# Unit 2 Review Answers

**Student Textbook pages 214–217**

## Answers to Knowledge/ Understanding Questions

### Multiple Choice

**1.** (d) **2.** (b) **3.** (b) **4.** (d) **5.** (b)

**6.** (e) **7.** (a) **8.** (c) and (e) **9.** (d) **10.** (b)

### Short Answers

**11.** The absorption spectrum for this element should show the complete visible spectrum between wavelengths 400 nm and 750 nm, with two black lines at 620 nm and 640 nm.

**12.** The emission spectrum is produced when excited electrons of the element drop from a higher energy level to a lower energy level. A quantum of light, with a specific wavelength, is given off. The absorption spectrum is produced when electrons of the element absorb a quantum of light, with a specific wavelength, to promote from a lower energy level to a higher energy level.

**13.** Nitrogen: [He] $2s^2\ 2p^3$

Phosphorus: [Ne] $3s^2\ 3p^3$

Arsenic: [Ar] $3d^{10}\ 4s^2\ 4p^3$

**14.** Mg (160 pm) < Na (186 pm) < Ca (197 pm) < K (227 pm)

**15.** $AX_2E$, and $AX_2E_2$ are the VSEPR notations for bent molecules.

**16.** Lone pairs of electrons occupy more space than bond pairs. They tend to push the bond pairs closer together and, as a result, the bond angle decreases.

**17. (a)** Rb (403.0 kJ/mol) < K (418.8 kJ/mol) < Na (495.8 kJ/mol) < Li (520.2 kJ/mol)

**(b)** Li (520.2 kJ/mol) < B (800.6 kJ/mol) < Be (899.5 kJ/mol) < C (1086.5 kJ/mol)

**18.** The two elements in Period 4 whose electron configurations are not correctly predicted by the aufbau principle are Cr and Cu.

**19.** An ionic solid's melting point and solubility are related to the size of the lattice energy.

**20.** $CO_2$ and $CS_2$ are linear non-polar molecules. Dispersion forces are the only intermolecular forces that determine their melting points and boiling points. Molecules of $CS_2$ (38 electrons) are larger than molecules of $CO_2$ (22 electrons). Therefore, dispersion forces are stronger in $CS_2$, and its boiling point is higher than the boiling point of $CO_2$ .

**21.** In the periodic table, atomic radius increases down a group and decreases across a period. First ionization energy decreases down a group and generally increases across a period. Electron affinity tends to decrease down a group and increase across a period. The trends for electron affinity are more irregular than the trends for atomic radius and ionization energy.

**22.** Elements down a group in the periodic table have similar electron configurations. Properties that relate to the electron arrangements of an atom thus repeat periodically.

**23.** The electron configuration for hydrogen is $1s^1$ and for lithium is $1s^2\ 2s^1$. In the hydrogen atom, there is one electron and one proton, so the hydrogen electron experiences the attractive force of the nuclear charge 1+. In the lithium atom, the nuclear charge is 3+. Because of the shielding of the two $1s$ electrons in the inner energy level, the valence electron of lithium does not experience a full positive charge of the nucleus. Instead, the electron experiences an effective charge close to 1+ by the nucleus. Thus, the effective charge that attracts electrons is not always the same as the nuclear charge.

**24.** The maximum number of electrons in a principal energy level is $2n^2$, where $n$ is the principal quantum number.

**25.** The fourth quantum number is the electron spin quantum number, $m_s$. Since an electron can only spin in two opposite directions, $m_s$ can only have two values $\pm\frac{1}{2}$.

**26.** The Heisenberg uncertainty principle states that it is impossible to know both the position and momentum of an object beyond a certain precision. Thus, electrons cannot be described as following a fixed path.

**27.** Dipole-dipole attractions are between molecules that have a permanent positive charge on one end and a negative charge on the other end. Dispersion forces are between molecules that have a temporary dipole caused by vibrations of electrons, setting up an uneven distribution of charge. For a small molecule, such as hydrogen chloride, the dipole-dipole forces are stronger than the dispersion forces. However, the dispersion forces can be quite large if there are a large number of electrons spread over a large molecule.

## Answers to Inquiry Questions

**28. (a)** Since the third ionization energy is significantly higher than the first and the second, an $s$ orbital is likely involved. Therefore, metal X likely belongs to Group 2 (IIA), the alkaline earth metals.

**(b)** The ground state valence electron configuration of X is $ns^2$, where $n$ is the principal quantum number of the outer energy level.

**29.** For $n = 3$, $l = 2$, $m_l = -1$, and $m_s = +\frac{1}{2}$, the $3d$ orbital is involved . The element belongs to the $d$, or transition metal, block elements of Period 4. Assuming that the $d$ orbital fills in order of increasing $m_l$ (−2, −1, 0 + 1 + 2), $m_l = -1$ is the second $d$ orbital to be filled.

Since $m_s = +\frac{1}{2}$, there is one electron in this $3d$ orbital.

The ground state electron configuration for the element is [Ar] $3d^1\ 4s^2$, and the element is titanium, Ti.

**30.** There is a slight increase between the second and the third ionization energies and a significant increase between the fourth and the fifth ionization energies. These increases suggest that the fifth electron is in an inner energy level, and the ground state valence electron configuration of this element is likely $ns^2\ np^2$. Thus, this element is a Group 14 (IVA) element.

**31.** The electronegativity difference for LiBr is 1.98 and for LiI is 1.68. LiBr should have more ionic character than LiI. Thus, LiBr should have a higher melting point than LiI.

**32. (a)** Sb is below As in Group 15 (VA) of the periodic table. With the same number of valence electrons, Sb has a larger positive charge that results in stronger metallic bonding forces in the metal.

**(b)** Si has greater number of valence electrons and a larger positive charge than P. Therefore, Si has stronger metallic bonding forces.

**(c)** For the same reason as in part (a), within the same group, Mg is more metallic than Be. For the same reason as in part (b), within the same period, Na is more metallic than Mg. Therefore, Na is more metallic than Be.

**33.** The first ionization energies of four Period 2 elements are given in the table below.

| Period 2 element | Li | Be | B | C |
|---|---|---|---|---|
| Electron configuration | $1s^2\,2s^1$ | $1s^2\,2s^2$ | $1s^2 2s^2 2p^1$ | $1s^2 2s^2 2p^2$ |
| $IE_1$ (kJ/mol) | 520 | 900 | 801 | 1087 |

The first ionization energy generally increases across a period. The discrepancy occurs at Be, where extra stability is associated with Be's filled 2s orbital. A big increase in first ionization energy from Li to Be is observed. As a result, the predicted increase from Be to B becomes a decrease.

**34.** Molecules with the VSEPR notation $AX_n$ are always non-polar if the molecule has only two elements. The symmetry of the molecule thus allows all the dipoles to cancel each other to result in a non-polar molecule.

**35. (a)** The electrons that will be available to participate in metallic bonding are most likely the $s^2$ electrons. These electrons are in the outer energy level and are most easily moved or removed.

**(b)** Since the $s^2$ electrons are easily removed, the charge carried by stable ions of this group is most likely to be 2+.

**(c)** The fact that mercury is a liquid and the other two metals are solids implies that the metallic bonding in mercury is not as strong as the metallic bonding in zinc or in cadmium.

**36.** Electronegativities: N (3.04), H (2.20), F (3.98), I (2.66), Cl (3.16)

**(a)** N is more electronegative, and it is the negative end of the dipole (N ← H).

**(b)** F is more electronegative, and it is the negative end of the dipole (F ← N).

**(c)** Cl is more electronegative, and it is the negative end of the dipole (I → Cl).

**37.** Electronegativities: F (3.98), Cl (3.16), Br (2.96), I (2.66), Si (1.50), P (2.19), S (2.58)

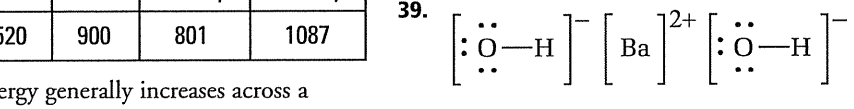
**(a)** The order of increasing bond polarity is Cl—Cl, Br→Cl, Cl→F. The arrow points toward the element with a partial negative charge. Cl—Cl is non-polar.

**(b)** The order of increasing bond polarity is Si—Si, S→Cl, P→Cl, Si→Cl. The arrow points toward the element with a partial negative charge. Si—Si is non-polar.

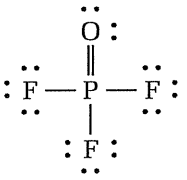
**38.** A binary compound with the formula $AY_3$ could have the following VSEPR notations: $AX_3$ (trigonal planar), $AX_3E$ (trigonal pyramidal), and $AX_3E_2$ (T-shaped). Since the molecule is polar, the molecular shape could be trigonal pyramidal or T-shaped. However, the molecular shape could not be trigonal planar, which would give a non-polar molecule.

## Answers to Communication Questions

**39.**

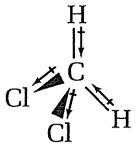



**40.**



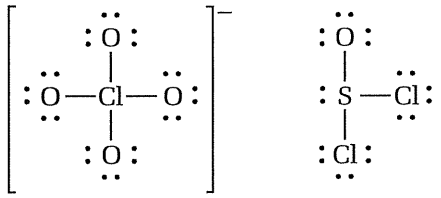

The shape of the molecule is tetrahedral. The polarity of the P—F bond is different from the polarity of the P = O bond. As a result, the dipoles do not cancel each other, and the molecule is polar.

**41.**



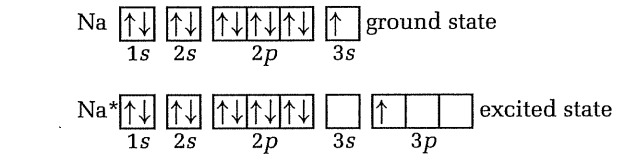

The two atoms of chlorine and the two atoms of hydrogen are arranged in a tetrahedral shape around the carbon. The C—Cl bonds are more polar than the C—H bonds. Therefore, the polarities of the C—H bonds and C—Cl bonds do not cancel each other, and the molecule is polar.

**42.**



Assuming that co-ordinate coralent bonds are formed between Cl and O, and between S and O, neither $ClO_4^-$ nor $OSCl_2$ have an expanded valence energy level.

**43.** The ground state electron configuration of Na is $1s^2\ 2s^2\ 2p^6\ 3s^1$. Assuming that the outermost electron is excited, the first possible excited state would be $1s^2\ 2s^2\ 2p^6\ 3s^0\ 3p^1$.




**44. (a)** The ground state electron configuration of Li is $1s^2\ 2s^1$. The quantum numbers for the outermost electron are $n = 2$, $l = 0$, $m_l = 0$, and $m_s = +\frac{1}{2}$.

**(b)** The ground state electron configuration of $Cl^-$ is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$. The quantum numbers for the outermost electron are $n = 3$, $l = 1$, $m_l = +1$, and $m_s = -\frac{1}{2}$.

**45.** Metallic character generally decreases across a period and increases down a group. Atomic radius also decreases across a period and increases down a group. Ionization energy, however, is related to overcoming the electrostatic forces holding electrons in atoms. The electrostatic forces of attraction vary directly with the charges on the particles, and inversely with the square of the distance between the charges. Thus, the general trends in ionization energies are opposite. Ionization energy increases across a period and decreases down a group.

**46.** Boiling point generally increases down a group. The expected trend in boiling points is $H_2S < H_2O < H_2Se$. However, $H_2O$ does not follow the trend. The intermolecular forces in $H_2O$ are much stronger than the intermolecular forces in $H_2S$ and $H_2Se$ because there is hydrogen bonding in $H_2O$. There is no hydrogen bonding in $H_2S$ and $H_2Se$ because neither $H_2S$ nor $H_2Se$ is small or electronegative enough to support hydrogen bonding. Thus, $H_2O$ has a much higher boiling point than $H_2S$ and $H_2Se$.

**47.** Covalent bonds are the simultaneous attractions between the nuclei and electrons of two or more atoms. In non-polar covalent bonds, the attractions of the different nuclei of the atoms for the electrons are the same. In polar covalent bonds, the attractions of the different nuclei of the atoms for the electrons are different. That is, the strength of attraction of the nucleus of one of the atoms for the electrons is greater than the strength of attraction of the nuclei of the other atoms. In the O—H bond in a water molecule, the oxygen nucleus attracts the pair of electrons between the two atoms more than the hydrogen nucleus does. Therefore, the O—H bond is polar covalent.

**48. (a)** A metallic element has metallic lustre (shininess), is a good conductor of electricity and heat, is malleable and ductile, and melts at a relatively high temperature.

**(b)** A metallic element reacts with oxygen to form oxides that dissolve in water to form strong bases. It reacts with chlorine to form salts that conduct electricity in molten state or in an aqueous solution.

**49. (a)** Successive ionization energies always increase because the effective charge increases as an electron is removed from an atom. After the first electron is removed from a neutral atom, the second electron is then removed from an ion with a charge of 1+, the third electron from an ion with a charge of 2+, and so on. If an electron is removed from an inner energy level, the increase in ionization energy is even more significant.

**(b)** All Group 1 (IA) elements have the ground state valence electron configuration of $ns^1$. These elements have a low first ionization energy and a much higher second ionization energy. The second electron has to be removed from an inner energy level that is closer to the nucleus of the atom. To do so, more energy is required. If an unknown element shows a big jump in the energy required to remove the second electron from an atom, you can conclude that the element is a member of Group 1 (IA).

**50.** Hund's rule states that, when filling orbitals, electrons try to occupy as many orbitals of the same energy first before they start to pair up in these orbitals. Single electrons in orbitals of the same energy must have the same spin. For the carbon atom, the electron configuration is $1s^2\ 2s^2\ 2p^2$. According to Hund's rule, the two $p$ electrons are in two different $p$ orbitals and have the same spin. That is, the quantum numbers for the first $p$ electron are $n = 2$, $l = 1$ $m_l = -1$, and $m_s = +\frac{1}{2}$, and the quantum numbers for the second $p$ electron are $n = 2$, $l = 1$ $m_l = 0$, $m_s = +\frac{1}{2}$.

**51. (a)** The percent ionic character increases with increasing $\Delta EN$. Thus, the greater the electronegativity difference between the elements, the more ionic character a bond has.

**(b)** Chlorine, $Cl_2$, has 0% ionic character. This implies that the two chlorine nuclei attract the pair of electrons between them equally. The bond is described as a non-polar covalent bond.

**(c)** The trend of the graph suggests that there is no compound with 100% ionic character. This implies that the model for covalent bonding applies to all compounds. Even for ionic compounds with large electronegativity difference, complete electron transfer (as suggested by the ionic bonding model) is unlikely.

**(d)** The arbitrary cutoff suggests that bonding is a continuum with the ionic bond on one extreme and the non-polar

covalent bond on the other extreme. Most of the bonds are somewhere in-between—polar covalent. Thus, all bonds are considered to be the simultaneous attractions between the nuclei and electrons of two or more atoms. The difference in electronegativity determines the nature of the bond, as shown by the graph.

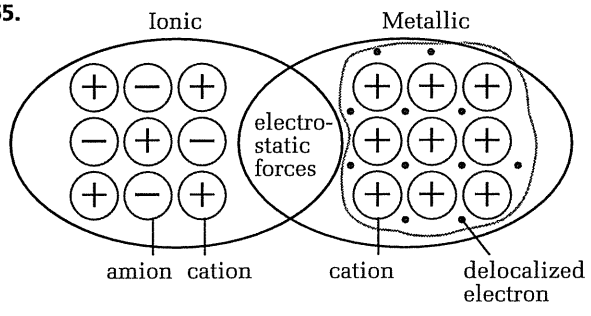
**52. (a)** Electronegativity increases across a period and decreases down a group in the periodic table.

**(b)** Atomic radius decreases across a period and increases down a group in the periodic table. The inverse relationship between electronegativity and atomic radius is expected because electronegativity is defined as the ability of an atom to attract electrons. The electrostatic force of attraction between two atoms that share a pair of electrons is inversely related to the distance between the atoms. Thus, as the atomic radius gets smaller, the distance between the shared pair of electrons gets shorter. The electrostatic force between the nuclei and the shared electrons increases. As a result, as the atomic radius decreases, the electronegativity of the atom increases.

**53.** Heisenberg's uncertainty principle states that it is impossible to know both the location and the momentum of a particle with any degree of certainty. De Broglie suggested that particles have wave properties. Since electrons are particles, the ideas of both scientists suggest that the location of an electron in an atom is uncertain.

**54.** Fluorine is much more reactive than oxygen or chlorine. Fluorine produces more energy than oxygen does when it reacts with the same elements. Fluorine reacts explosively with hydrogen. This means that fluorine forms strong bonds with other elements. Therefore, from a bonding point of view, fluorine is difficult to isolate. Once prepared, the fluorine gas attacks water and vessels. It cannot be contained, and it reacts with moisture in the air to form deadly hydrogen fluoride gas. The poisonous nature of hydrogen fluoride, which injured many people who worked with the gas in the past, makes the safe isolation of fluorine more remarkable. Moisson himself reported four delays in his studies of the electrolysis, due to severe poisoning. The fluorine gas produced was captured in a platinum-iridium alloy U-tube, which was cooled to reduce the rate of attack by fluorine on the vessel walls. Thus, the Nobel Prize was given to Moisson for his dedication and courage when working with extremely toxic materials to forward the cause of science.

**55.**




**56.** An examination of the trends in the periodic table indicates that metallic behaviour follows the same trends as atomic radius. As atomic radius increases down a group in the periodic table, metallic behaviour also increases down the group. The correlation between the increase in atomic radius down a group and the increase in metallic behaviour down a group suggests that as an atom gets smaller, the electrons are held more tightly and thus they are less able to move throughout the solid. Therefore, within the same group, carbon is a non-metal and lead is a metal.

## Answers to Making Connections Questions

**57.** There are no correct or typical answers to questions like this. Expect students to include the following ideas in their editorial:

- The connection is between a group of elements in the periodic table and their social significance to us.
- Transition metals, such as iron and manganese, are important for making steel, which is extremely important to any industrialized nation for the manufacture of cars and machinery.
- Most countries have the facilities to process iron ore into steel.
- Few countries have the resources to harvest the nodules.
- Perhaps the most humanitarian approach would be to allocate the nodules based on population.
- One solution would be to let a group of companies from different countries harvest and smelt the nodules. All the countries involved would be shareholders in this group of companies. These countries would share profits based on the value of their share of the nodules.

**58.** Debate: Military budgets should be exempt from cost-conservation measures that governments typically must consider. The connection here is between researching materials with special bonding properties and spending tax dollars.

Points supporting the position:

- Bonding theories are fairly sophisticated. They are not yet at the point where an investigator can whip up a molecule with a definite set of properties instantly. Most breakthroughs still come from painstaking research. Since research involves people, and people require compensation for their thoughts and deeds, this type of research is expensive.
- Fundamental research, as well as applied research, has to be supported. The latter is even more expensive because the outcome might not be useful. Of course, it can sometimes be quite spectacular: for example, the invention of the atomic bomb.

Points against the position:

- Creative thinking most often occurs when roadblocks to something are put up. If the military had all the money it needed, there would be less creative thought and, as a result, less useful research.
- Governments do not have unlimited funds.