

## 6 Relationships and Patterns in Chemistry

### 6.1 The Development of the Periodic Table

#### Warm Up (p. 286)

1. Students will often suggest that the word “periodic” means that the elements were discovered over different periods of time. The word actually means that similar elemental properties recur on a regular or *periodic* basis.
2. In the modern periodic table, the elements are arranged in order of increasing atomic number.
3. Elements belonging to the same vertical column or chemical family display similar properties.

#### Quick Check (p. 288)

1. English chemist William Olding was the first person to arrange the chemical elements into groups.
2. John Newlands noticed that similar properties seemed to repeat every 8<sup>th</sup> element in the same way that notes on a musical scale repeat every 8<sup>th</sup> tone.
3. As Mendeleev had arranged similar elements vertically on his table, the location of the blank spaces predicted the properties that elements would possess upon being discovered. When those elements were eventually discovered, Mendeleev’s predictions turned out to be extremely accurate. This helped his arrangement of the elements in his periodic table gain widespread acceptance.

#### Quick Check (p. 291)

1.

Element Name	Element Symbol	Group Number	Period Number	Metal/Non-metal/Metalloid
silicon	Si	14	3	metalloid
osmium	Os	8	6	metal
chromium	Cr	6	4	metal
meitnerium	Mt	9	7	metal
antimony	Sb	15	5	metalloid
iodine	I	17	5	non-metal

2. (a) The elements rearranged in order of least metallic to most metallic are:  
fluorine, oxygen, sulphur, aluminum, gallium, chromium, zirconium, cesium.
- (b) The element with the greatest tendency to gain an electron is fluorine.  
The element with the greatest tendency to lose an electron is cesium.

### Quick Check (p. 293)

1.

Property	Family Number	Family Name
(a) reactive non-metals possessing 7 valence electrons	17	halogens
(b) reactive solids that form 2+ cations during reactions	2	alkaline earth metals
(c) invisible gases that are almost totally unreactive	18	noble gases
(d) soft, very reactive silvery solids with 1 valence electron	1	alkali metals

2. The chemical formulas are:

- (a) NaI              (b) BaS              (c) GaCl<sub>3</sub>              (d) Rb<sub>2</sub>S              (e) Mg<sub>3</sub>P<sub>2</sub>              (f) H<sub>2</sub>Se

3. The two families of the periodic table containing the most reactive elements are the halogens and the alkali metals. Elements in each of these families are within one electron of being isoelectronic with the nearest noble gas and readily either gain or lose one electron respectively to achieve that electron arrangement.

### 6.1 Review Questions (p. 297)

1. The most important thing to know about the periodic table is that elements in the same chemical family have similar chemical properties.
2. In 1913, data gathered by the young British chemist Henry Moseley, combined with the discovery of isotopes, resulted in the elements of the periodic table being re-ordered according to their atomic numbers rather than their atomic masses.
3. Elements in the same chemical family have similar chemical properties because they have similar outer electron configurations and the valence electrons are the electrons involved in chemical reactions.
4. (a) The most metallic elements are located in the lower left portion of the periodic table.  
(b) The most non-metallic (or least metallic) elements are located in the upper right portion of the periodic table.

5.

Element	Properties	Letter and Symbol
(a) chlorine	found in carbohydrates and an elemental gas in 21% of the atmosphere	(e) O
(b) silver	soft conductor that reacts explosively with water producing H <sub>2</sub> gas	(d) Cs
(c) neon	less than 1 ounce of this solid radioactive nonconductor exists on Earth	(h) At
(d) cesium	waxy yellow solid non-metal found in match heads, fertilizers, and detergents	(f) P
(e) oxygen	blue-gray metalloid used extensively in the computer industry	(g) Si
(f) phosphorus	very reactive green gas used in the trenches in World War I	(a) Cl
(g) silicon	shiny solid that is the best conductor of heat and electricity	(b) Ag
(h) astatine	invisible unreactive gas used in lasers and some electric street signs	(c) Ne

6.

Element Properties	Family Number
unreactive gas used in electric street signs comprising 0.93 % of the atmosphere	18
shiny multivalent solid, good conductor, forms coloured compounds	6
soft silvery solid, good conductor, reacts vigorously with water	1
gray-white metalloid predicted by Mendeleev and discovered in 1886	14
reactive metal present in bones and teeth possessing two valence electrons	2
yellow-green gaseous non-metal and the most reactive of all the elements	17

7. Properties of metals include:

- solids at room temperature, except for mercury, which is a liquid
- generally shiny or lustrous when freshly cut or polished
- good conductors of heat and electricity
- generally malleable, which means they can be rolled or hammered into thin sheets
- generally ductile, which means they can be rolled or stretched into wires
- generally flexible as thin sheets or wires
- during chemical changes, tend to give up electrons relatively easily to form cations

8. Properties of non-metals include:

- usually gases or brittle solids at room temperature, except for liquid bromine
- solid non-metals can range in appearance from dull or lustrous and translucent to opaque
- poor conductors of heat and electricity
- during chemical changes, they tend to gain electrons from metals to form anions or share electrons with other non-metals.

9. Except where large ionic charges result, atoms of the main group elements will tend to give up or gain as many electrons as are necessary to acquire the valence electron configuration of the nearest noble gas.

10. The formulas for the stable ions are:

(a)  $\text{Be}^{2+}$       (b)  $\text{Te}^{2-}$       (c)  $\text{Cs}^+$       (d)  $\text{Ra}^{2+}$       (e)  $\text{Ga}^{3+}$       (f)  $\text{Se}^{2-}$       (g)  $\text{In}^{3+}$

11. (a) Properties of the alkali metals include:

- all soft, silvery solids
- the most reactive of all metals
- the oxide compounds of the alkali metals dissolve in water to produce strongly basic solutions
- all corrode rapidly in air to a dull gray appearance, react vigorously with water to produce hydrogen gas
- readily form compounds with non-metals
- readily lose that outer electron to form 1+ cations and so assume the electron configuration of previous noble gas

(b) Properties of the alkaline earth metals include:

- silver-coloured reactive metals.
- not as reactive as the alkali metals, but also readily form compounds with non-metals
- their oxides are also alkaline in solution but unlike alkali compounds, a number of group 2 compounds have a low solubility in water
- readily form 2+ cations by losing those two valence electrons and so will achieve the identical electron configuration of the nearest noble gas

(c) Properties of the halogens include:

- the most reactive family of elements in the periodic table
- the only chemical family in which all three states of matter are represented
- elemental halogens exist as diatomic molecules
- readily form compounds with metals, hydrogen, carbon, and other non-metals
- when reacting with metals, halogens typically gain a single electron forming 1- anions
- when reacting with non-metals, halogens will often share valence electrons

(d) Properties of the noble gases include:

- colourless gases
- generally unreactive
- all of the noble gases, except helium, have filled s and p sublevels and thus have “stable octets”

12.

Group 2	Core Notation	Group 17	Core Notation	Group 18	Core Notation
Be	[He] $2s^2$	F	[He] $2s^2 2p^5$	He	$1s^2$ (no core notation)
Mg	[Ne] $3s^2$	Cl	[Ne] $3s^2 3p^5$	Ne	[He] $2s^2 2p^6$
Ca	[Ar] $4s^2$	Br	[Ar] $4s^2 4p^5$	Ar	[Ne] $3s^2 3p^6$
Sr	[Kr] $5s^2$	I	[Kr] $5s^2 4d^{10} 5p^5$	Kr	[Ar] $4s^2 3d^{10} 4p^6$
Ba	[Xe] $6s^2$	At	[Xe] $6s^2 4f^{14} 5d^{10} 6p^5$	Xe	[Kr] $5s^2 4d^{10} 5p^6$
Ra	[Rn] $7s^2$			Rn	[Xe] $6s^2 4f^{14} 5d^{10} 6p^6$

## 6.2 Periodic Trends — Regular Changes in Elemental Properties

### Warm Up (p. 299)

1. The negatively charged electron clouds in atoms are held in place by the attractive force of positively charged protons in the nuclei of those atoms.
2. We might expect that the sizes of atoms would decrease if this attractive force increased in strength.
3. Increasing the positive charge on the nucleus should increase the attractive force on the electron cloud. This would occur as an element's atomic number increased.
4. Fewer protons in the nucleus would exert less of an attractive force on an atom's electron cloud, but we might also expect that as the distance between the nucleus and the valence electrons increased the attractive force exerted on that outer charge cloud by the nucleus would decrease.

### Quick Check (p. 302)

1. As we move left to right across a chemical period in the periodic table, the energy level remains the same and so the attractive force that the valence electrons experience from the nucleus is the predominant force operating that influences the size of the atoms. This force increases as we move across the period and so the atoms generally become smaller.
2. As we move down a chemical family, the number of energy levels present increases in atoms. Even though the nuclear charge also increases, the increasing energy levels are the predominating factor determining atomic size. As a result, the attractive force felt by the valence electrons decreases and the atomic size increases.
3. Effective shielding is most prevalent moving down a chemical family rather than across a chemical period. The evidence of this is that the sizes of atoms increase as we move down a chemical family and decrease as we move left to right across a chemical period.

### Practice Problems — Trends in Ionization Energy (p. 305)

- The elements ranked in order of decreasing ionization energy are:  
argon > sulphur > silicon > aluminum > magnesium > sodium > rubidium > cesium
- Members of this chemical family have the highest  $IE_1$  in their period. 18  
Members of this chemical family have the lowest  $IE_1$  in their period. 1  
Members of this chemical period have the highest  $IE_1$  in their family. The uppermost period in that family  
Members of this chemical period have the lowest  $IE_1$  in their family. 7
- If 2s electrons spend more of their time nearer the nucleus than 2p electrons, then 2s electrons are capable of partially shielding 2p electrons from the nuclear charge. This should make it easier to remove boron's single 2p electron, which explains boron's lower first ionization energy.

### 6.2 Activity (p. 308)

#### Procedure

1.

	Moving Across a Period	Moving Up a Chemical Family
Atomic Size	decreases	decreases
Ionization Energy	increases	increases
Electronegativity	increases	increases

### 6.2 Review Questions (p. 309)

- The regular and predictable changes in elemental properties as we move across a period or down a chemical family in the periodic table are known as periodic trends.
- The quantum mechanical model tells us that the outer boundaries of an atom are not hard and definite, but rather are the edges of charge clouds enclosing the regions of highest probability of finding an atom's outer electrons.
- Ionization energy and electronegativity both increase as we move left to right across a period and move up a chemical family in the periodic table.
- As we move in any direction horizontally across a period or vertically in a chemical family, the trend in atomic size is generally opposite to the trends seen in ionization energy and electronegativity.
- The valence electrons of both lithium and fluorine are in the second energy level and as a result, those valence electrons experience a similar amount of electron shielding. Because fluorine's nucleus has six more protons, the effective nuclear charge ( $Z_{\text{eff}}$ ) seen by fluorine's valence electrons is much greater

than that experienced by lithium's single valence electron. As a result, fluorine atoms are smaller than lithium atoms.

6. The largest atoms are located in the lower left region of the periodic table and the smallest atoms are located in the upper right region of the periodic table.

7. (a) A cation will always be smaller than its parent neutral atom because of increased attraction of the outer electrons for the nucleus and decreased repulsion of the electrons for each other.

(b) An anion will always be larger than its parent neutral atom because of decreased attraction of the outer electrons for the nucleus and increased repulsion of the electrons for each other.

8. Inner or "core" electrons are effective at shielding the outer electrons from the attractive force of the nucleus and exerting a repulsive force on those outer clouds. As the extent of that shielding increases, (which occurs when descending a chemical family), the atomic sizes increase and the ionization energies decrease.

9.

Where on the Periodic Table Elements Show:	
Largest Atomic Radii	<b>lower left</b>
Smallest Atomic Radii	<b>upper right</b>
Lowest Ionization Energy	<b>lower left</b>
Highest Ionization Energy	<b>upper right</b>
Lowest Electronegativity	<b>lower left</b>
Highest Electronegativity	<b>upper right</b>

10. Lithium's electron configuration is  $1s^2 2s^1$ . Lithium's single valence electron is shielded from the nuclear charge by the inner or core  $1s^2$  electrons. It is therefore relatively easy to remove that outer electron and so we see that the first ionization energy is low. After that electron is removed however, the next electron removed is an inner core electron, which is not shielded at all from the nuclear charge. As a result, lithium's second ionization energy is approximately 14 times greater than the first ionization energy!

11. When such reactions occur, elements with low ionization energies and low electronegativities will tend to lose valence electrons to elements whose ionization energies and electronegativities are high.

12. When such reactions occur, elements with high ionization energies and high electronegativities will tend to gain valence electrons from elements whose ionization energies and electronegativities are low. *(The results of the reactions described in questions 11 and 12 will be the formation of cations by atoms losing valence electrons and the formation of anions by atoms gaining valence electrons.)*

13. When two non-metal atoms, each with relatively high ionization energies and electronegativities react, as both attract their valence electrons strongly and neither have a tendency to lose them, they are likely to form bonds by sharing valence electrons rather than by transferring them.

14. Electron configuration for nickel:  $[\text{Ar}] 3d^8 4s^2$

Electron configuration for zinc:  $[\text{Ar}] 3d^{10} 4s^2$

Notice that when the electron configurations are written in order of increasing sublevel size, rather than increasing energy, we see that zinc has a filled 3d sublevel with 2 more electrons in the inner 3d cloud than nickel. These 2 additional electrons increase the effective shielding of zinc's 2 outer electrons and thus reduce the attractive force they feel from the nucleus. As a result, zinc atoms are larger than nickel atoms.

### 6.3 Describing Chemical Bonding

#### Warm Up (p. 311)

1. An atom's outermost electrons that take part in chemical bonding are known as valence electrons.
2. Electronegativity is defined as the relative ability of a bonded atom to attract shared electrons to itself.
3. (a) Elements that tend to lose outer electrons most easily during chemical changes are located in the lower left region of the periodic where.  
(b) Elements that tend to gain outer electrons most easily during chemical changes are located in the upper right region of the periodic table.

#### Quick Check (p. 314)

1. When metal atoms react with non-metal atoms, metal atoms tend to lose one or more valence electrons to non-metal atoms forming metal cations and non-metal anions.
2. Three types of chemical bonds based on the different elements involved:

Atoms Involved in Chemical Bond	Type of Chemical Bond
metal bonded to non-metal	ionic bond
non-metal bonded to non-metal	covalent bond
metal bonded to metal	metallic bond

3. Ionic bond formation is typically associated with metals from groups 1 and 2 reacting with non-metals from groups 16 and 17 of the periodic table.



**Practice Problems — Ionic Compounds (p. 315)**

1. (a) BaBr<sub>2</sub>                      (b) BeO                      (c) Sr<sub>3</sub>N<sub>2</sub>                      (d) MgCl<sub>2</sub>                      (e) FrF

2.

(a) RbF	As $\Delta\text{EN} = 3.2$ bond is ionic
(b) RaCl <sub>2</sub>	As $\Delta\text{EN} = 2.1$ bond is ionic
(c) KBr	As $\Delta\text{EN} = 2.0$ bond is ionic
(d) Na <sub>2</sub> O	As $\Delta\text{EN} = 2.6$ bond is ionic

3. (a) Na<sub>3</sub>N  $\Delta\text{EN} = 2.1$   
 (b) SrBr<sub>2</sub>  $\Delta\text{EN} = 1.8$   
 (c) LiCl  $\Delta\text{EN} = 2.0$   
 (d) CsF  $\Delta\text{EN} = 3.3$   
 (e) Rb<sub>2</sub>O  $\Delta\text{EN} = 2.7$

Compounds in order of increasing ionic bond character: SrBr<sub>2</sub> < LiCl < Na<sub>3</sub>N < Rb<sub>2</sub>O < CsF

**Practice Problems — Comparing Types of Chemical Bonds, p. 319**

1. (a) A compound with an ionic bond: NaCl ( $\Delta\text{EN} = 2.1$ )  
 (b) A compound with a polar covalent bond: AlN ( $\Delta\text{EN} = 1.5$ )  
 (c) A compound with a non-polar covalent bond NCl<sub>3</sub> ( $\Delta\text{EN} = 0$ )

2. (a) H<sub>2</sub>O  $\Delta\text{EN} = 1.4$   
 (b) PCl<sub>3</sub>  $\Delta\text{EN} = 0.9$   
 (c) Cl<sub>4</sub>  $\Delta\text{EN} = 0$   
 (d) SiO<sub>2</sub>  $\Delta\text{EN} = 1.7$   
 (e) AlN  $\Delta\text{EN} = 1.5$

Compounds in order of most equal to most unequal electron sharing: Cl<sub>4</sub> PCl<sub>3</sub> H<sub>2</sub>O AlN SiO<sub>2</sub>

3.

Elements Present	Formula	$\Delta\text{EN}$ Value	Nature of Bonds	Atom Possessing Greater Electron Density
C and S	CS <sub>2</sub>	0	covalent	neither
B and Cl	BCl <sub>3</sub>	1.0	polar covalent	chlorine
Al and O	Al <sub>2</sub> O <sub>3</sub>	2.0	ionic	oxygen
N and I	NI <sub>3</sub>	0.5	polar covalent	nitrogen
Ca and F	CaF <sub>2</sub>	3.0	ionic	fluorine

**6.3 Activity (p. 321)****Results and Discussion**

1. Some metal–non-metal combinations result in polar covalent bonds.

Examples include:

AlCl<sub>3</sub> ( $\Delta\text{EN} = 1.5$ )

GaBr<sub>3</sub> ( $\Delta\text{EN} = 1.4$ )

Some other examples are:

BeBr<sub>2</sub> ( $\Delta EN = 1.3$ )

MgS ( $\Delta EN = 1.3$ )

Ca<sub>3</sub>P<sub>2</sub> ( $\Delta EN = 1.1$ )

### 6.3 Review Questions (p. 322)

1. (a) The energy associated with the bonded atoms must be less than when the atoms are apart.

(b) This tells us that the attractive forces existing between the bonded atoms exceed the repulsive forces.

2. An ionic crystal lattice is the three-dimensional symmetrical arrangement of cations and anions in a solid ionic crystal. The vast number of interionic forces present in such a crystal must be overcome to melt an ionic compound and this explains the high melting points of such compounds.

3. A crystal lattice shows us that no neutral independent molecules exist in ionic compounds. The formulas that we write simply represent the smallest whole number ratios of cations to anions that exist in ionic compounds.

4. (a) The attractive forces associated with ionic bonds are the electrostatic forces between positively charged cations and negatively charged anions.

(b) The attractive forces associated with covalent bonds are the electrostatic forces between negatively charged electrons and adjacent positively charged nuclei.

5. (a) Similarities between ionic and covalent bonds include:

- Both involve valence electron clouds.
- Both involve electrostatic attractions between oppositely charged species.
- The formation of both bonds begins with the valence electrons of two atoms experiencing the attractive force of adjacent positive nuclei.
- Ionic and covalent bonds can both be strong

(b) Differences between ionic and covalent bonds include:

- Ionic bonds involve the transfer of valence electrons forming ions whereas covalent bonds involve the sharing of valence electrons with no ion formation.
- Ionic bonds form only between metals and non-metals whereas covalent bonds usually (but not always) form between two non-metals.
- Ionic bonding does not result in molecule formation whereas covalent bonding usually does.
- The attractive forces associated with covalent bonds are the electrostatic forces between negatively charged electrons and adjacent positively charged nuclei.
- The attractive forces associated with ionic bonds are the electrostatic forces between positively charged cations and negatively charged anions.

6.

Elements	Compound Formula	$\Delta\text{EN}$ Value	Nature of Bonds Present
(a) rubidium and oxygen	$\text{Rb}_2\text{O}$	2.7	ionic
(b) strontium and bromine	$\text{SrBr}_2$	1.8	ionic
(c) carbon and sulphur	$\text{CS}_2$	0	covalent
(d) silicon and chlorine	$\text{SiCl}_4$	1.2	polar covalent

7.  $\Delta\text{EN}$  for  $\text{MgS}$  = 1.3

$\Delta\text{EN}$  for  $\text{H}_2\text{O}$  = 1.4

We see that the bonds in water actually possess slightly more ionic character than those in magnesium sulphide even though the former compound contains two non-metals and the latter compound contains a metal and a non-metal.

8. The formula for glucose and other molecular compounds does not represent a ratio as do ionic formulas. Rather, they represent the actual number of atoms of each element existing in an individual molecule of the compound.

9. The melting of molecular covalent compounds does not involve the breaking of covalent chemical bonds within the molecules, but rather overcoming the relatively weak attractive forces between those molecules in the solid phase.

10. Diamond is a type of covalent substance is known as a network covalent solid. Rather than consisting of individual molecules as molecular covalent compounds do, these substances are held together by covalent bonds that extend throughout the entire sample. This means that the “molecule” is literally as big as the sample itself. To melt such a substance, all of the covalent bonds within this giant molecule must be broken and this accounts for the very high melting point.

11. The bond in  $\text{HCl}$  is a polar covalent bond ( $\Delta\text{EN} = 0.9$ ) whereas the bond in  $\text{N}_2$  must be pure covalent ( $\Delta\text{EN} = 0$ ). This means that electron density in the  $\text{HCl}$  molecule is concentrated on the chlorine side of the molecule making that end of the molecule somewhat negative and the hydrogen end of the molecule somewhat positive. We might expect therefore, that  $\text{HCl}$  molecules would attract each other more strongly than  $\text{N}_2$  molecules where the electron density is evenly distributed throughout the molecule. (The melting point of  $\text{HCl}$  is  $-114.2^\circ\text{C}$  and the melting point of  $\text{N}_2$  is  $-210^\circ\text{C}$ .)

## 6.4 Lewis Structure Diagrams

### Warm Up (p. 324)

1. Elements in the same chemical family display similar behaviour because they have the same number of electrons in their outermost or valence electron shells.

2. There are eight main group families including Groups 1, 2, 13, 14, 15, 16, 17, and 18.

Begin with the number 1 at Group 1 and count consecutively from 1 to 8. Each number will then represent the number of valence electrons present in the atoms of each family in the Group in question.

This gives us the following table:

Main Group Number	1	2	13	14	15	16	17	18
Valence Electrons	1	2	3	4	5	6	7	8

3. The term “stable octet” refers to a filled valence shell for an atom or ion consisting of 8 electrons in the outer s and p sublevels. This gives the valence shell the “ $ns^2np^6$ ” configuration.

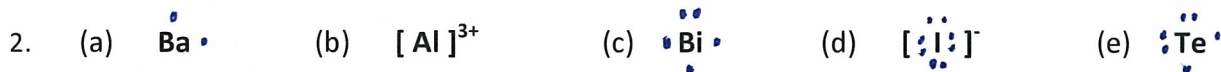
### Quick Check (p. 326)

1. The dots in a Lewis dot symbol for an atom or ion represent the number of electrons in the valence shell – the number of electrons in the outer s and p sublevels.

2. For the elements in Groups 1, 2, and 13 the number of dots tells us the number of electrons that each atom loses when forming a cation – this also equals the magnitude of the resulting positive charge on that cation.

3. In a correctly drawn Lewis symbol for a nonmetal, the number of unpaired dots shown represents either the number of electrons that atom must gain when forming an anion, or the number the electrons the atom must share to complete its octet when forming covalent bonds.

### Practice Problems -- Building Lewis Structures (p. 324)

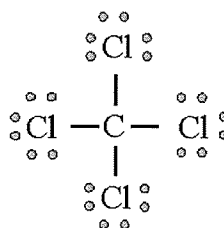


**Quick Check** (p. 330)

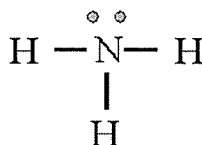
1. Because the elements Be, B, and Al each possess fewer than 4 valence electrons, (2, 3, and 3 respectively) they are unable to achieve a valence octet of four electron pairs by sharing all of their valence electrons when forming covalent compounds.
2. The skeleton "S – H – H " is incorrect because it shows a central hydrogen with two bonds. Hydrogen normally forms only one bond by sharing its single electron.
3. The total number of valence electrons in each of the following molecules are:
  - (a)  $\text{H}_2\text{Se}$   $(2 \times 1) + (6) = 8$  valence electrons
  - (b)  $\text{CCl}_4$   $(4) + (4 \times 7) = 32$  valence electrons
  - (c)  $\text{NF}_3$   $(5) + (3 \times 7) = 26$  valence electrons
  - (d)  $\text{PCl}_5$   $(5) + (5 \times 7) = 40$  valence electrons
  - (e)  $\text{SF}_6$   $(6) + (6 \times 7) = 48$  valence electrons

**Practice Problems -- Drawing Lewis Structures for Molecules** (p. 331)

1. Lewis structure for carbon tetrachloride:

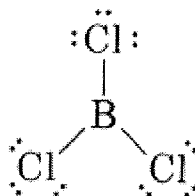


- 2.(a) Lewis structure for ammonia:



- (b) There are 3 bonding electron pairs  
and 1 lone pair of electrons.

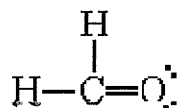
- 3.(a) Lewis structure for boron trichloride:



- (b) The central boron possesses an  
incomplete octet of 6 electrons.

Practice Problems -- Lewis Structures Containing Multiple Bonds and Expanded Octets (p. 335)

1. Lewis structure for formaldehyde:



2. Lewis structure for carbon monoxide:



3. Lewis structure for bromine trifluoride:



6.4 Activity answers are on page 17

6.4 Review Questions (p. 339)

1.

Incomplete Valence Octet	Valence Octet	Expanded Valence Octet
H	C	Si
Be	N	P
B	O	S
Al	F	Cl

2. Helium and neon are both invisible totally unreactive gases whose similar behaviour qualifies each of them as members of the family which includes other unreactive gases. They do have a different number of outer electrons, but the outer electron cloud of each is a particularly stable configuration.

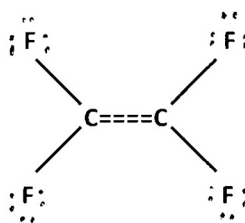
Helium only has 2 electrons, and those are represented by the 2 dots in its Lewis structure. That single pair of outer electrons fills the 1s sublevel and also completes the 1<sup>st</sup> energy level, which explains why helium is unreactive.

Neon has 8 outer electrons represented as 4 pairs of dots. Those 8 electrons fill the 2s and 2p sublevels, complete the 2<sup>nd</sup> energy level, and give neon a stable valence octet consisting of 4 pairs of electrons. As a result, neon is also unreactive.

3.

Element Pair	Probable Central Atom	Probable Peripheral Atom
(a) phosphorus and chlorine	phosphorus	chlorine
(b) nitrogen and oxygen	nitrogen	oxygen
(c) carbon and sulphur	carbon	sulphur
(d) nitrogen and hydrogen	nitrogen	hydrogen
(e) oxygen and fluorine	oxygen	fluorine

4.



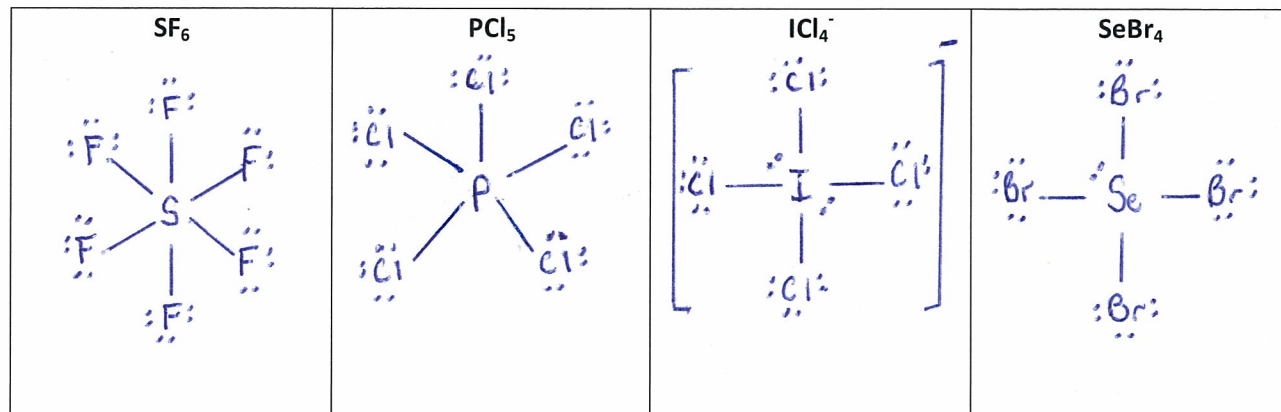
5.

OF <sub>2</sub>	H <sub>2</sub> S	PCl <sub>3</sub>	CCl <sub>2</sub> F <sub>2</sub>

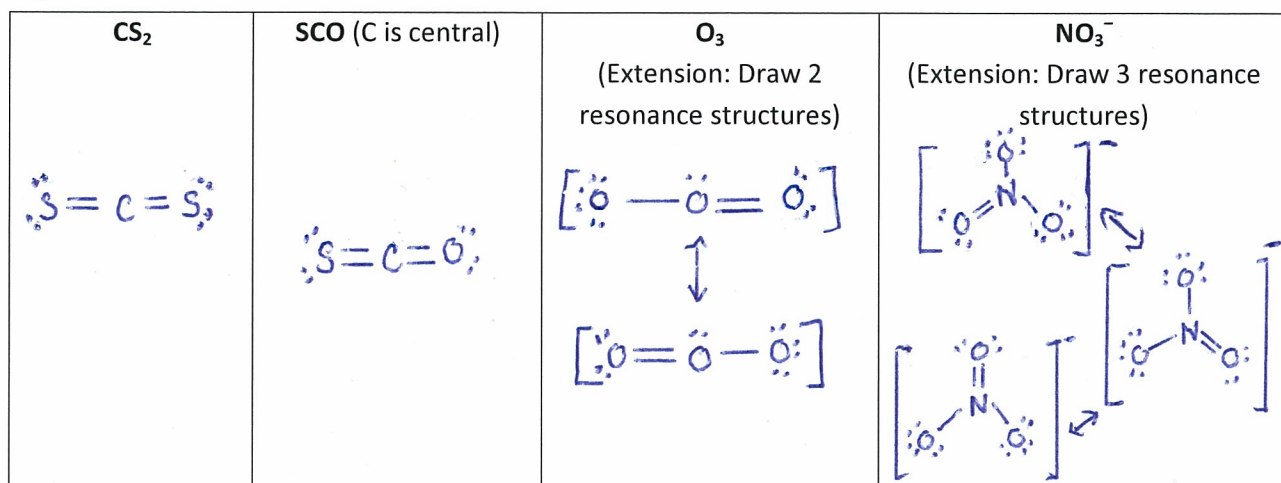
6.

OH <sup>-</sup>	AlH <sub>4</sub> <sup>-</sup>	CN <sup>-</sup>

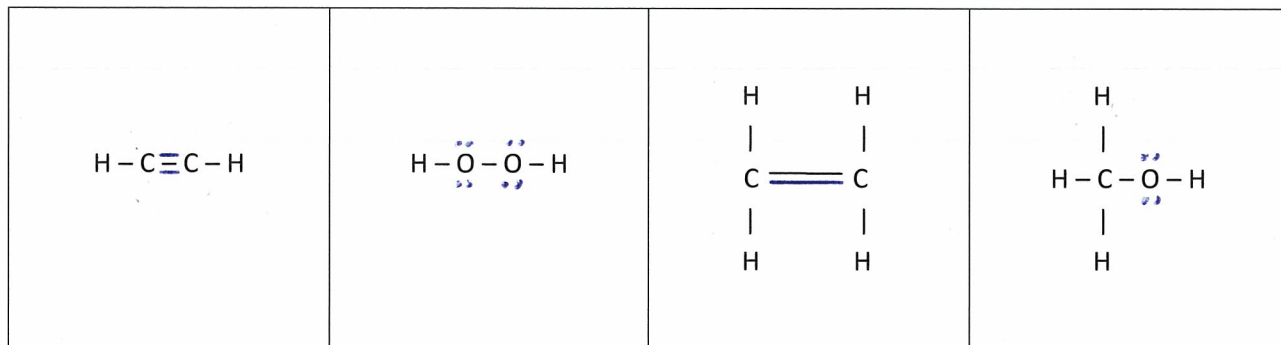
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8.



9.





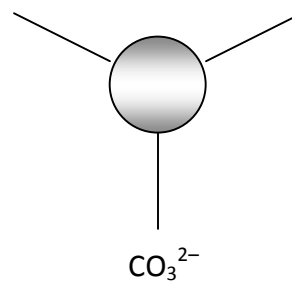
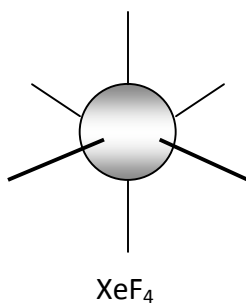
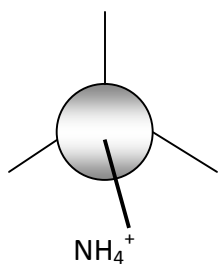
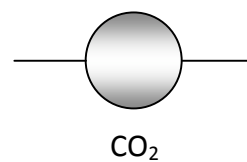
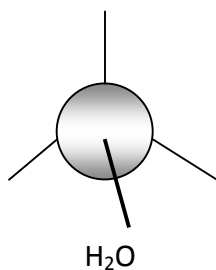
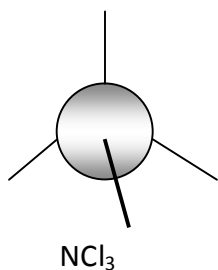
## 6.4 Activity (p. 337)

### Procedure

1.

1. Chemical Formula	2. Number of Atoms Bonded to Central Atom	3. Number of Lone Pairs on Central Atom	4. Sum of Columns 2 and 3
$\text{NCl}_3$	3	1	$3 + 1 = 4$
$\text{H}_2\text{O}$	2	2	4
$\text{CO}_2$	2	0	2
$\text{NH}_4^+$	4	0	4
$\text{XeF}_4$	4	2	6
$\text{CO}_3^{2-}$	3	0	3

4.



## 6.5 The Shape and Behaviour of Molecules (Extension)

### Warm Up (p. 341)

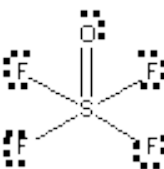
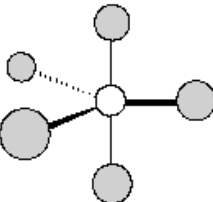
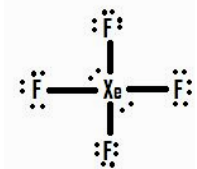
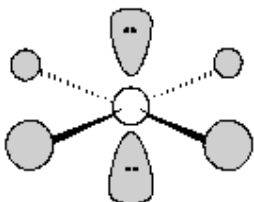
1. An element's electronegativity tells us how well it attracts a pair of bonded electrons to itself in a chemical bond.
2. A polar covalent bond is a bond in which electron pairs are unequally shared between two different atoms. The atom having the higher electronegativity pulls the bonded electrons closer to itself and away from the adjacent atom resulting in a partial positive charge on the atom with the lower electronegativity and a partial negative charge on the atom with the higher electronegativity.
3. Substances composed of polar molecules would be expected to have higher melting and boiling points because the negative ends of the polar molecules will be attracted to the positive ends of neighbouring polar molecules.

### Quick Check (p. 346)

1. Any group of valence electrons associated with a central atom (such as a lone pair, bonding pair, or multiple pairs involved in a double or triple bond) will tend to orient themselves in three-dimensional space around that atom so as to minimize the repulsion between them.
2. **BP-BP < BP-LP < LP-LP**  
  
least intense → most intense
3. In methane ( $\text{CH}_4$ ) all of the electron pair interactions are BP-BP because methane is an  $\text{AX}_4$  molecule. As these are the least intense, the H-C-H bond angles are those of a regular tetrahedron ( $109.5^\circ$ ). Ammonia ( $\text{NH}_3$ ) is an  $\text{AX}_3\text{E}$  molecule and the lone pair on the central nitrogen exerts more of a repulsive force on the N-H bonding electron pairs than a bonding pair of electrons would and so the H-N-H bond angles are smaller at  $107^\circ$  because of that LP-BP interaction. Water is an  $\text{AX}_2\text{E}_2$  molecule and the two lone pairs on the central oxygen exert more repulsive forces on the O-H bonding electrons than a single lone pair would. In water, LP-LP as well as LP-BP interactions exists. As a result, the H-O-H bond angle is even smaller at  $104.5^\circ$ .

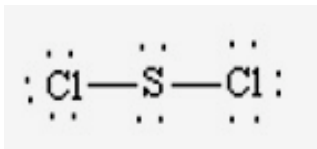
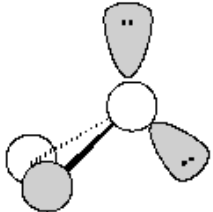
Practice Problems (p. 348)

1.

Lewis Structure	$AX_mE_n$ Notation	Molecular Shape (Name and Diagram)
<p>(a)</p> 	$AX_6$	 <p>octahedral</p>
<p>(b)</p> 	$AX_4E_2$	 <p>square planar</p>

2.

Chemical Formula	Lewis Structure	$AX_mE_n$ Notation	Molecular Shape (Name and Diagram)
(a)  $CCl_4$		$AX_4$	 tetrahedral
(b)  $PF_3$		$AX_3E$	 trigonal pyramidal

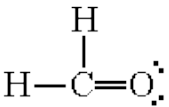
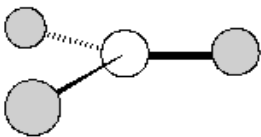
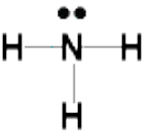
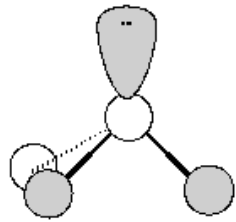
(c)			
$\text{SCl}_2$		$\text{AX}_2\text{E}_2$	 bent or angular

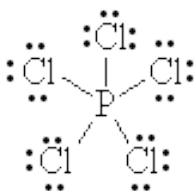
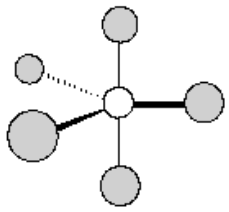
### Practice Problems (p. 353)

1.

<i>Symmetric Molecules</i>		<i>Asymmetric Molecules</i>	
$\text{AX}_m\text{E}_n$ Notation	Shape of Molecule	$\text{AX}_m\text{E}_n$ Notation	Shape of Molecule
$\text{AX}_2$	linear	$\text{AX}_2\text{E}$	bent or angular
$\text{AX}_3$	trigonal planar	$\text{AX}_2\text{E}_2$	bent or angular
$\text{AX}_4$	tetrahedral	$\text{AX}_3\text{E}$	trigonal pyramidal
$\text{AX}_5$	trigonal bipyramidal	$\text{AX}_4\text{E}$	seesaw
$\text{AX}_6$	octahedral	$\text{AX}_3\text{E}_2$	t-shaped
$\text{AX}_2\text{E}_3$	linear	$\text{AX}_5\text{E}$	square pyramidal
$\text{AX}_4\text{E}_2$	square planar		

2.

Lewis Structure	$\text{AX}_m\text{E}_n$ Notation	Molecular Shape (Name and Diagram)	Polar Molecule? (Yes / No)
(a) 	$\text{AX}_3$	 trigonal planar	Yes
(b) 	$\text{AX}_3\text{E}$	 trigonal pyramidal	Yes

(c)			 <p>trigonal bipyramidal</p>	No
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### 6.5 Activity (p. 359)

#### Procedure

3.

AX <sub>m</sub> E <sub>n</sub> Notation	Sample Molecule	Name of Shape	Polar Molecule? Yes/No
AX <sub>2</sub>	CO <sub>2</sub>	linear	no
AX <sub>3</sub>	BF <sub>3</sub>	trigonal planar	no
AX <sub>2</sub> E	SO <sub>2</sub>	bent or angular	yes
AX <sub>4</sub>	CH <sub>4</sub>	tetrahedral	no
AX <sub>3</sub> E	NH <sub>3</sub>	trigonal pyramidal	yes
AX <sub>2</sub> E <sub>2</sub>	H <sub>2</sub> O	bent or angular	yes
AX <sub>5</sub>	PCl <sub>5</sub>	trigonal bipyramidal	no
AX <sub>4</sub> E	SF <sub>4</sub>	see-saw	yes
AX <sub>3</sub> E <sub>2</sub>	BrF <sub>3</sub>	t-shaped	yes
AX <sub>2</sub> E <sub>3</sub>	XeF <sub>2</sub>	linear	no
AX <sub>6</sub>	SF <sub>6</sub>	octahedral	no
AX <sub>5</sub> E	BrF <sub>5</sub>	square pyramidal	yes
AX <sub>4</sub> E <sub>2</sub>	XeF <sub>4</sub>	square planar	no

4. As XeF<sub>4</sub> is a symmetric square planar molecule, any bond dipoles that might exist would cancel out. This means that the electronegativities of the elements in the molecule and also the corresponding  $\Delta EN$  values are irrelevant when assessing the molecular polarity.

### 6.5 Review Questions (p. 360)

1. (a) The letters stand for: **V**alence **S**hell **E**lectron **P**air **R**epulsion.

(b) VSEPR theory allows us to use two-dimensional Lewis diagrams for molecules to quite accurately predict the three-dimensional shapes of those molecules.

2. As lone-pair electrons are attracted to only one atomic nucleus, they are held less tightly than bonding electron groups. Their electron clouds therefore occupy more space and exert more repulsive force on bonding electron groups than those groups exert on each other.

3. Both  $\text{BF}_3$  and  $\text{CH}_2\text{O}$  have the same shape because both of them are  $\text{AX}_3$  molecules. Even though the peripheral atoms in each molecule possess different numbers of non-bonding electrons and one of the AX bonds on  $\text{CH}_2\text{O}$  is a double bond, the shapes are identical.

4.

$\text{AX}_m\text{E}_n$ Notation	Molecular Shape	$\text{AX}_m\text{E}_n$ Notation	Molecular Shape
$\text{AX}_3$	angular	$\text{AX}_4\text{E}$	T-shaped
$\text{AX}_2\text{E}_3$	trigonal bipyramidal	$\text{AX}_2\text{E}$	octahedral
$\text{AX}_4$	trigonal pyramidal	$\text{AX}_3\text{E}_2$	square pyramidal
$\text{AX}_3\text{E}$	trigonal planar	$\text{AX}_6$	square planar
$\text{AX}_2\text{E}_2$	tetrahedral	$\text{AX}_5\text{E}$	angular
$\text{AX}_5$	linear	$\text{AX}_4\text{E}_2$	seesaw

5. (a) The  $\text{X}-\text{A}-\text{X}$  bond angles in ammonia will be smaller than in methane ( $107^\circ$  vs.  $109.5^\circ$ ) because ammonia is an  $\text{AX}_3\text{E}$  molecule and methane is an  $\text{AX}_4$  molecule. The lone pair on the central nitrogen in ammonia occupies more space than the bonded pairs on the central carbon in methane and will therefore force the bonded electron pairs in ammonia closer together.

(b) Methane is a symmetrical  $\text{AX}_4$  molecule and is non-polar. The intermolecular forces acting are therefore weak London dispersion forces. Ammonia is an asymmetrical  $\text{AX}_3\text{E}$  polar molecule that exhibits much stronger hydrogen bonding.

6.

$\text{AX}_m\text{E}_n$ Category	$\text{AX}_2$	$\text{AX}_3$	$\text{AX}_4$	$\text{AX}_5$		$\text{AX}_2\text{E}_3$	$\text{AX}_6$	$\text{AX}_4\text{E}_2$
$\text{X}-\text{A}-\text{X}$ Bond Angle	$180^\circ$	$120^\circ$	$109.5^\circ$	$120^\circ$	$180^\circ$	$180^\circ$	$90^\circ$	$90^\circ$

7. The molecules of each compound are asymmetric AX<sub>3</sub>E molecules. If we consider the  $\Delta$ EN values as we move up the family from bottom to top, we see the following:

Compound	SbH <sub>3</sub>	AsH <sub>3</sub>	PH <sub>3</sub>	NH <sub>3</sub>
$\Delta$ EN Value	0.2	0.1	0.0	0.9

Even though asymmetry exists in each of the molecules, because the  $\Delta$ EN values for the first three compounds are so small, all of the bonds in those molecules are non-polar and so the molecules themselves are non-polar. As a result, the intermolecular forces acting between them are weak London dispersion forces. However, the bonds in ammonia are quite polar and so the molecule itself is also polar. As nitrogen is bonded to hydrogen in a polar molecule, hydrogen bonds exist between the ammonia molecules. We would therefore expect ammonia to have the highest boiling point.

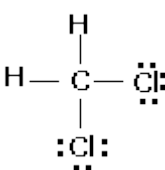
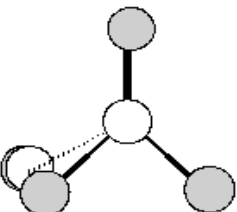
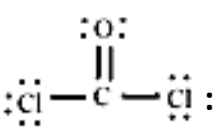
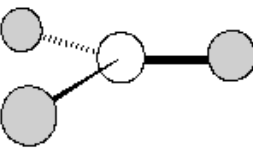
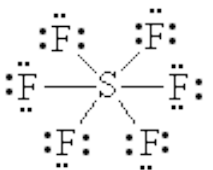
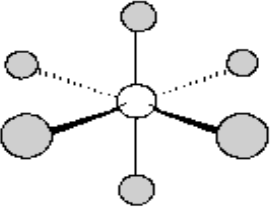
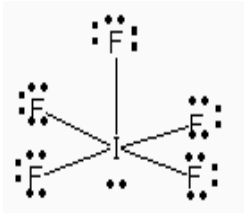
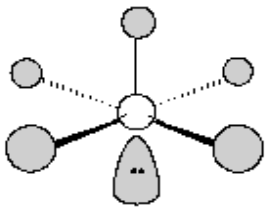
8. Because hydrogen bonds exist between ammonia molecules but only London dispersion forces exist between molecules of nitrogen and hydrogen, the ammonia has a higher boiling point and thus liquefies (condenses) at a higher temperature than nitrogen and hydrogen, which remain in the gaseous phase upon cooling.

9. As a result of hydrogen bonding, water remains liquid until 100°C. This is very important given the fact that most of our Earth and most of our bodies are composed of and require liquid water to survive. The three-dimensional structure of proteins and the base-pairing in the double helix of DNA molecules depend on hydrogen bonds. Hydrogen bond formation in ice makes it less dense than liquid water, which is crucial to aquatic life when bodies of water freeze.

10. I<sub>2</sub> is a relatively large and massive diatomic molecule and the strength of London dispersion forces increase as the size of the molecules involved increases. This is because large electron clouds are more loosely held than smaller clouds and thus more easily deformed or polarized by a nearby dipole than compact tightly held clouds. In addition, large molecules with more surface area have electron clouds that are spread out and so are more easily distorted by neighboring dipoles. As a result, the dispersion forces are strong enough to keep the molecules of I<sub>2</sub> attached to each other even at room temperature.

11. Although the ionic bonds holding a crystal lattice together are strong, when the surface of that lattice is in contact with water, each ion on that surface will attract the oppositely charged end of polar water molecules near them. That attraction between an ion and a polar molecule is called an **ion-dipole force**. These attractive forces soon overcome those existing between the ions themselves and so the crystal structure begins to break down and the ionic compound dissolves. As the ions move away from the lattice surface, they immediately become surrounded or enclosed in what chemists call a hydration shell. Ion-dipole forces are the primary force responsible for the solubility of ionic compounds in water.

12.

Lewis Structure	$AX_mE_n$ Notation	Shape of Molecule (Name and Diagram)	Type of Intermolecular Force Acting Between Molecules
(a)  dichloromethane	$AX_4$	 tetrahedral	dipole-dipole
(b)  phosgene (a gas used in the trenches in WWI)	$AX_3$	 trigonal planar	dipole-dipole
(c)  sulphur hexafluoride	$AX_6$	 octahedral	London dispersion
(d)  iodine pentafluoride	$AX_5E$	 square pyramidal	dipole-dipole