As, is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$. Each filled orbital should contain two electrons. According to Hund's rule, each of the three $4p$ orbitals has one un-paired electron. The total number of orbitals is 18.
- (b) Other than the three $4p$ orbitals that are half filled, 15 of the orbitals are completely filled.
- (c) There are four orbitals associated with the fourth energy level: one $4s$ and three $4p$.
- (d) Theoretically speaking, one electron can occupy one orbital in the excited state. The maximum number would be 33 and the minimum is 17.

Answers to Inquiry Questions

13. The electron configuration for vanadium, V, is $[\text{Ar}] 4s^2 3d^3$. Therefore, diagram (c) would be the correct orbital diagram. According to the aufbau principle, the $4s$ orbital fills first, before the $3d$ orbital. When the $4s$ orbital is filled, the three electrons fill the $3d$ orbitals following Hund's rule. That is, the three electrons of the same spin are added separately to the three $3d$ orbitals before any pairing of electrons occurs.
14. (a) The electron configuration for carbon is $1s^2 2s^2 2p^2$. According to Hund's rule, the two $2p$ electrons both have the same spin and should be in two different orbitals. Thus, in the diagram, the last two arrows should point upward and there should only be one arrow in each of the first two boxes in the last row of $2p$ orbitals.
- (b) The electron configuration for iron is $[\text{Ar}] 3d^6 4s^2$. Since the $3d$ orbitals fill before the $4p$ orbitals, the arrow in one of the $4p$ orbitals should point downward and move to pair up with the $3d$ electron in the first box of the row of five $3d$ orbitals.
- (c) The electron configuration for bromine is $[\text{Ar}] 3d^{10} 4s^2 4p^5$. Thus, two arrows pointing downward should be added to the first two boxes of the $4p$ orbitals to pair up the electrons.
15. (a) magnesium: ground state electron configuration is $1s^2 2s^2 2p^6 3s^2$
- (b) chlorine: ground state electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^5$
- (c) manganese: ground state electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
- (d) yttrium: ground state electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^1$

16. and 17.

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1																		
2													f				c	
3														h				a
4		g					j				j							i
5		b					j				j							
6		d									j			e				
7																		

Code:

	s block
	p block
	d block
	f block

18. For element 126, the ground state electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^{10} 7p^6 8s^2 5g^6$ or $[\text{Rn}] 7s^2 5f^{14} 6d^{10} 7p^6 8s^2 5g^6$. Thus, the outermost electron is in $8s^2$, so $n = 8$ and $l = 0$. There would be just one s orbital in this energy sublevel.

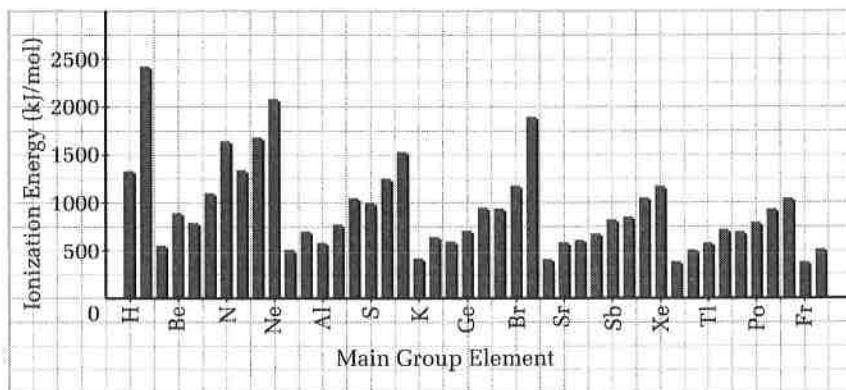
Answers to Communication Questions

19. (a) K is in group 1 (IA), period 4, and Sr is in group 2 (IIA), period 5. The atomic size of K is smaller than the atomic size of Rb, which is larger than the atomic size of Sr. Between K and Rb, an additional energy level makes Rb larger. An increase in the effective charge makes Sr smaller than Rb. However, neither effect gives a quantitative measure, so it is hard to know whether K or Sr has a larger atomic size. In actual fact, the atomic size of K is 227 pm, of Rb is 248 pm, and of Sr is 215 pm. Thus, the effective charge is the more important factor here.
- (b) The electron configuration for Mn is $[\text{Ar}] 3d^5 4s^2$. The electron configuration for Fe is $[\text{Ar}] 3d^6 4s^2$. Both Mn and Fe have two electrons in the $4s$ orbital. Ranking these two elements according to their first ionization energies could be tricky because of the half-filled d type orbital. The atomic radius of Mn is 127 pm and the atomic radius of Fe is 126 pm, as expected by the increase in the effective

charge. Thus, the first ionization energy of Fe is likely to be higher than the first ionization energy of Mn. From the graph of first ionization energies, Fe has a higher first ionization energy than Mn. Apparently, the half-filled d type orbital does not alter the trend for these elements.

- (c) Na is in group 1 (IA), period 3, and Ca is in group 2 (IIA), period 4. The first ionization energy generally decreases down a group and increases across a period. Within the same group, K has a lower first ionization energy than Na, since K is in the next higher energy level. Also, across the same period, K has a lower first ionization energy than Ca. Again, a quantitative measure is needed to rank the ionization energies, and thus the metallic character, of Na and Ca.
20. Atoms are not solid spheres, and their volume is described in terms of probabilities. Therefore, the size of an atom has to be measured indirectly by first measuring the distance between the nuclei of the bonded, neighbouring atoms. The atomic radius is then calculated from the bond length.
21. (a) For valence electron configuration $5s^1$, the period number of the element comes from the 5 of the $5s$ orbital. The position in the period is determined by the total number of valence electrons, which is 1. The element is the first element of period 5, Rb.
- (b) For valence electron configuration $4s^2 3d^2$, the period is determined by the 4 in the $4s$ orbital and the position in the period by the total number of valence electrons, which is 4. Thus, the element is in the fourth element of period 4, Ti.
- (c) For valence electron configuration $3s^2$, the element is the second element in period 3, Mg.
- (d) For valence electron configuration $4s^2 3d^{10} 4p^3$, the element is the fifteenth element in period 4, As.
22. When the wave equations were solved for an electron, the rules for the first three quantum numbers were determined: $n = 1, 2, 3 \dots n$, and for every value of n , $l = 0, \pm 1, \pm 2, \pm 3 \dots \pm(n - 1)$, and for every value of l , $m_l = 0, \pm 1, \pm 2, \pm 3 \dots \pm l$. When $n = 1$, l can only have one value, 0. Thus, there is only one type of orbital, s orbital, in the first energy level (period 1).
23. (a) The ground state electron configuration for scandium is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$. The ground state electron configuration for aluminum is $1s^2 2s^2 2p^6 3s^2 3p^1$.
- (b) Since there are three outermost, or valence, electrons, scandium and aluminium might be expected to form $3+$ ions. Since the composition of the oxides of both elements implied a $3+$ ion, this would further encourage chemists to group them together.
- (c) The placement of scandium and aluminum in different groups is based on trends in properties (such as atomic radii and ionization energies) and on electron configurations. For example, if scandium was between aluminum and gallium in group 13 (IIIA), then the trend in atomic radius and ionization energy would have a discontinuity at scandium. The radius would gradually increase from boron to aluminum, and then take a big increase in size to scandium and a big decrease in size to gallium. There is a similar discontinuity in ionization energies. Thus, scandium fits in with the trends better where it is.

24.



Answers to Making Connections Questions

25. There are no correct or typical answers to questions like this, but there are some things to look for in the argument presented by students.

The first is a list of applications of plutonium. Plutonium is used for

- nuclear weapons
- a source of nuclear power
- pacemakers

Students should then list properties that make plutonium useful in these applications:

- The most important isotope of plutonium is plutonium-239, which has a half-life of 24 360 years.
- Plutonium-239 readily undergoes fission. It can be both used and produced in quantity in nuclear reactors. In addition to being fissionable, it can be split by neutrons of very low (ideally zero) energy. Thus, it can be assembled into a critical mass, and hence can sustain a chain reaction without an external source of neutrons. This makes it useful for nuclear weapons.
- Plutonium-238 has been used to power pacemakers and equipment on the moon, using the heat it emits.

Next, students should discuss the risks and the benefits:

- The risk of working with plutonium is the high radioactivity. The benefit is the energy produced.
- Plutonium-239 is an extremely hazardous poison due to its high radioactivity.
- To reduce the risks, companies need to invest in costly protection for workers and the environment. Since

radioactivity is very penetrating, thick concrete walls, huge vats of heavy water, and massive steel containers are required to contain it. Therefore, building a reactor and its containment systems is very expensive.

Dismantling the facility once it can no longer be used is also expensive, since much of the waste material is radioactive.

- Economists say that power plants are cheaper to operate than coal-fired or natural-gas-fired electric generating plants. However, when higher construction costs for new reactors and the price of decommissioning old plants and disposing of waste are factored in, coal and natural gas plants are cheaper overall.

Finally, students have to justify their opinions on whether the benefits outweigh the risks.

- Their opinions will be very personal and not likely to be based on evidence alone. There is much material, both for and against the use of nuclear energy. Students may have difficulty sorting through all the arguments regarding the costs and the alternatives, but the issue is important and worth the effort.

26. Again there are no right or wrong answers to this question. Because of the enormity of the issue, students should focus on the research conducted by one institution. They might first look at the mission statement and goals of the institution and then justify if the money should be spent elsewhere.

7. The VSEPR notation for the molecule is AX_4 . The arrangement of the electron pairs about the germanium atom is tetrahedral, and the shape of the molecule is also tetrahedral. The Ge—H bond is essentially non-polar ($\Delta EN = 0.19$). The molecule is symmetrical, so the polarities of the four bonds cancel each other. Therefore, the molecule is non-polar.
8. The electronegativity differences of the four halides decrease in this order: Cs—F ($\Delta EN = 3.19$) > Cs—Cl ($\Delta EN = 2.37$) > Cs—Br ($\Delta EN = 2.17$) > Cs—I ($\Delta EN = 1.87$). The decrease in ΔEN implies a decrease in the attractive force between the cation and the anion, because the effective charge on the ions decreases. The size of the halide ions varies and $I^- > Br^- > Cl^- > F^-$. Since the electrostatic force of attraction varies inversely as the square of the distance between the charges, the electrostatic force between a cation and an anion decreases for larger ions. As a result, the force of attraction between the Cs^+ ion and the halide ion decreases in the following order: $CsF > CsCl > CsBr > CsI$. To melt an ionic solid, the bond between the cation and the anion needs to be broken. Therefore, the melting points of the halides, from highest to lowest, are in this order: $CsF > CsCl > CsBr > CsI$.
9. A canoe often comes into contact with rocks when, for example, going aground at a portage or running rapids. The material for making a canoe has to be strong enough to resist possible punctures as a result of this contact. As well, a canoe has to be light enough to be carried. A material that has the strength of steel and the weight of feathers is ideal. KEVLAR[®] is used to make canoes because it is produced from very light fibres, which can withstand the impact of a bullet without shredding. KEVLAR[®] spreads any impact over a wide area, so that no part of the canoe punctures or breaks.
10. Diamond is a network solid. Each carbon atom of diamond is joined to four other carbon atoms in the shape of a tetrahedron, so each diamond is a single molecule.
- (a) An electrical insulator is a poor conductor of electricity. It is used to suppress the flow of electricity. Diamond does not conduct electricity because all its electrons are tied up in the covalent bonds between carbon atoms. Thus, diamond would make an excellent insulator. Diamond is not flexible, however, so it cannot be wrapped around wires. Currently, industrial diamonds are very small and are only used for abrasives.
- (b) To create an optical fibre, the core of the fibre must have a very high refractive index. The refractive index of glass fibre is 1.512. The refractive index of diamond is 2.417. This optical property of diamond would make it an ideal core fibre. However, no technology has yet been developed to grow or machine diamond into long thin threads. Thus, diamond is not yet a suitable material for making cables for fibre optics.

Chapter 4 Review Answers

Student Textbook pages 209-211

Answers to Knowledge/Understanding Questions

- An octet in the valence level of atoms is a stable arrangement—an arrangement that corresponds to a minimum total energy for the system. For example, the noble gases are chemically unreactive, and each atom of a noble gas element (other than He) has eight valence electrons.
 - All the halogen hydrides are polar. They experience dispersion forces as well as dipole-dipole forces. The dispersion forces are greatest in HI and smallest in HF, because the HI molecule is the largest and contains the most electrons. ΔEN is greatest in HF and smallest in HI, so the dipole-dipole forces are greatest in HF and smallest in HI. However, ΔEN is small for the halogens, so the difference in the strength of the dipole-dipole forces is not as significant as the difference in the strength of the dispersion forces. Therefore, the boiling point should increase from HF to HI. In HF, though, the relative sizes of the hydrogen atom and fluorine atom allows hydrogen bonding to occur in this molecule. Hydrogen bonding does not occur in the other hydrogen halides. As a result, HF has the strongest intermolecular forces and its boiling point is higher than the boiling points of the other hydrogen halides.
3. (a) $CHBr_3$
- ```

 ..
 :Br:
 |
H — C — Br:
 |
 :Br:
 ..

```
- (b)  $HS^-$
- ```

  [ H — S : ]-
      ..
  
```
- An induced dipole exists when a non-polar molecule comes close to an ion or a polar molecule. As soon as the ion or polar molecule moves away, the induced dipole disappears. A permanent dipole is produced by the difference in the electronegativities of the atoms within a molecule. The dipole does not disappear, and its existence is independent of the presence of other molecules.
 - The intermolecular forces of non-polar molecules are dispersion forces. The strength of these forces depends on the number of electrons and the shape and size of the molecule. Oil is a mixture of large organic molecules, which have many car-

bon atoms and a large number of electrons. As a result, the dispersion forces between the molecules are very strong, even stronger than the hydrogen bonding in water. Therefore, cooking oil has a higher boiling point than water.

6. (a) dispersion forces and dipole-dipole forces

(b) ionic bond

(c) metallic bond

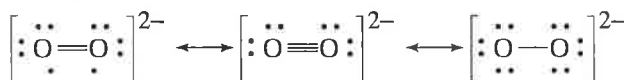
(d) dispersion forces

7. Cesium is a metallic solid, and sulfur is a molecular solid.

The other two elements are gases. Thus, cesium has the highest boiling point, and sulfur has the next highest boiling point. Krypton has more electrons (36) in its atom than oxygen gas (32). The dispersion forces of krypton are greater than the dispersion forces of O_2 . Thus, the order of decreasing boiling points is cesium > sulfur > krypton > oxygen. This order agrees with the measured boiling points: Cs (944 K), S (717.8 K), Kr (119.9 K), and O_2 (90.2 K).

8. H_2O and NH_3 have the same number of electrons in their molecules. ΔEN is greater for O—H (1.24) than for N—H (0.84). Therefore, the O—H bond is more polar and forms stronger hydrogen bonds. Thus, hydrogen bonding is stronger in water than in ammonia.

9. O_2^{2-}



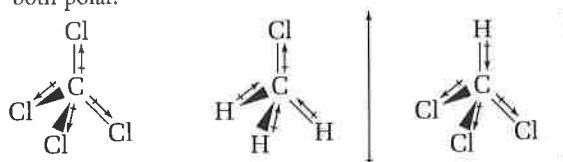
10. The term “crystal” describes something macroscopic—something we can see. The unit cell of a crystal is microscopic. It cannot be seen with the unaided eye or even with a microscope. X-ray diffraction is used to make a model of the crystal, from which the unit cell can be determined. The unit cell is the structure that, if repeated many times, makes up the crystal.

11. If an ionic crystal is broken, each of the broken pieces is still made up of repeating unit cells of the original crystal. Thus, on a microscopic scale, each broken piece has the same shape as the original crystal. Glass is an amorphous solid. If glass is shattered, the silicate tetrahedra do not arrange themselves identically in the broken pieces.

12. Viscosity is the ability of a liquid to resist flow. In flowing, the molecules must move past one another. A liquid with low viscosity thus experiences weaker intermolecular forces between liquid molecules. Surface tension arises due to an imbalance of forces between molecules at the surface of the liquid and the forces between the liquid molecules in the interior. Regardless of the strength of the intermolecular forces, this net force always exists at the surface of the liquid. Therefore, a liquid with low viscosity can have a high surface tension.

13. SiO_2 is a network solid, so it has the highest boiling point. K is a metallic solid. The boiling point of K is therefore relatively lower. C_3H_8 and C_2H_5OH are covalent solids, which have lower boiling points than SiO_2 and K. C_2H_5OH is polar, and it may have hydrogen bonding in its solid form. C_3H_8 is non-polar and will not form hydrogen bonds. C_2H_5OH probably has a higher boiling point than C_3H_8 . Thus, the order of increasing boiling points is $C_3H_8 < C_2H_5OH < K < SiO_2$. Actual boiling point data support this prediction: C_2H_5OH (351 K), SiO_2 (2503 K), C_3H_8 (231 K), K (1039 K).

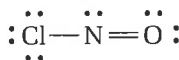
14. The molecular shapes of CCl_4 , CH_3Cl , and $CHCl_3$ are all tetrahedral. CCl_4 is non-polar, and CH_3Cl and $CHCl_3$ are both polar.



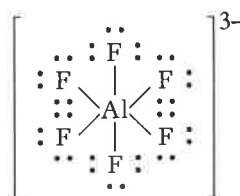
CH_3Cl and $CHCl_3$ would have dipole-dipole forces between molecules in liquid samples.

15. Dipole-dipole attractions are intermolecular forces between two polar molecules. An ionic bond is the attraction between two ions: one positively charged and the other negatively charged. Ionic bonds are usually thought of as intramolecular forces. In an ionic crystal, however, it is impossible to differentiate between the intermolecular and intramolecular forces.

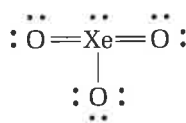
16. (a) The shape of the $NOCl$ molecule is bent.



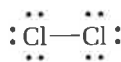
- (b) The shape of the AlF_6^{3-} ion is octahedral.



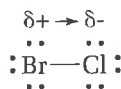
- (c) The shape of the XeO_3 molecule is trigonal pyramidal.



17. (a) Cl_2



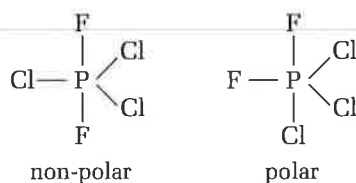
- (b) $BrCl$



18. Substances that have the following VSEPR notations have the same name for the molecular shape and the electron group arrangement: AX_2 (linear), AX_3 (trigonal planar), AX_4 (tetrahedral), AX_5 (trigonal bipyramidal), and AX_6 (octahedral). In other words, a molecule that does not have any lone pairs of electrons has the same name for its molecular shape and electron group arrangement.
19. If the polar bonds in a molecule are arranged around the central atom in such a way that the dipoles can cancel each other, then the molecule is non-polar. Non-polar molecules with polar covalent bonds include binary (two-element) molecules with no electron lone pairs. These molecules have the VSEPR notation AX_n , where $n = 2, 3, 4, 5,$ and 6 . Linear molecules that have the VSEPR notation AX_2E_3 (XeF_2 , for example) and square planar molecules that have the VSEPR notation AX_4E_2 (XeF_4 , for example) are also non-polar.

Answers to Inquiry Questions

20.

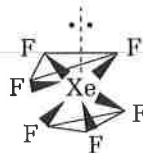


When the two fluorine atoms are bonded to the central atom along the vertical axis of the trigonal bipyramidal arrangement, the dipoles cancel and the molecule is non-polar. When one fluorine atom and one chlorine atom are bonded to the central atom along the vertical axis, however, the molecule is polar.

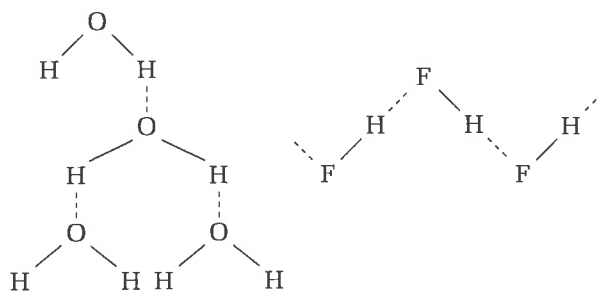
21. At temperatures above 4°C , water cools. The molecules move closer together, so the density of the water increases. This increase in density continues until the temperature of the water is 4°C , when the open hexagonal structure (due to hydrogen bonding) stops further contraction in the volume of the water. When the water cools further, below 4°C , the water molecules form more of the open hexagons that are found in ice. The volume begins to increase, so the density begins to decrease.
22. The electronegativity difference between the elements in a compound determines the type of bonding present. ΔEN of NaCl is 2.23, and the compound is ionic. ΔEN of CaCl_2 is 2.16. The compound is also ionic, but it has less ionic character than NaCl . ΔEN of AlCl_3 is 1.55, and the compound is polar covalent. Al^{3+} and Cl^- ions do not form the lattice structure that is characteristic of ionic solids. The electrostatic force of attraction is greatest in NaCl , so NaCl requires the most energy to melt (to break its ionic lattice). Therefore, NaCl has the highest melting point. Since AlCl_3 is more of a covalent compound, the melting point of AlCl_3 is the lowest.
23. (a) The decomposition of Au_2S_3 occurs at a lower temperature than the decomposition of Au_2S . Thus, Au_2S has

greater ionic character. That is, more energy is required to break the ionic lattice of Au_2S .

- (b) The decomposition of the sulfide is the reverse of the formation reaction. Since the decomposition temperature of Au_2S is higher than the decomposition temperature of Au_2S_3 , Au_2S is thermally a more stable compound than Au_2S_3 . Since stability is related to lattice energy, the lattice energy of Au_2S is greater than the lattice energy of Au_2S_3 .
24. The ground state electron configuration of Xe is $[\text{Kr}] 4d^{10} 5s^2 5p^6$. If six fluorine atoms are bonded to Xe , the resulting molecule will have the VSEPR notation AX_6E . In total, there will be one non-bonding pair and six bonding pairs of electrons surrounding the central atom. Thus, the shape is not octahedral. This molecule has the xenon atom in the centre, with a lone pair of electrons directly above it. The six fluorine atoms are arranged at the three points of each of two triangles above and below the xenon atom.



25. From the electronegativity values, the $\text{H}-\text{F}$ bond ($\Delta EN = 1.78$) is more polar than the $\text{O}-\text{H}$ bond ($\Delta EN = 1.24$). This suggests that hydrogen fluoride may have a higher boiling point than water. However, hydrogen bonding is much stronger in water than in hydrogen fluoride. The bent shape of the water molecule allows for two $\text{O}-\text{H}$ bonds to form hydrogen bonds. Each water molecule can be linked to three other molecules, but each hydrogen fluoride molecule can be linked to only two other molecules. As a result, water should have a higher boiling point. From Figure 4.16 on page 193 of the student textbook, the boiling point of water is 373 K and the boiling point of hydrogen fluoride is 293 K.



Answers to Communication Questions

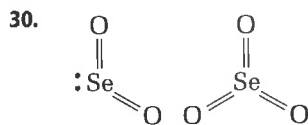
26. The ground state electron configuration of S is $[\text{Ne}]3s^2 3p^4$. To form four bonds, as in SF_4 , one electron of S has to be excited to a $3d$ orbital to give $[\text{Ne}]3s^2 3p^3 3d^1$. As a result, the electron arrangement around the central S atom involves four bonding pairs and one non-bonding pair of electrons. The VSEPR notation for this molecule is AX_4E . The electron

group arrangement is trigonal bipyramidal, and the shape of the molecule is seesaw. The molecule is slightly polar.

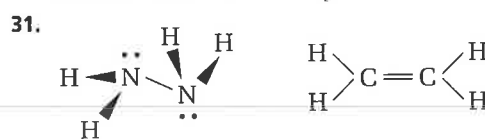
The ground state electron configuration of Si is $[\text{Ne}]3s^2 3p^2$. To form four bonds, as in SiF_4 , one electron of Si has to be excited to a $3p$ orbital to give $[\text{Ne}]3s^1 3p^3$. As a result, the electron arrangement around the central Si atom involves four bonding pairs of electrons. The VSEPR notation for this molecule is AX_4 . The electron group arrangement is tetrahedral, and the shape of the molecule is also tetrahedral. The molecule is non-polar due to the symmetry of the four identical Si—F bonds around the central atom.

27. The electronegativity difference between Na and I is 1.73. NaI is an ionic solid. Therefore, the forces of attraction that must be overcome to melt this solid are the electrostatic attractions due to ionic bonding. The electronegativity difference between H and I is 0.46. HI is a covalent solid that is only slightly polar. The intermolecular forces in HI are dispersion forces and dipole-dipole forces. To melt solid HI, these forces of attraction have to be overcome.
28. (a) Diamond is a crystalline solid that does not melt. It sublimates at a temperature around 4000 K. It is one of the hardest substances known. 2,2-dimethylpropane is a liquid at room temperature. Its melting point is 257 K, and its boiling point is 283 K.
- (b) The C—C bonds in both diamond and 2,2-dimethylpropane are the same. The bond energy is 348 kJ/mol, and the bond length is 154 pm. In both substances, a carbon atom is bonded to four other carbon atoms in the shape of a tetrahedron. In diamond, however, each carbon atom is bonded to four other carbon atoms in a huge network. In 2,2-dimethylpropane, each outer carbon is bonded to one carbon atom and three hydrogen atoms. Thus, 2,2-dimethylpropane exists as discrete molecules. The diamond molecule is the diamond itself. The entire crystal is a molecule because every carbon atom in the crystal is linked to four other carbon atoms within the network.
- (c) The C—C bonds in diamond are networked. To cleave, to smash, or to scratch a diamond, you have to break thousands of C—C bonds. This makes diamond a very hard solid. For 2,2-dimethylpropane, the forces that hold the molecules in a solid are dispersion forces. Thus, the solid is much softer and melts at a much lower temperature.
29. Carbon dioxide is a linear, non-polar covalent molecule. The only intermolecular forces that act between the molecules are dispersion forces. These forces are very weak because the carbon dioxide molecule has only 22 electrons and is quite small. Therefore, carbon dioxide exists as free molecules (a gas) at room temperature. In contrast, each silicon atom is strongly bonded to four oxygen atoms in the form of a tetrahedron, and each oxygen atom is bonded to two silicon atoms to form a network. To melt silica, a lot of energy is needed to break the bonds in the network.

Therefore, silica has a very high melting point (1873K), and it exists as a solid at room temperature.



The electronegativity difference between selenium and oxygen is 0.89. Thus, each Se=O bond is polar. The shape of a SeO_2 molecule is bent, and a net dipole results. Therefore, the molecule is polar. In SeO_3 , the selenium atom forms three double bonds with three oxygen atoms. The shape of the molecule is trigonal planar. Although the Se=O bonds are polar, the symmetry of the molecule causes the dipoles of the Se=O bonds to cancel one another. Therefore, the SeO_3 molecule is non-polar.



Ethene is a planar molecule with a double bond between the two carbon atoms. In this arrangement, the slight polarities of the C—H bond cancel each other. Therefore, ethene is a non-polar molecule. The intermolecular forces in ethene are dispersion forces only. Hydrazine (18 electrons) has stronger dispersion forces than ethene (14 electrons). The N—H bond is also more polar than the C—H bond. A hydrazine molecule consists of two trigonal pyramidal structures linked through the nitrogen atoms. In this arrangement, the polarities of the N—H bonds do not cancel, and the molecule is polar. As well as dipole-dipole forces and dispersion forces, hydrogen bonding occurs between the hydrogen atoms of one molecule and the nitrogen atoms of another molecule. Thus, the intermolecular forces in liquid hydrazine are considerably larger than the intermolecular forces in ethene. As a result, the boiling point of hydrazine should be higher than the boiling point of ethene.

Answers to Making Connections Questions

32. There are no correct or typical answers to questions like this. However, students should include the following ideas and provide reasons for their opinion.

Risks associated with inert substances:

- Materials that are very inert present problems with disposal. A common example is foam cups, which do not decompose and therefore end up in landfill sites. Imagine what would happen if every 12-year-old bought a KEVLAR® jacket for protection. Three years later, the landfill sites would be filled with KEVLAR® jackets that the children outgrew.

Benefits associated with inert substances: