EOT COVERAGE – CHEMISTRY
GRADE 10 A AND 10 HSA

1. **Differentiate between the ground state and excited state of an atom. (Text book pages 14-15)**

GROUND STATE VERSUS EXCITED STATE

GROUND STATE	EXCITED STATE
Ground state is the state in which all electrons in a system are in the lowest possible energy levels	Excited state is any state of the system that has a higher energy than the ground state
Known to be having a "zero" energy	Has a high energy
Highly stable	Highly unstable
Has a long lifetime	Has a short lifetime
The distance between ground state electron and the atomic nucleus is the least possible distance	The distance between the excited state electron and the atomic nucleus is comparatively higher
Electrons are located in the lowest possible energy levels	Electrons are located in higher energy levels

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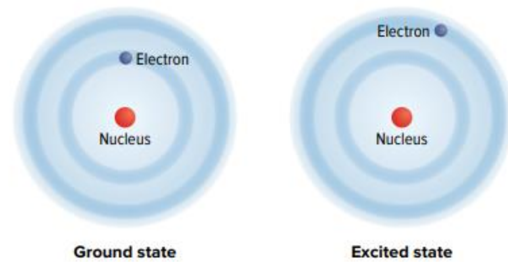


Figure 10 The figure shows an atom that has one electron. Note that the illustration is not to scale. In its ground state, the electron is associated with the lowest energy level. When the atom is in an excited state, the electron is associated with a higher energy level.

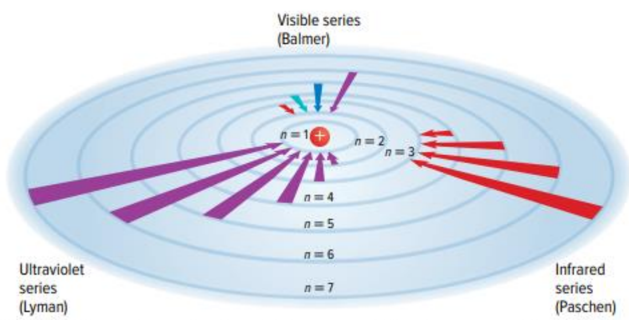


Figure 11 When an electron drops from a higher-energy orbit to a lower-energy orbit, a photon is emitted. The ultraviolet (Lyman), visible (Balmer), and infrared (Paschen) series correspond to electrons dropping to $n = 1$, $n = 2$, and $n = 3$, respectively.

2. **Explain Aufbau principle, Pauli's exclusion principle and Hund's rule. (Text book pages 25,26 and 27)**

Aufbau Principle:
Electrons will occupy the lowest energy level orbital that will receive it.

Pauli Exclusion Principle:
No two electrons within an atom may have an identical set of all four quantum numbers..

Hund's Rule:
Electrons will occupy all empty orbitals in a subshell with single electrons having parallel spins before entering half-filled orbitals.

3. **Use the electron configuration notation, orbital notation and noble gas notation of an element (atomic numbers 1 to 36) to identify the location of the element in the periodic table (period, group, block). (Page 47,49)**

PRACTICE Problems ADDITIONAL PRACTICE

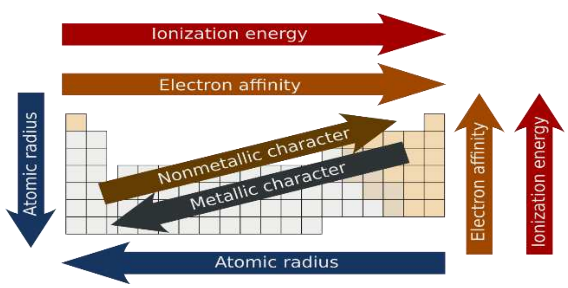
8. Without using the periodic table, determine the group, period, and block of an atom with the following electron configurations.
a. $[Ne]3s^2$ **b.** $[He]2s^2$ **c.** $[Kr]5s^24d^{10}5p^5$

9. What are the symbols for the elements with the following valence electron configurations?
a. s^2d^1 **b.** s^2p^3 **c.** s^2p^6

10. CHALLENGE Write the electron configuration of the following elements.
a. the group 2 element in the fourth period
b. the group 12 element in the fourth period
c. the noble gas in the fifth period
d. the group 16 element in the second period

8. a. group 2; period 3; block s
   b. group 2; period 2; block s
   c. group 17; period 5; block p
9. a. Sc, Y, La, Ac b. N, P, As, Sb, Bi c. Ne, Ar, Kr, Xe, Rn
10. a. $1s^22s^22p^63s^23p^64s^2$
    b. $1s^22s^22p^63s^23p^64s^23d^{10}$
    c. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6$
    d. $1s^22s^22p^4$

4. **Predict the trends of the properties of elements across a period and a group of the periodic table. (Page 50 to 54)**



5. **Describe how ions (cations and anions) form to fulfill the octet rule. (Page 65 to 68)**
6. **Describe how ions (cations and anions) form to fulfill the octet rule. (Page 67)**
   - **Ions are formed when an atom gains or loses electrons from the outermost energy shell.**
   - **The tendency of atoms to have eight electrons in their outer shell is known as the octet rule.**

- Ionic Bond – form between a **metal cation** and **nonmetal anion**
- To follow the octet rule, the cation gives up an electron(s) to the anion, which bonds the two ions together.
- When the two ions bond, they become a **neutrally charged** compound.

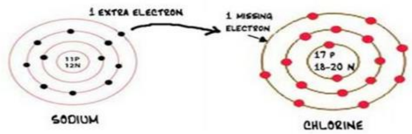


Table 1 Electron-Dot Structures

Group	1	2	13	14	15	16	17	18
Diagram	Li·	·Be·	·B·	·C·	·N·	·O:	:F:	:Ne:

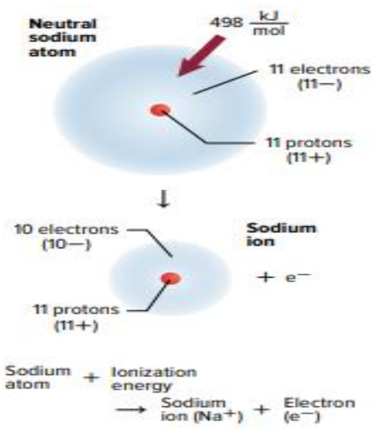


Figure 2 In the formation of a positive ion, a neutral atom loses one or more valence electrons. The atom is neutral because it contains equal numbers of protons and electrons; the ion, however, contains more protons than electrons and has a positive charge.

Table 2 Group 1, 2, and 13 Ions

Group	Configuration	Charge of Ion Formed
1	[noble gas]ns^1	1+ when the s^1 electron is lost
2	[noble gas]ns^2	2+ when the s^2 electrons are lost
13	[noble gas]ns^2np^1	3+ when the s^2p^1 electrons are lost

Negative Ion Formation

Nonmetals, which are located on the right side of the periodic table, easily gain electrons to attain a stable outer electron configuration. Examine **Figure 4**. To attain a noble-gas configuration, chlorine gains one electron, forming an ion with a 1– charge. After gaining the electron, the chloride ion has the electron configuration of an argon atom.

Chlorine atom (Cl)	$1s^22s^22p^63s^23p^5$
Argon atom (Ar)	$1s^22s^22p^63s^23p^6$
Chloride ion (Cl^-)	$1s^22s^22p^63s^23p^6$

An **anion** is a negatively charged ion. To designate an anion, the ending *-ide* is added to the root name of the element. Thus, a chlorine atom becomes a chloride anion.

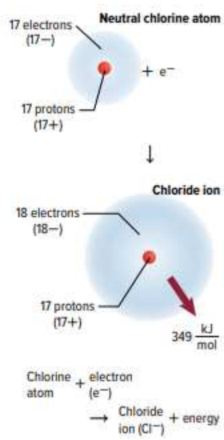


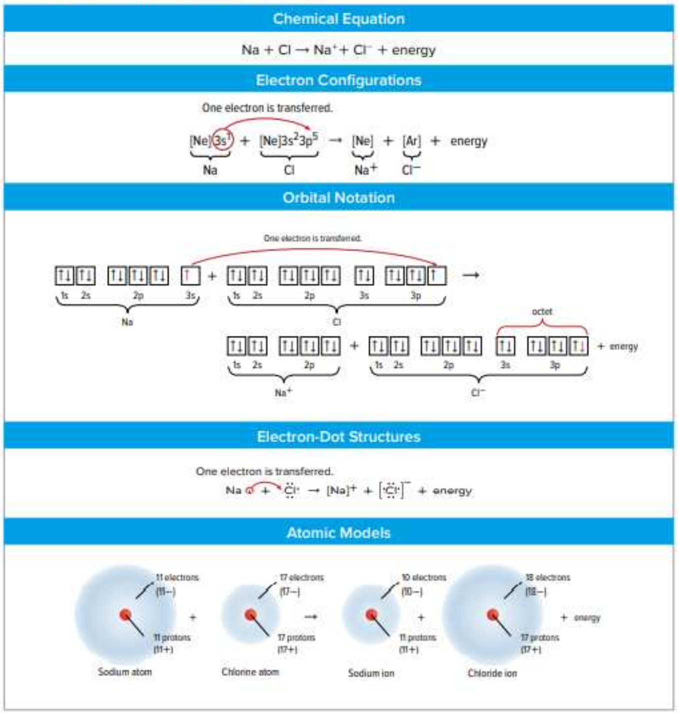
Figure 4 During the formation of the negative chloride ion, a neutral atom gains one electron. The process releases 349 kJ/mol of energy.

Table 3 Group 15–17 Ions

Group	Configuration	Charge of Ion Formed
15	[noble gas]ns^2np^3	3– when three electrons are gained
16	[noble gas]ns^2np^4	2– when two electrons are gained
17	[noble gas]ns^2np^5	1– when one electron is gained

7. Use the Lewis diagram (electron-dot structure) to explain how elements from the periodic table groups combine to form an ionic compound. (Page 70,71)

Table 4 Formation of Sodium Chloride



PRACTICE Problems ADDITIONAL PRACTICE

Explain how an ionic compound forms from these elements.

7. sodium and nitrogen
8. lithium and oxygen
9. strontium and fluorine
10. aluminum and sulfur
11. **CHALLENGE** Explain how elements in the two groups shown on the periodic table at the right combine to form an ionic compound.

7. Three Na atoms each lose 1e⁻, forming 1+ ions. One N atom gains 3e⁻, forming a 3− ion. The ions attract, forming Na_3N. The overall charge on one formula unit of Na_3N is zero.

8. Two Li atoms each lose 1e⁻, forming 1+ ions. One O atom gains 2e⁻, forming a 2− ion. The ions attract, forming Li_2O. The overall charge on one formula unit of Li_2O is zero.

9. One Sr atom loses 2e⁻, forming a 2+ ion. Two F atoms each gain 1e⁻, forming 1− ions. The ions attract, forming SrF_2. The overall charge on one formula unit of SrF_2 is zero.

10. Two Al atoms each lose 3e⁻, forming 3+ ions. Three S atoms each gain 2e⁻, forming 2− ions. The ions attract, forming Al_2S_3. The overall charge on one formula unit of Al_2S_3 is zero.

11. One group 2 atom loses 2e⁻, forming a 2+ ion. Two group 17 atoms each gain 1e⁻, forming 1− ions. The ions attract, forming XY_2, where X represents a group 2 atom and Y represents a group 17 atom.

8. Describe the relationship between lattice energy and the charge of ions. (Page 74,75)

- **Lattice energy increases for ions with higher charges and shorter distances between ions.**

LATTICE ENERGY

Compound	Name:	Charge on each ion	lattice energy (kJ/mol)
NaCl	sodium chloride	1, -1	-787.5
NaBr	sodium bromide	1, -1	-751.4
CaF2	calcium flouride	+2, -1	-2634.7
MgO	magnesium oxide	+2, -2	-3760

Lattice energy

Because the ions in an ionic compound are arranged in a crystal lattice, the energy required to separate 1 mol of the ions of an ionic compound is referred to as the **lattice energy.** The strength of the electrical forces holding ions in place is reflected by the lattice energy. The greater the lattice energy, the stronger the force of attraction.

Lattice energy is directly related to the size of the ions bonded. Smaller ions form compounds with more closely spaced ionic charges. Because the electrostatic force of attraction between opposite charges increases as the distance between the charges decreases, smaller ions produce stronger attractions and greater lattice energies. For example, the lattice energy of a lithium compound is greater than that of a potassium compound with the same anion because a lithium ion is smaller than a potassium ion.

As the lattice energy increases, the ion size decreases, and as the lattice energy decreases, the ion size increases.
The relationship is inverse.

Table 6 Lattice Energies of Some Ionic Compounds

Compound	Lattice Energy (kJ/mol)	Compound	Lattice Energy (kJ/mol)
KI	632	KF	808
KBr	671	AgCl	910
RbF	774	NaF	910
NaI	682	LiF	1030
NaBr	732	$SrCl_2$	2142
NaCl	769	MgO	3795

9. Write the chemical name of an ionic compound containing monoatomic and polyatomic ions (including oxyanions). (Pages 79 to 81)

10. Write the chemical name of an ionic compound containing monoatomic and polyatomic ions (including oxyanions). (Page 82)

Table 9 Common Polyatomic Ions

Ion	Name	Ion	Name
NH_4^+	ammonium	IO_4^-	periodate
NO_2^-	nitrite	$C_2H_3O_2^-$	acetate
NO_3^-	nitrate	$H_2PO_4^-$	dihydrogen phosphate
OH^-	hydroxide	CO_3^{2-}	carbonate
CN^-	cyanide	SO_3^{2-}	sulfite
MnO_4^-	permanganate	SO_4^{2-}	sulfate
HCO_3^-	hydrogen carbonate	$S_2O_3^{2-}$	thiosulfate
ClO^-	hypochlorite	O_2^{2-}	peroxide
ClO_2^-	chlorite	CrO_4^{2-}	chromate
ClO_3^-	chlorate	$Cr_2O_7^{2-}$	dichromate
ClO_4^-	perchlorate	HPO_4^{2-}	hydrogen phosphate
BrO_3^-	bromate	PO_4^{3-}	phosphate
IO_3^-	iodate	AsO_4^{3-}	arsenate

Ammonium ion (NH_4^+)

Phosphate ion (PO_4^{3-})

Figure 9 Ammonium and phosphate ions are polyatomic; that is, they are made up of more than one atom. Each polyatomic ion, however, acts as a single unit and has one charge.

PRACTICE Problems ADDITIONAL PRACTICE

Write formulas for ionic compounds composed of the following ions.
Use units as a guide to your solutions.

24. sodium and nitrate
25. calcium and chlorate
26. aluminum and carbonate
27. **CHALLENGE** Write the formula for an ionic compound formed by ions from a group 2 element and polyatomic ions composed of only carbon and oxygen.

24. $NaNO_3$
25. $Ca(ClO_3)_2$
26. $Al_2(CO_3)_3$
27. Answers will vary; example: $MgCO_3$

Naming an oxyanion

An **oxyanion** is a polyatomic ion composed of an element, usually a nonmetal, bonded to one or more oxygen atoms. More than one oxyanion exists for some nonmetals, such as nitrogen and sulfur. These ions are easily named using the rules in **Table 10**.

Table 10 Oxyanion Naming Conventions for Sulfur and Nitrogen

• Identify the ion with the greatest number of oxygen atoms. This ion is named using the root of the nonmetal and the suffix *-ate*.
• Identify the ion with fewer oxygen atoms. This ion is named using the root of the nonmetal and the suffix *-ite*.
Examples: NO_3^- nitr*ate*; NO_2^- nitr*ite*; SO_4^{2-} sulf*ate*; SO_3^{2-} sulf*ite*

As shown in **Table 11**, chlorine forms four oxyanions that are named according to the number of oxygen atoms present. Names of similar oxyanions formed by other halogens follow the rules used for chlorine. For example, bromine forms the bromate ion (BrO_3^-), and iodine forms the periodate ion (IO_4^-) and the iodate ion (IO_3^-).

Table 11 Oxyanion Naming Conventions for Chlorine

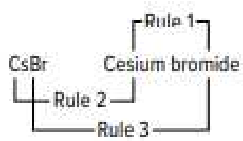
• The oxyanion with the greatest number of oxygen atoms is named using the prefix *per-*, the root of the nonmetal, and the suffix *-ate*.
• The oxyanion with one fewer oxygen atom is named using the root of the nonmetal and the suffix *-ate*.
• The oxyanion with two fewer oxygen atoms is named using the root of the nonmetal and the suffix *-ite*.
• The oxyanion with three fewer oxygen atoms is named using the prefix hypo-, the root of the nonmetal, and the suffix *-ite*.
Examples: ClO_4^- perchlorate; ClO_3^- chlorate; ClO_2^- chlorite; ClO^- hypochlorite

Naming ionic compounds

Chemical nomenclature is a systematic way of naming compounds. Now that you are familiar with chemical formulas, you can use the following five rules to name ionic compounds.

1. Name the cation followed by the anion. Remember that the cation is always written first in the formula.
2. For monatomic cations, use the element name.
3. For monatomic anions, use the root of the element name plus the suffix *-ide*.

   Example:



4. To distinguish between multiple oxidation numbers of the same element, the name of the chemical formula must indicate the oxidation number of the cation. The oxidation number is written as a Roman numeral in parentheses after the name of the cation.

   Note: This rule applies to the transition metals and metals on the right side of the periodic table, which often have more than one oxidation number. See **Table 8** earlier in this lesson. It does not apply to group 1 and group 2 cations, as they have only one oxidation number.

   Examples: Fe^{2+} and O_2^- ions form FeO, known as iron(II) oxide.
   Fe^{3+} and O_2^- ions form Fe_2O_3, known as iron(III) oxide.

5. When the compound contains a polyatomic ion, simply use the name of the polyatomic ion in place of the anion or cation.

   Examples: The name for NaOH is sodium hydroxide.
   The name for $(NH_4)_2S$ is ammonium sulfide.

PRACTICE Problems ADDITIONAL PRACTICE

Interpret the formula representations and name these compounds.

28. NaBr
29. $CaCl_2$
30. KOH
31. $Cu(NO_3)_2$
32. Ag_2CrO_4
33. **CHALLENGE** The ionic compound NH_4ClO_4 is a key reactant used in solid rocket boosters, such as those that powered the Space Shuttle into orbit. Name this compound.

28. sodium bromide
29. calcium chloride
30. potassium hydroxide
31. copper(II) nitrate
32. silver chromate
33. ammonium perchlorate

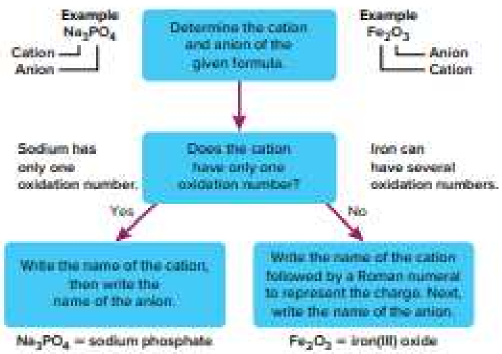
PROBLEM-SOLVING STRATEGY

Naming Ionic Compounds

Naming ionic compounds is easy if you follow this naming-convention flowchart.

Apply the Strategy

Name the compounds KOH and Ag_2CrO_4 using this flowchart.



11. Explain some physical properties of metals (melting and boiling points, thermal and electrical conductivity, malleability, ductility, durability, hardness, and strength). (Pages 83 and 84)

Physical Properties of Metals	
STATE	All metals are solid at room temperature except mercury and gallium. Mercury is liquid at room temperature, whereas gallium is a soft solid at a temperature below 29.76°C and liquid at a temperature above it.
LUSTER	All metals are lustrous in nature i.e. they show brightness on reflection of light.
HARDNESS	It is the property that exhibits resistance to indentation by an external force. Almost all metals except zinc, mercury, and gallium are hard.
MALLEABILITY	It is the property of metals to be beaten into thin sheets on hammering. Except for zinc, all metals are malleable.
DUCTILITY	It is the property of metals to be drawn into wires. Except for zinc and mercury, all metals are ductile.
TENSILE STRENGTH	It is the property of metals to withstand longitudinal pull when they are in the form of a wire. Almost all metals except zinc & mercury exhibit very high tensile strength.
SONORITY	Metals produces resonant sound when struck. This property of a metal is called sonority. Almost all metals are sonorous, however, the level of sonority varies from metal to metal. For example, when we drop a small iron nail, it produces a ringing sound.
CONDUCTIVITY	It is the property of metals to transmit heat or electricity through a unit area in unit time. Almost all metals are good conductors of heat and electricity, however, the level of conductivity varies from metal to metal.
MELTING & BOILING POINT	Usually metals have very high melting and boiling point, however exception to this is sodium, potassium, and mercury.
DENSITY	Almost all metals have very high density with an exception of sodium, potassium, and calcium.

12. Differentiate between sigma and pi bonds. (Pages 96 to 98)

Sigma (σ) Bond	Pi (π) Bond
(a) It is formed by the end to end overlap of orbitals.	It is formed by the lateral overlap of orbitals.
(b) The orbitals involved in the overlapping are *s-s*, *s-p*, or *p-p*.	These bonds are formed by the overlap of *p-p* orbitals only.
(c) It is a strong bond.	It is weak bond.
(d) The electron cloud is symmetrical about the line joining the two nuclei.	The electron cloud is not symmetrical.
(e) It consists of one electron cloud, which is symmetrical about the internuclear axis.	There are two electron clouds lying above and below the plane of the atomic nuclei.
(f) Free rotation about σ bonds is possible.	Rotation is restricted in case of pi-bonds.

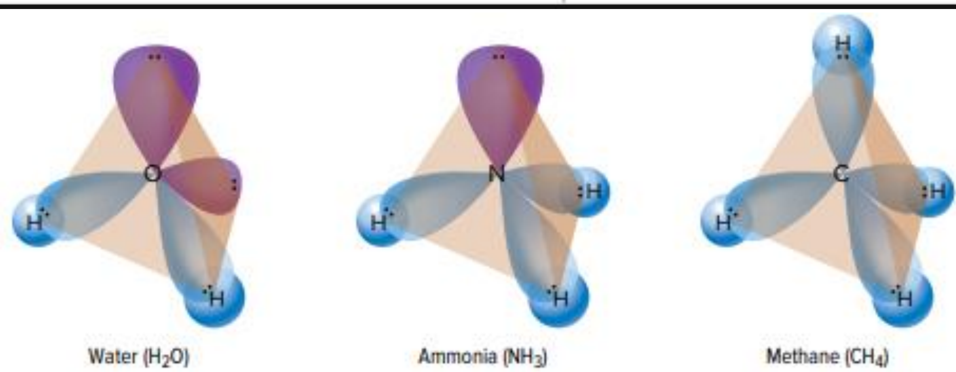


Figure 7 Sigma bonds formed in each of these molecules when the atomic orbital of each hydrogen atom overlapped end-to-end with the orbital of the central atom.

List the orbitals that can form sigma bonds in a covalent compound.

Sigma bonds can form from the overlap of an s orbital with another s orbital, an s orbital with a p orbital, or a p orbital end-to-end with another p orbital.

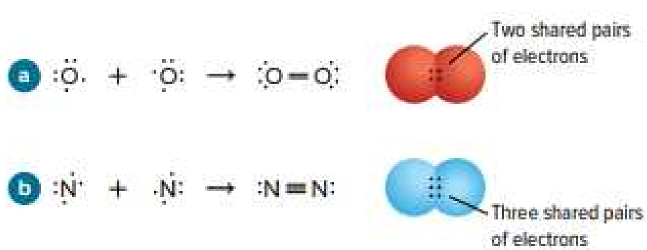


Figure 8 Multiple covalent bonds form when two atoms share more than one pair of electrons.

a. Two oxygen atoms form a double bond.

b. A triple bond forms between two nitrogen atoms.

Double bonds

A double covalent bond forms when two pairs of electrons are shared between two atoms. For example, atoms of the element oxygen only exist as diatomic molecules. Each oxygen atom has six valence electrons and must obtain two additional electrons for a noble-gas configuration, as shown in **Figure 8a**. A double covalent bond forms when each oxygen atom shares two electrons; a total of two pairs of electrons are shared between the two atoms.

Triple bonds

A triple covalent bond forms when three pairs of electrons are shared between two atoms. Diatomic nitrogen (N_2) molecules contain a triple covalent bond. Each nitrogen atom shares three electron pairs, forming a triple bond with the other nitrogen atom as shown in **Figure 8b**.

The pi bond

A multiple covalent bond consists of one sigma bond and at least one pi bond. A **pi bond,** represented by the Greek letter pi (π), forms when parallel orbitals overlap and share electrons. The shared electron pair of a pi bond occupies the space above and below the line that represents where the two atoms are joined together.

It is important to note that molecules having multiple covalent bonds contain both sigma and pi bonds. A double covalent bond, as shown in **Figure 9**, consists of one pi bond and one sigma bond. A triple covalent bond consists of two pi bonds and one sigma bond.

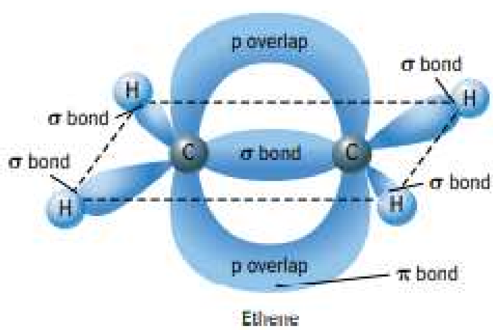
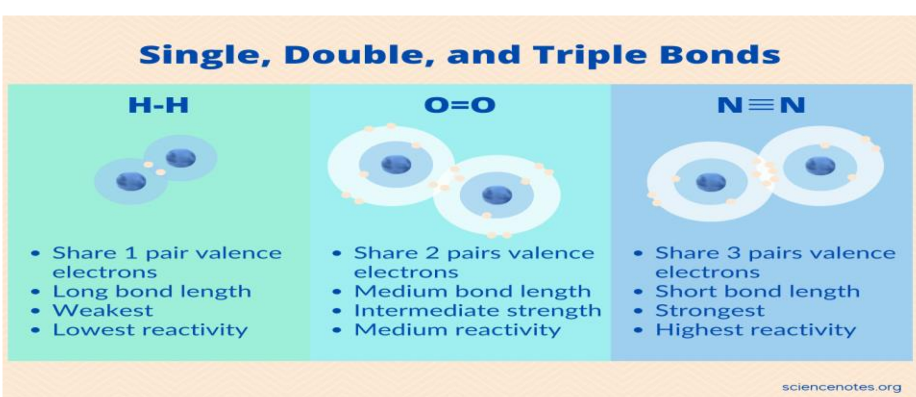


Figure 9 Notice how the multiple bond between the two carbon atoms in ethene (C_2H_4) consists of a sigma bond and a pi bond. The sigma bond is formed by the end-to-end overlap of orbitals directly between the two carbon atoms. The carbon atoms are close enough that the side-by-side p orbitals overlap and form the pi bond. This results in a doughnut-shaped cloud around the sigma bond.

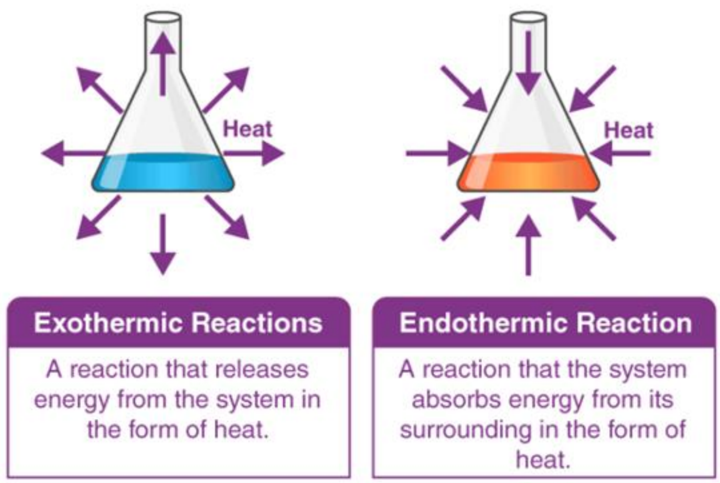
13. Identify the relationship between the type of a covalent bond (single, double, triple) and its bond length, bond strength and the bond dissociation energy. (Page 99,100)



- **The triple bond is shorter than a double bond, and a double bond is shorter than a single bond.**
- **Weaker bond strength results in the bond having less bond dissociation energy.**
- **On the other hand, stronger bond strength results in the bond having stronger bond dissociation energy.**

Table 1 Covalent Bond Type, Bond Length, and Bond-Dissociation Energy

Molecule	Bond Type	Bond Length	Bond-Dissociation Energy
F_2	single covalent	1.43×10^{-10} m	159 kJ/mol
O_2	double covalent	1.21×10^{-10} m	498 kJ/mol
N_2	triple covalent	1.10×10^{-10} m	945 kJ/mol



14. Name a binary molecular compound based on its molecular formula. (Page 102)

1. The first element in the formula is always named first, using the entire element name. N is the symbol for *nitrogen*.
2. The second element in the formula is named using its root and adding the suffix *-ide*. O is the symbol for oxygen so the second word is *oxide*.
3. Prefixes are used to indicate the number of atoms of each element that are present in the compound. **Table 2** lists the most common prefixes used. There are two atoms of nitrogen and one atom of oxygen, so the first word is *dinitrogen* and the second word is *monoxide*.

Table 2 Prefixes in Covalent Compounds

Number of Atoms	Prefix	Number of Atoms	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

PRACTICE Problems ADDITIONAL PRACTICE

Name each of the binary covalent compounds listed below.

14. CO_2

15. SO_2

16. NF_3

17. CCl_4

18. CHALLENGE What is the formula for diarsenic trioxide?

14. carbon dioxide
15. sulfur dioxide
16. nitrogen trifluoride
17. carbon tetrachloride
18. As_2O_3

15. Name an acid (binary acid and oxyacid) given its chemical formula and vice versa. (Page 104)

Naming binary acids

A binary acid contains hydrogen and one other element. The naming of the common binary acid known as hydrochloric acid is explained in the following rules.

1. The first word has the prefix *hydro-* to name the hydrogen part of the compound. The rest of the first word consists of a form of the root of the second element plus the suffix *-ic*. HCl (hydrogen and chlorine) becomes *hydrochloric*.
2. The second word is always *acid*. Thus, HCl in a water solution is called *hydrochloric acid*.

Table 3 Naming Oxyacids

Compound	Oxyanion	Acid Suffix	Acid Name
$HClO_3$	chlorate	-ic	chloric acid
$HClO_2$	chlorite	-ous	chlorous acid
HNO_3	nitrate	-ic	nitric acid
HNO_2	nitrite	-ous	nitrous acid

1. First, identify the oxyanion present. The first word of an oxyacid's name consists of the root of the oxyanion and the prefix *per-* or *hypo-* if it is part of the oxyanion's name. The first word of the oxyacid's name also has a suffix that depends on the oxyanion's suffix. If the oxyanion's name ends with the suffix *-ate*, replace it with the suffix *-ic*. If the name of the oxyanion ends with the suffix *-ite*, replace it with the suffix *-ous*. NO_3^-, the nitrate ion, becomes *nitric*.
2. The second word of the name is always *acid*. HNO_3 (hydrogen and the nitrate ion) becomes *nitric acid*.

PRACTICE Problems ADDITIONAL PRACTICE

Name the following acids. Assume each compound is dissolved in water.

19. HI

20. $HClO_3$

21. $HClO_2$

22. H_2SO_4

23. H_2S

24. CHALLENGE What is the formula for periodic acid?

19. hydroiodic acid

20. chloric acid

21. chlorous acid

22. sulfuric acid

23. hydrosulfuric acid

24. HIO_4

Table 4 Formulas and Names of Some Covalent Compounds

Formula	Common Name	Molecular Compound Name
H_2O	water	dihydrogen monoxide
NH_3	ammonia	nitrogen trihydride
N_2H_4	hydrazine	dinitrogen tetrahydride
HCl	muriatic acid	hydrochloric acid
$C_9H_8O_4$	aspirin	2-(acetyloxy)benzoic acid

16. Draw Lewis structure for a number of covalent compounds with single and multiple bonds. (Page 109)

EXAMPLE Problem 3

LEWIS STRUCTURE FOR A COVALENT COMPOUND WITH SINGLE BONDS Ammonia is a raw material used in the manufacture of many products, including fertilizers, cleaning products, and explosives. Draw the Lewis structure for ammonia (NH_3).

1 ANALYZE THE PROBLEM

Ammonia molecules consist of one nitrogen atom and three hydrogen atoms. Because hydrogen must be a terminal atom, nitrogen is the central atom.

2 SOLVE FOR THE UNKNOWN

Find the total number of valence electrons available for bonding.

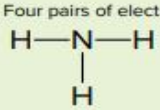
$$1 \text{ N atom} \times \frac{5 \text{ valence electrons}}{1 \text{ N atom}} + 3 \text{ H atoms} \times \frac{1 \text{ valence electron}}{1 \text{ H atom}} = 8 \text{ valence electrons}$$

There are 8 valence electrons available for bonding.

$$\frac{8 \text{ electrons}}{2 \text{ electrons/pair}} = 4 \text{ pairs}$$

Determine the total number of bonding pairs. To do this, divide the number of available electrons by two.

Four pairs of electrons are available for bonding.



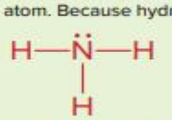
Place a bonding pair (a single bond) between the central nitrogen atom and each terminal hydrogen atom.

Determine the number of bonding pairs remaining.

4 pairs total − 3 pairs used = 1 pair available

Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.

The remaining pair—a lone pair—must be added to either the terminal atoms or the central atom. Because hydrogen atoms can have only one bond, they have no lone pairs.



Place the remaining lone pair on the central nitrogen atom.

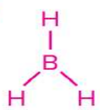
3 EVALUATE THE ANSWER

Each hydrogen atom shares one pair of electrons, as required, and the central nitrogen atom shares three pairs of electrons and has one lone pair, providing a stable octet.

PRACTICE Problems ADDITIONAL PRACTICE

37. Draw the Lewis structure for BH_3.

38. CHALLENGE A nitrogen trifluoride molecule contains numerous lone pairs. Draw its Lewis structure.

37.


38.


17. Determine the exceptions to the octet rule (add number of valence electrons, sub octets and coordinate covalent bonds, expanded octets). (Page 113,114 and 115)

Exceptions to the Octet Rule

Generally, atoms attain an octet when they bond with other atoms. Some molecules and ions, however, do not obey the octet rule. There are several reasons for these exceptions.

Odd number of valence electrons

First, a small group of molecules might have an odd number of valence electrons and be unable to form an octet around each atom. For example, NO_2 has five valence electrons from nitrogen and 12 from oxygen, totaling 17, which cannot form an exact number of electron pairs.

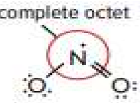
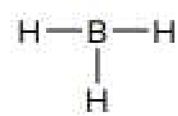


Figure 15 The central nitrogen atom in this NO_2 molecule does not satisfy the octet rule; the nitrogen atom has only seven electrons in its outer energy level.

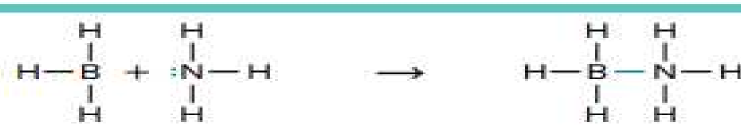
ClO_2 and NO are other examples of molecules with odd numbers of valence electrons.

Suboctets and coordinate covalent bonds

Another exception to the octet rule is due to a few compounds that form suboctets—stable configurations with fewer than eight electrons present around an atom. This group is relatively rare, and BH_3 is an example. Boron, a group 13 metalloid, forms three covalent bonds with other nonmetallic atoms. The boron atom shares only six electrons—too few to form an octet. Such compounds tend to be reactive and can share an entire pair of electrons donated by another atom.



A **coordinate covalent bond** forms when one atom donates both of the electrons to be shared with an atom or ion that needs two electrons to form a stable electron arrangement with lower potential energy. Refer to **Figure 16**. Atoms or ions with lone pairs often form coordinate covalent bonds with atoms or ions that need two more electrons.



The boron atom has no electrons to share, whereas the nitrogen atom has two electrons to share.

The nitrogen atom shares both electrons to form the coordinate covalent bond.

Figure 16 In this reaction between boron trihydride (BH_3) and ammonia (NH_3), the nitrogen atom donates both electrons that are shared by boron and nitrogen, forming a coordinate covalent bond.

Expanded octets

The third group of compounds that does not follow the octet rule has central atoms that contain more than eight valence electrons. This electron arrangement is referred to as an expanded octet. An expanded octet can be explained by considering the d orbitals that occur in the energy levels of elements in period three or higher. An example of an expanded octet, is the bond formation in the molecule PCl_5.

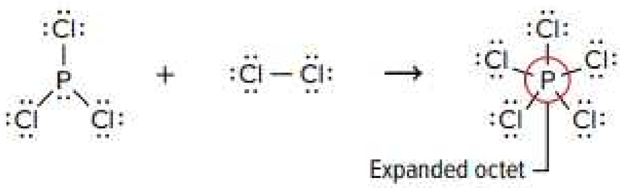


Figure 17 Prior to the reaction of PCl_3 and Cl_2, every reactant atom follows the octet rule. After the reaction, the product, PCl_5, has an expanded octet containing ten electrons.

EXAMPLE Problem 6

LEWIS STRUCTURE: EXCEPTION TO THE OCTET RULE Xenon is a noble gas that will form a few compounds with nonmetals that strongly attract electrons. Draw the correct Lewis structure for xenon tetrafluoride (XeF_4).

1 ANALYZE THE PROBLEM

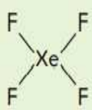
You are given that a molecule of xenon tetrafluoride consists of one xenon atom and four fluorine atoms. Xenon has less attraction for electrons, so it is the central atom.

2 SOLVE FOR THE UNKNOWN

First, find the total number of valence electrons.

$$1 \text{ Xe atom} \times \frac{8 \text{ valence electrons}}{1 \text{ Xe atom}} + 4 \text{ F atoms} \times \frac{7 \text{ valence electrons}}{1 \text{ F atom}} = 36 \text{ valence electrons}$$

$$\frac{36 \text{ electrons}}{2 \text{ electrons/pair}} = 18 \text{ pairs}$$ Determine the total number of bonding pairs.

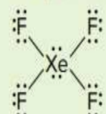


Use four bonding pairs to bond the four F atoms to the central Xe atom.

EXAMPLE Problem 6 (continued)

18 pairs available − 4 pairs used = 14 pairs available — Determine the number of remaining pairs.

$14 \text{ pairs} - 4 \text{ F atoms} \times \frac{3 \text{ pairs}}{1 \text{ F atom}} = 2 \text{ pairs unused}$ — Add three pairs to each F atom to obtain an octet. Determine how many pairs remain.



Place the two remaining pairs on the central Xe atom.

3 EVALUATE THE ANSWER

This structure gives xenon 12 total electrons, an expanded octet. Xenon compounds, such as the XeF_4 shown here, are toxic because they are highly reactive.

PRACTICE Problems ADDITIONAL PRACTICE

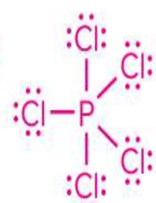
Draw the expanded octet Lewis structure for each molecule.

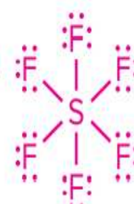
47. ClF_3

48. PCl_5

49. CHALLENGE Draw the Lewis structure for the molecule formed when six fluorine atoms and one sulfur atom bond covalently.

47.


48.


49.


18. Using the VESPR model the electron domain geometry and molecular geometry for different molecule and ions. (Page 119)

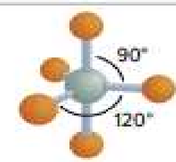
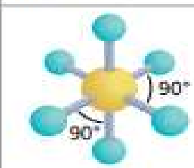
Table 5 **Molecular Shapes: 2 or 3 Total Pairs**

Molecule	Total Pairs	Shared Pairs	Lone Pairs	Hybrid Orbitals	Molecular Shape*
$BeCl_2$	2	2	0	sp	 180° Linear
$AlCl_3$	3	3	0	sp^2	 120° Trigonal planar

The $BeCl_2$ molecule contains only two pairs of electrons shared with the central Be atom. These bonding electrons have the maximum separation, a bond angle of 180°, and the molecular shape is linear.

The three bonding electron pairs in $AlCl_3$ have maximum separation in a trigonal planar shape with 120° bond angles.

Table 6 **Molecular Shapes: 4, 5, or 6 Total Pairs**

Molecule	Total Pairs	Shared Pairs	Lone Pairs	Hybrid Orbitals	Molecular Shape*
CH_4	4	4	0	sp^3	 Tetrahedral
NH_3	4	3	1	sp^3	 Trigonal pyramidal
H_2O	4	2	2	sp^3	 Bent
$NbBr_5$	5	5	0	sp^3d	 Trigonal bipyramidal
SF_6	6	6	0	sp^3d^2	 Octahedral

When the central atom in a molecule has four pairs of bonding electrons, as CH_4 does, the shape is tetrahedral. The bond angles are 109.5°.

NH_3 has three single covalent bonds and one lone pair. The lone pair takes up a greater amount of space than the shared pairs. There is stronger repulsion between the lone pair and the bonding pairs than between two bonding pairs. The resulting geometry is trigonal pyramidal, with 107.3° bond angles.

Water has two covalent bonds and two lone pairs. Repulsion between the lone pairs causes the angle to be 104.5°, less than both tetrahedral and trigonal pyramid. As a result, water molecules have a bent shape.

The $NbBr_5$ molecule has five pairs of bonding electrons. The trigonal bipyramidal shape minimizes the repulsion of these shared electron pairs.

As with $NbBr_5$, SF_6 has no unshared electron pairs on the central atom. However, six shared pairs arranged about the central atom result in an octahedral shape.

EXAMPLE Problem 7

FIND THE SHAPE OF A MOLECULE Phosphorus trihydride, a colorless gas, is produced when organic materials, such as fish flesh, rot. What is the shape of a phosphorus trihydride molecule? Predict the bond angle and identify hybrid orbitals.

1 ANALYZE THE PROBLEM

A phosphorus trihydride molecule has three hydrogen atoms bonded to a central phosphorus atom.

EXAMPLE Problem 7 (continued)

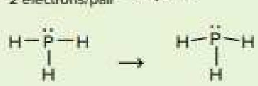
2 SOLVE FOR THE UNKNOWN

Find the total number of valence electrons and the number of electron pairs.

$$1 \text{ P atom} \times \frac{5 \text{ valence electrons}}{1 \text{ P atom}} + 3 \text{ H atoms} \times \frac{1 \text{ valence electron}}{1 \text{ H atom}} = 8 \text{ valence electrons}$$

$$\frac{8 \text{ electrons}}{2 \text{ electrons/pair}} = 4 \text{ pairs}$$

Determine the total number of bonding pairs.



Lewis structure → **Molecular shape**

Draw the Lewis structure, using one pair of electrons to bond each H atom to the central P atom and assigning the lone pair to the P atom.

The molecular shape is trigonal pyramidal with a predicted 107° bond angle and sp^3 hybrid orbitals.

3 EVALUATE THE ANSWER

All electron pairs are used and each atom has a stable electron configuration.

PRACTICE Problems ADDITIONAL PRACTICE

Determine the molecular shape, bond angle, and hybrid orbitals for each molecule.

56. BF_3 **57.** OCl_2 **58.** BeF_2 **59.** CF_4

60. CHALLENGE For a NH_4^+ ion, identify its molecular shape, bond angle, and hybrid orbitals.

56. trigonal planar, 120°, sp^2
57. bent, 104.5°, sp^3
58. linear, 180°, sp
59. tetrahedral, 109°, sp^3
60. tetrahedral, 109°, sp^3

19. Categorize bond types using electronegativity difference into mostly ionic, polar covalent, mostly covalent, and non-polar covalent bonds. (Pages 122 and 123)

Electronegativity Values for Selected Elements

Metal · Metalloid · Nonmetal

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
1	1 H 2.20																
2	3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98
3	11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16
4	19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96
5	37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 2.10	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66
6	55 Cs 0.79	56 Ba 0.89	57 La 1.10	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2
7	87 Fr 0.7	88 Ra 0.9	89 Ac 1.1														

Table 7 EN Difference Between Atoms within a Compound and Bond Character

Electronegativity Difference	Bond Character
> 1.7	mostly ionic
0.4 – 1.7	polar covalent
< 0.4	mostly covalent
0	nonpolar covalent

20. Describe conditions needed for a molecule to be polar. (Page 125)

- **A molecule will be polar if the bonds are polar, and these polar bonds are arranged in such a geometry that the dipole moments do not cancel one another.**
- **For a molecule to be polar, the difference in the electronegativity of the bonded atoms must be between 0.4 and 1.7. Then, if the molecule is symmetric, then the molecules are usually nonpolar.**
- **Molecules that are asymmetric are polar if the bond type is polar.**
- **Example: CCl_4. In CCl_4, the bond type is polar, but it is a symmetric molecule making it non-polar whereas NH_3 has polar bond and is asymmetric making it polar.**