

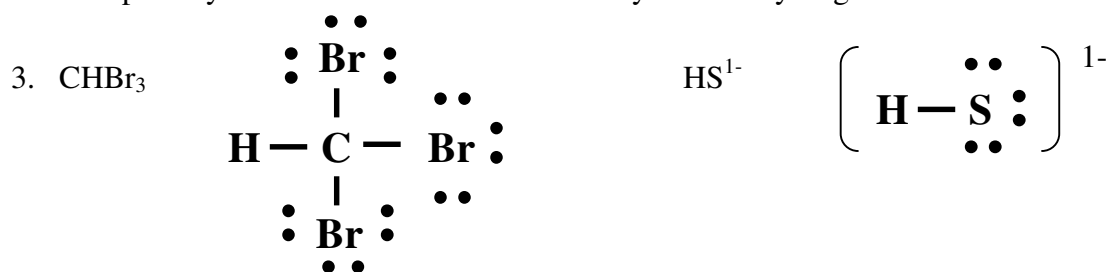
## Answers to Review for Chapter 4: Structure and Properties of Substances

**Chapter 4 review (p. 209 – 211):** Q 2, 3, 6, 8, 9, 13 – 17, 19, 20, 24 – 27, 29 – 31

2. Inter-molecular forces in the hydrogen halides :

	<b>HF</b>	<b>HCl</b>	<b>HBr</b>	<b>HI</b>
molecular shape	linear	linear	linear	linear
symmetrical	no	no	no	no
$\Delta EN$	1.78	0.96	0.76	0.46
inter-molecular forces	LDFs dipole-dipole H-bonding	LDFs dipole-dipole	LDFs weak dip-dip	LDFs

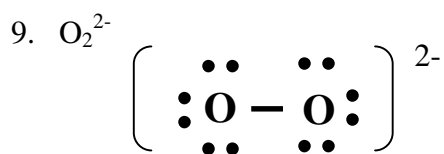
HF has the highest boiling point because it has the strongest inter-molecular attraction due to the extreme polarity of the H – F bond and the ability to form hydrogen bonds between molecules.



6. What types of inter-molecular attraction must be broken to melt the following?

	<b>NH<sub>3</sub></b>	<b>NaI</b>	<b>Fe</b>	<b>CH<sub>4</sub></b>
molecular shape	AX <sub>3</sub> E <sub>1</sub> trigonal pyramidal	linear	not applicable	AX <sub>4</sub> E <sub>0</sub> tetrahedral
symmetrical	no	no	not applicable	yes
$\Delta EN$	0.84	1.73	0	0.35
type of bonding	polar covalent	ionic	metallic	non-polar covalent
inter-molecular forces to break so substance can melt	LDFs dipole-dipole H-bonding	ion to ion (LDFs)	metallic bonding	LDFs only

8. Hydrogen bonding will be stronger in H<sub>2</sub>O than in NH<sub>3</sub> because the  $\Delta EN$  for the O – H bond is 1.24 while the  $\Delta EN$  for the N – H bond is only 0.84. While both molecules are asymmetrical and capable of hydrogen bonding, the more polar O – H bond will cause a stronger dipole and stronger hydrogen bonding.



13.  lowest to highest boiling points →

	<b>C<sub>3</sub>H<sub>8</sub></b>	<b>C<sub>2</sub>H<sub>5</sub>OH</b>	<b>K</b>	<b>SiO<sub>2</sub></b>
type of bonding	non-polar covalent	polar covalent	metallic	covalent network
inter-molecular forces to break so substance can melt	LDFs only, so low boiling point	dip-dips, H-bonding and LDFs, so fairly high bp	metallic bonding in 3-D, so quite high boiling point	covalent bonding in 3-D, very high boiling point

**Answers to questions on page 209 and 210:**

14.

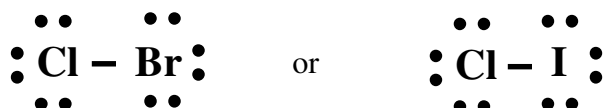
	CCl <sub>4</sub>	CH <sub>3</sub> Cl	CHCl <sub>3</sub>
molecular shape	AX <sub>4</sub> E <sub>0</sub> tetrahedral	AX <sub>4</sub> E <sub>0</sub> tetrahedral	AX <sub>4</sub> E <sub>0</sub> tetrahedral
symmetrical	yes	no	no
ΔEN	0.61	0.35 and 0.61	0.35 and 0.61
type of bonding	non-polar covalent	polar covalent	polar covalent
inter-molecular forces to break so substance can melt	LDFs only	LDFs slight dipole-dipole	LDFs slight dipole-dipole

15. Dipole-dipole attraction is the inter-molecular attraction between partially charged ( $\delta^-$  and  $\delta^+$ ) regions on polar molecules. They are relatively strong forces, and the strength increases as  $\Delta EN$  increases. Ionic bonds are the intra-molecular forces of attraction between fully charged ions ( $-$  and  $+$ ) in an ionic compound and are very strong. (Technically, the inter-molecular attraction between ionic compounds in a crystal lattice are the same as an ionic bond.)

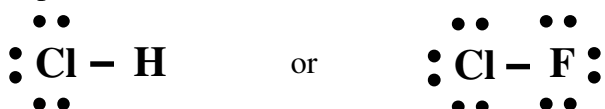
16.

	NOCl	AlF <sub>6</sub> <sup>3-</sup>	XeO <sub>3</sub>
Lewis diagram			
molecular shape	AX <sub>2</sub> E <sub>1</sub> bent (V-shaped)	AX <sub>6</sub> E <sub>0</sub> octahedral	AX <sub>3</sub> E <sub>1</sub> trigonal pyramidal
symmetrical	no	yes	no

17. a) A compound of chlorine (therefore more than one type of atom) that is diatomic (contains 2 atoms) and has only dispersion forces between the molecules when in the liquid state (so  $\Delta EN$  must be less than 0.50). Either Cl-Br or Cl-I would be correct answers:



b) A compound of chlorine (therefore more than one type of atom) that is diatomic (contains 2 atoms) and has dipole-dipole attraction between the molecules when in the liquid state (so  $\Delta EN$  must be between 0.50 and 1.70). Either H-Cl or Cl-F would be correct answers (Cl-Br would also technically have a very slight dipole).

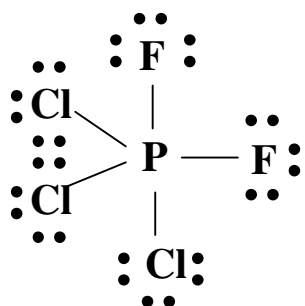


19. A molecule with polar bonds can be a non-polar molecule when the molecule is symmetrical in all planes, because then the dipoles act equally in opposite directions and cancel out.

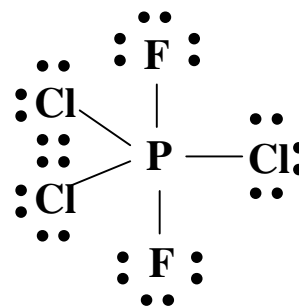
Answers to questions on page 210 :

20.  $\text{PF}_2\text{Cl}_3$  can be drawn so that is both polar and non-polar.

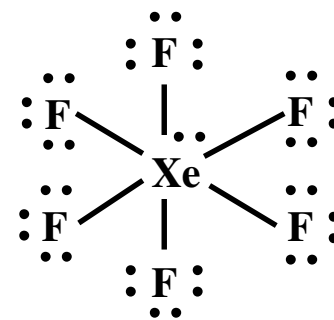
Polar because asymmetrical



Non-polar because symmetrical



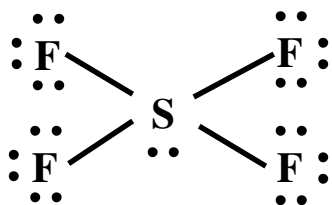
24.  $\text{XeF}_6$  will NOT have an octahedral shape because there is a lone pair of electrons on the central Xe atom. This is because xenon has 8 valence electrons. Six of these electrons are found in the bonds with fluorine, leaving 2 electrons left over as a lone pair.



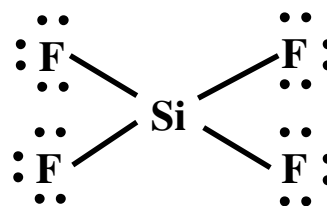
25. Based on  $\Delta\text{EN}$  values, HF has more polar bonds, so it should have a higher boiling point than water. However, water actually has the higher boiling point because of its bent shape. Each water molecule can form hydrogen bonds with several other water molecules, making a strong network of hydrogen bonds. HF, because of its linear shape, can only form hydrogen bonds with a few surrounding molecules, so it has smaller inter-molecular forces and a lower boiling point than water.

26.  $\text{SF}_4$  has the VSEPR notation  $\text{AX}_4\text{E}_1$ , so its shape is seesaw. Seesaw is an asymmetrical shape, so this molecule is polar.  $\text{SiF}_4$  has the VSEPR notation  $\text{AX}_4\text{E}_0$  so it is a tetrahedral shape. This is a symmetrical shape, so  $\text{SiF}_4$  is a non-polar molecule.

$\text{SF}_4$



$\text{SiF}_4$

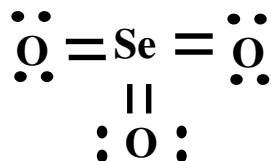


27.  $\text{NaI}$  is an ionic compound, so ion to ion attractions (from the crystal lattice) must be overcome in order to melt samples of this compound.  $\text{HI}$  is a slightly polar covalent molecule (asymmetrical and  $\Delta\text{EN} = 0.46$ ), so there will be weak dipole-dipole attractions and LDFs that must be overcome to allow this substance to melt.

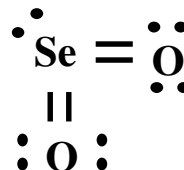
29.  $\text{CO}_2$  is a linear molecule that is symmetrical, so it is non-polar which makes it a gas at room temperature.  $\text{SiO}_2$  is a covalent network solid that is bonded in three dimensions, which makes it a solid at room temperature.

**Answers to questions on page 210:**

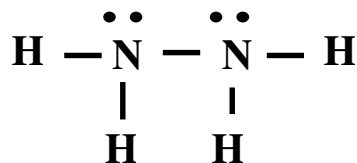
30.  $\text{SeO}_3$  has VSEPR notation  $\text{AX}_3\text{E}_0$  so while it contains polar bonds, it is trigonal planar in shape, symmetrical and non-polar.



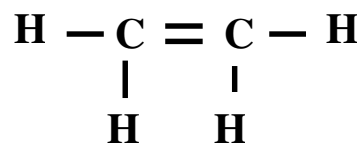
$\text{SeO}_3$  has VSEPR notation  $\text{AX}_2\text{E}_1$ , so it contains polar bonds. It is bent in shape, so it is asymmetrical which means it is a polar molecule.



31.  $\text{N}_2\text{H}_4$  has non-polar covalent bonds between the two nitrogen atoms and polar covalent bonds between the nitrogen and hydrogen atoms. Because of the lone pairs of electrons on the nitrogen atoms, the molecule is asymmetrical, which makes it a polar so it will have dipole-dipole inter-molecular attractions, LDFs and hydrogen bonding.



$\text{C}_2\text{H}_4$  has a non-polar double covalent bond between the two carbon atoms and non-polar covalent bonds between the carbon and hydrogen atoms. Because there are no lone pairs of electrons on the carbon atoms, the molecule is symmetrical, which makes it non-polar. It will have only LDFs between molecules.



**Answers to Unit 2 Review (p. 214-217):** Q 1 – 10, 13 – 19, 21 – 32, 35 – 38, 40 – 43, 45, 47 – 53

Be able to apply what we have learned. For example, predict four physical properties of phosgene ( $\text{COCl}_2$ ), an extremely toxic compound that has been used for chemical warfare.

Phosgene is a polar covalent compound (it is asymmetrical and has polar bonds) so it will have fairly strong dipole-dipole attractions in addition to LDFs. This means that phosgene will probably be a gas or liquid at SATP, have fairly high melting and boiling points, is soluble in water, has a distinctive odour and is a non-conductor of electricity both in its pure form and in solution.

**Answers to questions on page 214 :**

1. d	3. b	5. c (and also b?)	7. a	9. d
2. b	4. d (and also c?)	6. e	8. c	10. b

13.  ${}_{7}\text{N}$  [He]  $2s^2 2p^3$

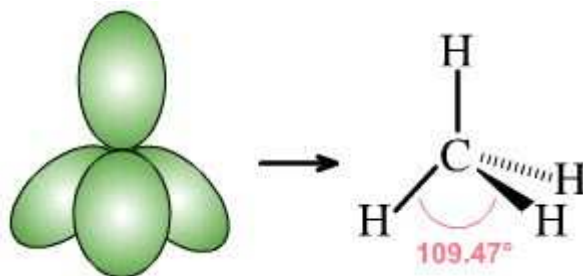
${}_{15}\text{P}$  [Ne]  $3s^2 3p^3$

${}_{33}\text{As}$  [Ar]  $4s^2 3d^{10} 4p^3$ . Because the d sub-level is full, can also write:  ${}_{33}\text{As}$  [Ar]  $3d^{10} 4s^2 4p^3$

14. In order of increasing atomic radius (size):  $\text{Mg} < \text{Na} < \text{Ca} < \text{K}$

15. Molecules with bent shape have the VSEPR notation:  $\text{AX}_2\text{E}_1$  or  $\text{AX}_2\text{E}_2$

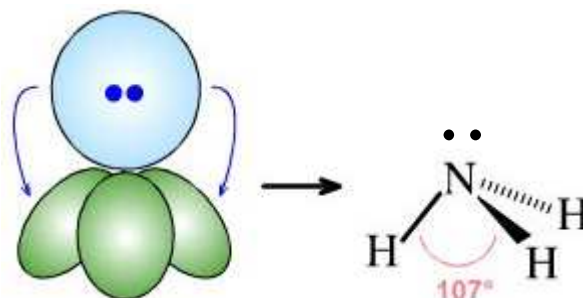
16. Lone pairs of electrons take up more space than bonded pairs, so they tend to repel the electrons in bonds and push them closer together, decreasing the bond angles. For example, the bond angle in a tetrahedral molecule ( $\text{AX}_4\text{E}_0$ ) is  $109.5^\circ$ , but the bond angle for a trigonal pyramidal molecule ( $\text{AX}_3\text{E}_1$ ) with one lone pair is about  $107^\circ$ , and for a bent molecule ( $\text{AX}_2\text{E}_2$ ) with two lone pairs is about  $104.5^\circ$ .



17. In order of increasing  $\text{IE}_1$ :

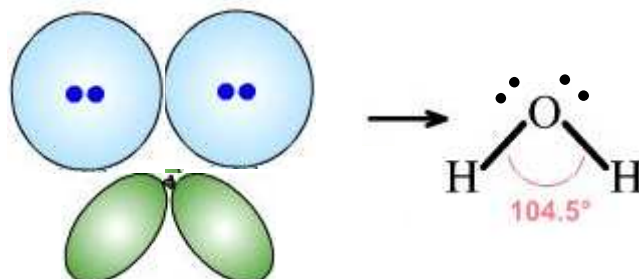
a)  $\text{Rb} < \text{K} < \text{Na} < \text{Li}$

b)  $\text{Li} < \text{Be} < \text{B} < \text{C}$



18. In period 4, chromium and copper both have different electron configurations than is predicted because these elements promote a 4s electron to half-fill or fill the d-sublevel which gives a lower energy electron arrangement.

19. Lattice energy is related to the inter-molecular attraction between the ionic compounds in the crystal lattice. The larger the lattice energy, the higher the melting and boiling points and the harder the crystal.

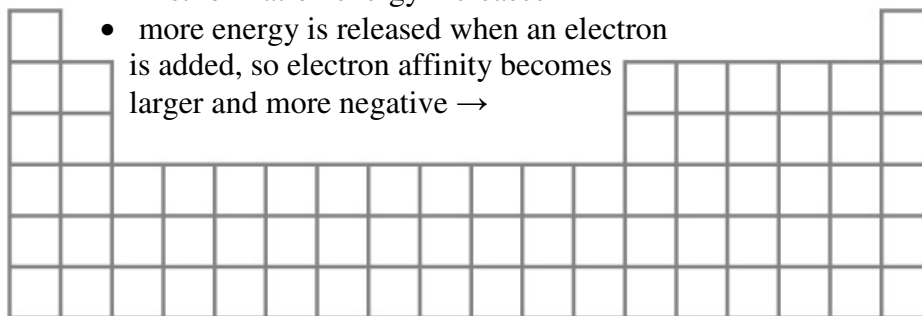


Page 215 Answers to Unit Review in Text:

21.

- atomic radius increases ←
- first ionization energy decreases ←
- less energy is released when an electron is added, so electron affinity becomes smaller and less negative ←

- atomic radius decreases →
- first ionization energy increases →
- more energy is released when an electron is added, so electron affinity becomes larger and more negative →



22. The properties of the elements repeat in a periodic fashion because the properties depend on how many valence electrons are in the outer shell. All elements with one valence electron have similar properties, all elements that need only one valence electron to complete a full valence shell also have similar properties. Once the valence shell is full, the pattern starts again at the beginning of the next period.
23. The nuclear charge refers to the atomic number ( $Z$ ), which is the number of protons in the nucleus of an atom. Hydrogen has a nuclear charge of  $1+$  and lithium has a nuclear charge of  $3+$ .
- The effective nuclear charge ( $Z_{\text{eff}}$ ) refers to the actual "pull" exerted by the nucleus on the valence electrons. This is calculated by subtracting the number of shielding electrons from the nuclear charge. Both hydrogen and lithium have an effective nuclear charge of  $1+$ .
24. The maximum number of electrons in an energy level is calculated using the equation  $2n^2$ , where "n" is the principal quantum number of the energy level you are considering.
25. The fourth quantum number ( $m_s$ ) indicates the spin of the electrons in an orbital. The possible values for  $m_s$  are either  $+\frac{1}{2}$  (up) or  $-\frac{1}{2}$  (down). There are only two possible values for this number because there are only two electrons in an orbital, and only two possible spin patterns.
26. The uncertainly principle states that is impossible to know BOTH where an electron is (its nergy or position away from the nucleus) and where it is going (its path or trajectory). By measuring one of these parameters, we change the other. The implications of the uncertainty principle are that we can't predict exactly where the electrons in an atom are going to be- there are no such thing as electron *orbits*. Instead, we talk about orbitals, which are three-dimensional regions where we are likely to find a particular electron 95% of the time.
27. Dipole-dipole attractions and LDFs are both types of inter-molecular attractions. Dip-dips are associated with permanently charged regions of molecules (full or partial charges) and are quite strong. Their strength increases with the  $\Delta EN$  of the bonds in an asymmetrical molecule. LDFs are temporary very slightly charged regions on the atoms in molecules due to the random motion of electrons in the atom. LDFs are very weak, but increase as the size of atoms increases (because they have more electrons) and increase as the size and "straightness" of molecules increases.
28. A metal "X" forms the compound  $XCl_2$  and its  $IE_3$  is very large.
- a) This indicates that the metal has two valence electrons, so it probably belongs to Group II (the alkaline earth metals).
  - b) In its ground state, the valence electron configuration will be  $ns^2$ .

**Pages 215 and 216 Answers to Unit Review in Text:**

29. For  $n = 3$ ,  $l = 2$ ,  $m_l = -1$  and  $m_s = +\frac{1}{2}$

The last electron is in the third energy level ( $n = 3$ ), and the d orbital ( $l = 2$ ). The possible values for  $m_l$  are  $-2, -1, 0, +1, +2$ . If we follow the convention that the electrons fill in order, then  $m_l$  tells us that this is the  $d^2$  electron. Altogether, the last electron is  $3d^2$ , so the element must be titanium.

30. Given

$IE_1 = 1.09$	easy to remove, so valence electron
$IE_2 = 2.35$	easy to remove, so valence electron
$IE_3 = 4.62$	easy to remove, so valence electron
$IE_4 = 6.22$	easy to remove, so valence electron
$IE_5 = 37.83$	hard to remove, so breaking stable octet

there are four valence electrons, so this element is in Group IVA (4) or IVB (14)

31. Of the ionic solids LiI and LiBr, I would expect LiBr to have the higher melting point because LiBr has a higher  $\Delta EN$ , so it will have stronger inter-molecular attraction.

32. Which element in each pair is less metallic?

- Sb or As? Sb will have more metallic character because it is lower in the Group (has higher shielding effect), so it will lose its valence electrons more easily.
- Si or P? Silicon will have more metallic character because it has weaker  $Z_{eff}$ , so it will lose its valence electrons more easily.
- Be or Na? Na will have more metallic character because it has both more shielding effect and less  $Z_{eff}$ , so it will lose its valence electron much more easily than Be.

35. Zn, Cd and Hg all have the electron configuration ending in  $d^{10}s^2$ .

- the valence electrons to participate in bonding are the  $s^2$  electrons.
- all three elements lose these  $s^2$  electrons to form 2+ ions (mercury can also form 1+ ions)
- mercury is a liquid at SATP, so it has weaker metallic bonding than Zn or Cd

36. a)  $\Delta EN$  N – H is  $(3.04 - 2.20) = 0.84$  so this is a polar bond, N has  $\delta^-$  charge

b)  $\Delta EN$  F – N is  $(3.98 - 3.04) = 0.94$  so this is a polar bond, F has  $\delta^-$  charge

c)  $\Delta EN$  I – Cl is  $(3.16 - 2.66) = 0.50$  so this is defined as a non-polar bond, no partial charges (because it is right on the line, you could also say that this is a very slightly polar bond)

37. In order of increasing polarity:

a) Cl – Cl (no dipole) < Cl – Br (Cl is  $\delta^-$ ) < Cl – F (F is  $\delta^-$ )

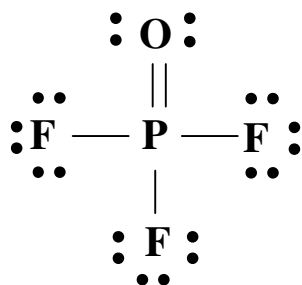
b) Si – Si (no dipole) < S – Cl (Cl is  $\delta^-$ ) < P – Cl (Cl is  $\delta^-$ ) < Si – Cl (Cl is  $\delta^-$ )

38. Molecule  $AY_3$  is polar, therefore it must be an asymmetrical molecule.

It could be:  $AX_3E_1$  (trigonal pyramidal),  $AX_3E_2$  (T-shaped) or  $AX_3E_3$

It can't be  $AX_3E_0$  (trigonal planar) because this would be non-polar

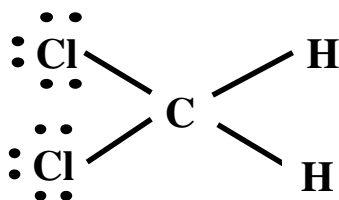
40.  $POF_3$



The molecule is asymmetrical and contains polar bonds, so it is a polar molecule and will have dipole.

**Page 216 Answers to Unit Review in Text:**

41.  $\text{CH}_2\text{Cl}_2$

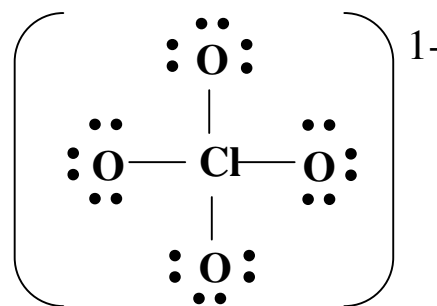


The molecule is tetrahedral shape ( $\text{AX}_4\text{E}_0$ ). It is asymmetrical and contains polar bonds, so it is a polar molecule and will have dipole.

42. Compare  $\text{ClO}_4^{1-}$  and  $\text{OSCl}_2$ . Use the “system” for bonding by the octet rule to determine the coordinate bonds:

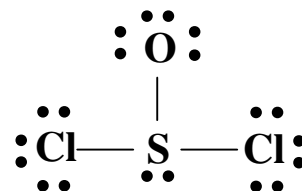
$\text{ClO}_4^{1-}$

- have: 32 electrons
- need: 40 electrons to give all atoms a stable octet
- there are 8 electrons in bonds, or 4 bonds
- shape is tetrahedral
- there is no expanded valence because all of the bonds on chlorine are coordinate covalent



$\text{OSCl}_2$ :

- have: 26 electrons
- need: 32 electrons to give all atoms a stable octet
- there are 6 electrons in bonds, or 6 bonds
- shape is trigonal pyramidal
- there is no expanded valence because all of the bonds on sulfur are regular or coordinate covalent bonds



43. Draw the first excited state for a sodium atom :

Ground state configuration :  ${}_{11}\text{Na } 1s^2 2s^2 2p^6 3s^1$

First excited state :  ${}_{11}\text{Na } 1s^2 2s^2 2p^6 3p^1$  (the 3s electron jumps to the next sub-level, the 3p)

45. Across the periodic table from left to right : metallic character decreases  
atomic size decreases and  
ionization energy increases

Down the periodic table from top to bottom : metallic character increases  
atomic size increases and  
ionization energy decreases

47. Non-polar covalent bonds have  $\Delta\text{EN}$  of less than 0.50

Polar covalent bonds have  $\Delta\text{EN}$  of 0.50 – 1.70

Ionic bonds have  $\Delta\text{EN}$  of greater than 1.70

Because the  $\Delta\text{EN}$  of the O – H bond in water is 1.24, this bond is classified as polar covalent.

48. Physical properties of solid metallic elements are that they are shiny, ductile, malleable, have relatively high melting and boiling points and are good conductors of heat and electricity.

Chemical properties of metallic elements are that they react with water to produce basic solutions, react with chlorine to form ionic compounds (which are electrolytes) and react with acids to produce hydrogen gas (from grade 11). All of these properties are due to fact that metals tend to lose electrons to form positive ions.



## Pages 216 and 217 Answers to Unit Review in Text:

49. Successive ionization energies for any element always increase because you are pulling an electron away from an increasingly positively charged ion.
- b) Any time you break a stable octet, the successive ionization energies will jump considerably. For example, sodium's second ionization energy will be much higher than its first IE. Magnesium's third ionization energy will be much higher than its second IE.
50. Hund's rule states that electrons will half-fill the orbitals of the same sub-level before electrons double up in any of the orbitals. For a carbon atom, this would predict that the 1s orbital will be filled first, the 2s orbital will be filled next, and then one electron will go into each of the 2p<sub>x</sub> and 2p<sub>y</sub> orbitals rather than putting two electrons into the 2p<sub>x</sub> orbital.
51. a) As  $\Delta EN$  increases, the ionic character of the bond also increases.
- b) Ionic character is zero for any bond in which two bonding atoms are identical, for example a Cl – Cl bond. We can infer that the electrons in this molecule are equally attracted to both chlorine atoms, so they are equally shared and the Cl – Cl bond is completely non-polar.
- c) No bonds have 100% ionic character. This implies that although electrons in a bond may be very much more strongly attracted to one atom than the other, there is always at least a tiny amount of attraction to the other atom. Electrons are never 100% transferred to one atom.
- d) There is no single exact "cut-off" point at which bonds can be defined as non-polar, polar or ionic. Instead, all bonds are on a continuum from completely non-polar ( $\Delta EN=0$ ) to increasingly polar. As bonds become more polar, their electrons are shifted more and more towards the atom with higher electronegativity but they are always somewhat attracted to their original atom.
52. Electronegativity increases from left to right across the periodic table, while atomic radius decreases.
- Electronegativity decreases from top to bottom down the periodic table, while atomic radius increases.
- These trends make sense because as an atom has increasing electronegativity, it has a stronger pull on both its own electrons and on new electrons. This exerts a strong pull on new electrons as well as pulling its own electrons in tight to the nucleus, which decreases atomic size.
- Similarly, going down a group, the atoms get bigger which means that the valence electrons are further from the nucleus. This means that the atom will exert a weaker pull on both new electrons and on its own valence electrons, so atomic size increases.
53. Heisenberg suggested the uncertainty principle, which implies that we never know for sure where an electron is going to be in an atom.
- De Broglie suggested that electrons have wave-like properties, which also implies that we never know for sure where an electron is going to be.

**Answers to Practice multiple choice:**

1. c	19. b	37. d	55. c	73. b
2. d	20. d	38. a	56. b	74. d
3. c	21. d	39. d	57. c	75. c
4. d	22. b	40. b	58. a	76. b
5. b	23. a	41. c	59. b	77. b
6. d	24. c	42. a	60. b	78. c
7. c	25. b	43. c	61. b	79. c
8. b	26. b	44. b	62. c	80. d
9. a	27. b	45. d	63. d	81. c
10. d	28. c	46. b	64. a	82. a
11. a	29. c	47. c	65. b	83. a
12. d	30. b	48. a	66. b	84. a
13. a	31. b	49. c	67. b	85. d
14. d	32. c	50. b	68. a	86. c
15. d	33. d	51. b	69. b	87. c
16. b	34. b	52. a	70. b	
17. d	35. a	53. d	71. b	
18. c	36. d	54. d	72. d	