Chapters 6

Chemical Bonding

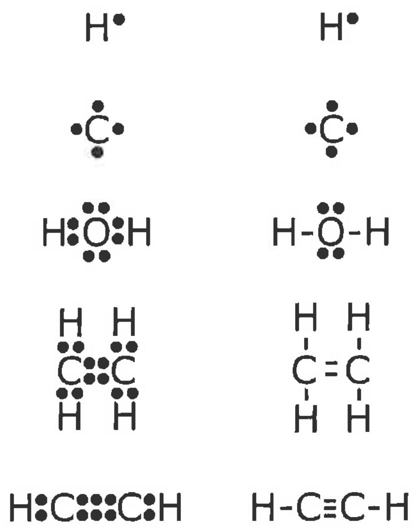
Hydrogen

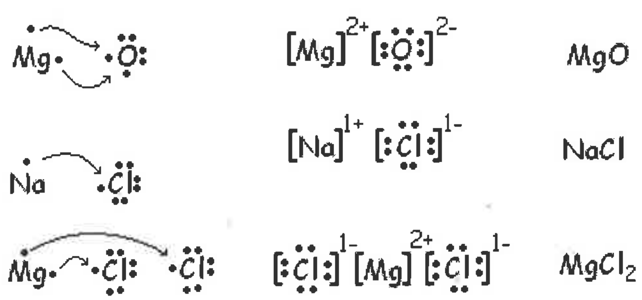
Carbon

Water

Ethylene

Acetylene





cation	anion	compound
Ca^{+2}	Cl^{-1}	$CaCl_2$
Ba^{+2}	O^{-2}	BaO
K^{+1}	S^{-2}	K_2S
Fe^{+3}	Br^{-1}	$FeBr_3$
Cr^{+3}	O^{-2}	Cr_2O_3

Chapter 7

Chemical Formulas

Numbers and Their Corresponding Prefixes:	
Prefix	**Means:**
mono- (or mon-)	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

CHAPTER SIX

CHEMICAL BONDING

Name ____ Date____________

Chapter 6: Chemical Bonding

ection one: Pages 175-177

A ______________________ ____________________ is a mutual electrical attraction between the nuclei and valence electrons of different atoms that bond the atoms together. Why do atoms bond? ______________________________

Types of bonding:

Ionic: Results from a **transfer** of electrons. The result is an electrical attraction between cations and anions.

Covalent: Result of **sharing** electrons between two atoms.

Types of bonding as determined by electronegativity difference: SEE YELLOW PACKET.

Nonpolar covalent bond : shared electrons are shared equally between two atoms. The result is a balanced distributior of electrical charge (E.D is 0.3 or less.)

Polar covalent bond: Shared electrons are NOT shared equally between two atoms. The result is an uneven distributior of electrical charge. (E.D is 0.4 to 1.7)

'onic bond: E.D of 1.8 and higher.

he atom with the higher electronegativity pulls the shared electrons closer to it. Consequently, this atom has a partial negative charge. To denote this charge a _______ is used. The symbol _________ denotes the partial positive charge.

Example: HCl

What type of bonding would you expect between the following atoms?

Li and F __________________________ Cu and S ______________ Cs and S _____________

H and S __________________________ H and O _______________ I and Br _____________

Section 2 Covalent Bonding and Molecular Compounds (pages 178-189)

A neutral group of atoms that are held together with covalent bonds are called a _____________________.

A chemical compound whose simplest units are molecules is called a _______________________ compound.

Write down the seven common diatomic elements.

____

Why does nature favor chemical bonding?

The distance between two bonded atoms at their minimum potential energy = ____________ ________________.

Bond energy is the energy required to break a chemical bond and form neutral isolated atoms. Reported in kJ/mol. Example: 436kJ is required to break the hydrogen-hydrogen bond in a one mole of hydrogen molecules to form two moles of separated hydrogen atoms.

Atoms share electrons by **overlapping their orbitals.**

Octet Rule: Chemical compounds are formed so that each atom can obtain an octet of electrons in its highest occupied energy level.

Example: F_2 molecule

Exceptions to the octet rule: Boron ([He]$2s^2$ $2p^1$) combining with 3 fluorine atoms.

Expanded valence involves the s, p, and d orbitals. This happens when an atoms combines with a highly electronegative element (F,O,Cl). The result is an atom with more than eight electrons around it. **Examples: PF_5 and SF_6**

Electron Dot Notation shows only the valence electrons of an atom of a particular element. Think of each side as representing the s and three p orbitals. Draw electron dot notation for : Na Cl O Al Ar

Lewis structures: When dot notation is used to represent the bonding between atoms in molecular compounds. Each pair of dots between two atoms can be changed into a line. Cl-Cl F-F H-H

Lone pairs: Two electrons in dot notation that are not shared between two atoms.

Rules:

1. Add up the valence electrons. This is how many you use. No more.....No less.

2. Determine the central atom. (If carbon is involved, it is always central. If carbon is not in formula, choose the least electronegative to be central.) Hydrogen is NEVER in the center.

3. Think symmetry and place the other atoms around the central atom.

4. Place eight electrons around the outer atoms. (Except Hydrogen....It only wants one more for a total of 2)

5. Add the remaining electrons to your central atom.

6. Count the electrons in the structure to be sure that the number of valence electrons used equals the number available. Be sure the central atom has eight and the other atoms (besides hydrogen) have eight.

7. If carbon is in the middle, it can double or triple bond to allow the attached atoms to obtain eight. Carbon can single, double, or triple bond with itself, nitrogen, and oxygen. Move one or more lone electron pairs to existing bonds between non-hydrogen atoms until the outer shells of all atoms are completely filled.

Practice: H_2O F_2 H_2 SO_3^{2-}

NH_3 H_2S CH_3I N_2

Resonance: Bonding in molecules or ions that cannot be correctly represented by a single Lewis structure.

Example: Ozone can be represented by or

A double headed arrow is placed between a molecule's resonance structures.

Covalent-Network bonding. (Example: Diamond)

This occurs between many identical atoms (or molecules)held together by forces acting between the atoms or molecules.

Section 3 Ionic Bonding and Ionic Compounds (pages 190-194)

An ___________compound is composed of positive and negative ions that are combined so that the numbers of positive and negative charges are equal. Most ionic compounds are crystalline solids.

A crystal is a three-dimensional network of positive and negative ions mutually attracted to one another. The chemical formula of an ionic compound represents the simplest ratio of the compound's combined ions. This is called the formula unit.

Cations take the name of the element. Anions end in –ide.

Representing cations and anions:

Sodium atom sodium ion chlorine atom chlorine ion

Since nature favors a system of lower potential energy, the orderly arrangement within a crystal results in a **crystal lattice structure.** The three dimensional shape of the lattice structure depends on the strengths of attraction and the size and charge of the cation and anion. When calcium and fluorine combine, the calcium ion is surrounded by eight fluoride anions and each fluoride ion is surrounded by four calcium cations.

To compare the bond strengths in ionic compounds, chemists compare the amounts of energy released when separated ions in a gas come together to make a crystal lattice structure. Negative values indicate that energy is released when crystals form.

The force that holds ions together in ionic compounds is a very strong overall attraction between ions. The force of attraction between molecules is much weaker . Therefore, ionic and molecular compounds have different properties.

	Molecular	Ionic
Melting Point		
Boiling Point		
Hardness		
State of matter		
Conductivity as solids		
Conductivity as liquids		
Conductivity (aq)		

Polyatomic ion: Charged group of covalently bonded atoms.

	Formula	Lewis structure
Ammonium		
Nitrate		
Sulfate		
Phosphate		

Metallic Bonding:

Metals have mobile valence electrons. **Metals are described as atoms immersed in a sea of mobile electrons.** The mobile valence electrons are the reason metals are lustrous.

Metals are malleable and ductile because one plane of atoms can slide past another without resistance or breaking bonds.

The strength of the metallic bond is determined by the nuclear charge of the metal atoms and the number of electrons in the metal's electron sea. Both of these factors are reflected in the metal's **enthalpy of vaporization**. This is the amount of energy absorbed when a specified amount changes to a gas.

Name ______________________________ Date_________

er 6 Section 5

VSEPR theory:

Lone pairs:

Shapes to know: See page 200 of text

	Formula example	Lewis structure	Hybrid
Linear			sp
Trigonal planar			sp^2
Bent/angular			sp^3
hedral			sp^3
Trigonal pyramidal			sp^3
Trigonal bipyramidal			
Octahedral			

Hybridization: When there is a mixing of similar energy orbitals on the same atom to produce new hybrid atomic orbitals.

Carbon:

):

Hybrid orbitals are orbitals of equal energy produced by the combination of two or more orbitals on the same atom.

How to figure out if a molecule is polar (a permanent dipole):

1. Draw the structural formula.
2. Find electronegativity of each atom.
3. Draw arrows to represent difference.
4. Review formula to see if one end of the molecule is more negative than another.

Examples:

H_2O NH_3 HF

F_2 CBr_4 NH_2Cl

CO_2 SCl_2 BF_3

Intramolecular forces: Forces between atoms.

Intermolecular forces: Forces between molecules.

Strongest act between polar molecules.

Hydrogen bonding (The strongest dipole-dipole force.) Exists between hydrogen and F,O,and N.

Therefore, it is a force that exists when a hydrogen atom that is bonded to a highly electronegative atom is attracted to an unshared pair of electrons of an electronegative atom in a nearby molecule.

London Dispersion Forces:

- Since electrons are in continuous motion, any change which causes the distribution to be unequal will create (induce) a temporary dipole. The temporary dipole induces another molecule to become a temporary dipole.
- The intermolecular attractions resulting from the constant motion of electrons and the creation of instantaneous dipoles are called London Dispersion Forces.
- These are the only forces acting on noble gases and nonpolar atoms. They become stronger with atoms with lots of electrons. (Increase with atomic mass).

Bond length:

Bond length varies depending on what atom you are dealing with.

There are two trends to remember:

1. Down a group: bond length increases
2. Multiple bonds are shorter than single bonds. The more electrons in a bond, the more it is attracted to the nucleus.

Why does ice float?

When the polar molecules slow down at low temperatures, they form hexagonal rings due to the intermolecular forces.

Water when frozen take sup more space than as a liquid.

The shape of small molecules helps to determine their polarity.

It is the polarity of large molecules that determines their shape.

Proteins are composed of individual subunits linked in chains. Some are polar and some are nonpolar. The polarities cause them to bend and twist.

Nonpolar subgroups tend to go inside while polar groups collect on the outside.

The bending and twisting results in a specific shape.

These shapes allow a protein to have a specialized function.

Determining Molecular Polarity

1. Draw correct structural formula
2. Calculate the bond polarities
   a. Look up the electronegativies of the elements
   b. Subtract the electronegativities of two atoms that are bonded together
   c. Repeat for all different bonds
   d. If the difference is 0.2 or less then the bond is NONPOLAR
   e. If the difference is more than 0.2 then the bond is POLAR
3. Add arrows and charges to your structural formula
4. Determine the shape of the molecule. Is the shape a polar shape or a nonpolar shape?
   a. If it has lone pairs on the center it should be POLAR
      i. With 2 exceptions: linear and square planar are both nonpolar shapes
   b. If there are no lone pairs on the center it is NONPOLAR
5. Decide if the molecule is polar or nonpolar
   a. All NONPOLAR bonds: NONPOLAR
   b. POLAR bond with POLAR shape: POLAR
   c. POLAR bonds with NONPOLAR shape: look at direction of arrows
      i. All bond arrows point the same way (out or in): NONPOLAR
      ii. Bond arrows point different directions (some in and some out): POLAR

Bonds	**Shape**	**Molecule**	
Nonpolar	All shapes	Nonpolar	
Polar	Polar	Polar	
Polar	Nonpolar	Look at directions of polarity. What charges are on the outside?	
		All same: Nonpolar	Different: Polar

1A						**ELECTRONEGATIVITIES**											8A
H 2.1	2A											3A	4A	5A	6A	7A	He
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	Lu 1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn
Fr 0.7	Ra 0.9	Lr 1.3	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cp	Uut	Uuq	Uup	Uuh		Uuo
		La 1.1	Ce 1.1	Pr 1.1	Nd 1.1	Pm 1.1	Sm 1.1	Eu 1.1	Gd 1.1	Tb 1.2	Dy 1.2	Ho 1.2	Er 1.2	Tm 1.2	Yb 1.2		
		Ac 1.1	Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am 1.3	Cm 1.3	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3		

Name ______________________ Date ______________ Class ______________

CHAPTER 6 REVIEW

Chemical Bonding

SECTION 1

SHORT ANSWER **Answer the following questions in the space provided.**

1. ______ A chemical bond between atoms results from the attraction between the valence electrons and ____ of different atoms.

(a) nuclei
(b) inner electrons
(c) isotopes
(d) Lewis structures

2. ______ A covalent bond consists of

(a) a shared electron.
(b) a shared electron pair.
(c) two different ions.
(d) an octet of electrons.

3. ______ If two covalently bonded atoms are identical, the bond is identified as

(a) nonpolar covalent.
(b) polar covalent.
(c) ionic.
(d) dipolar.

4. ______ A covalent bond in which there is an unequal attraction for the shared electrons is

(a) nonpolar.
(b) polar.
(c) ionic.
(d) dipolar.

5. ______ Atoms with a strong attraction for electrons they share with another atom exhibit

(a) zero electronegativity.
(b) low electronegativity.
(c) high electronegativity.
(d) Lewis electronegativity.

6. ______ Bonds that possess between 5% and 50% ionic character are considered to be

(a) ionic.
(b) pure covalent.
(c) polar covalent.
(d) nonpolar covalent.

7. ______ The greater the electronegativity difference between two atoms bonded together, the greater the bond's percentage of

(a) ionic character.
(b) nonpolar character.
(c) metallic character.
(d) electron sharing.

8. The electrons involved in the formation of a chemical bond are called

______________________________.

9. A chemical bond that results from the electrostatic attraction between positive and negative ions is called a(n) ______________________________.

SECTION 1 continued

10. If electrons involved in bonding spend most of the time closer to one atom rather than the other, the bond is ______________________________.

11. If a bond's character is more than 50% ionic, then the bond is called a(n) ______________________________.

12. A bond's character is more than 50% ionic if the electronegativity difference between the two atoms is greater than ______________________________.

13. Write the formula for an example of each of the following compounds:

______________________ **a.** nonpolar covalent compound

______________________ **b.** polar covalent compound

______________________ **c.** ionic compound

14. Describe how a covalent bond holds two atoms together.

____

____

____

15. What property of the two atoms in a covalent bond determines whether or not the bond will be polar?

____

____

16. How can electronegativity be used to distinguish between an ionic bond and a covalent bond?

____

____

____

____

17. Describe the electron distribution in a polar-covalent bond and its effect on the partial charges of the compound.

____

____

____

Name ______________________ Date ______________ Class ______________

CHAPTER 6 REVIEW

Chemical Bonding

SECTION 2

SHORT ANSWER **Answer the following questions in the space provided.**

1. Use the concept of potential energy to describe how a covalent bond forms between two atoms.

____

____

____

____

____

____

2. Name two elements that form compounds that can be exceptions to the octet rule.

____

____

3. Explain why resonance structures are used instead of Lewis structures to correctly model certain molecules.

____

____

____

____

4. Bond energy is related to bond length. Use the data in the tables below to arrange the bonds listed in order of increasing bond length, from shortest bond to longest.

a.

Bond	Bond energy (kJ/mol)
H—F	569
H—I	299
H—Cl	432
H—Br	366

____

b.

Bond	Bond energy (kJ/mol)
C—C	346
C≡C	835
C=C	612

____

____

5. Draw Lewis structures to represent each of the following formulas:

a. NH_3

b. H_2O

c. CH_4

d. C_2H_2

e. CH_2O

Name ______________________ Date ______________ Class ______________

CHAPTER 6 REVIEW

Chemical Bonding

SECTION 3

SHORT ANSWER **Answer the following questions in the space provided.**

1. ______ The notation for sodium chloride, NaCl, stands for one

(a) formula unit.
(b) molecule.
(c) crystal.
(d) atom.

2. ______ In a crystal of an ionic compound, each cation is surrounded by a number of

(a) molecules.
(b) positive ions.
(c) dipoles.
(d) negative ions.

3. ______ Compared with the neutral atoms involved in the formation of an ionic compound, the crystal lattice that results is

(a) higher in potential energy.
(b) lower in potential energy.
(c) equal in potential energy.
(d) unstable.

4. ______ The lattice energy of compound A is greater in magnitude than that of compound B. What can be concluded from this fact?

(a) Compound A is not an ionic compound.
(b) It will be more difficult to break the bonds in compound A than those in compound B.
(c) Compound B has larger crystals than compound A.
(d) Compound A has larger crystals than compound B.

5. ______ The forces of attraction between molecules in a molecular compound are generally

(a) stronger than the attractive forces among formula units in ionic bonding.
(b) weaker than the attractive forces among formula units in ionic bonding.
(c) approximately equal to the attractive forces among formula units in ionic bonding.
(d) equal to zero.

6. Describe the force that holds two ions together in an ionic bond.

____

____

____

7. What type of energy best represents the strength of an ionic bond?

____

____

8. What types of bonds are present in an ionic compound that contains a polyatomic ion?

____

____

____

____

9. Arrange the ionic bonds in the table below in order of increasing strength from weakest to strongest.

Ionic bond	Lattice energy (kJ/mol)
NaCl	−787
CaO	−3384
KCl	−715
MgO	−3760
LiCl	−861

____

____

10. Draw Lewis structures for the following polyatomic ions:

a. NH_4^+

b. SO_4^{2-}

11. Draw the two resonance structures for the nitrite anion, NO_2^-.

Name ______________________ Date ______________ Class ______________

CHAPTER 6 REVIEW

Chemical Bonding

SECTION 4

SHORT ANSWER **Answer the following questions in the space provided.**

1. ______ In metals, the valence electrons are considered to be

(a) attached to particular positive ions.
(b) shared by all surrounding atoms.
(c) immobile.
(d) involved in covalent bonds.

2. ______ The fact that metals are malleable and ionic crystals are brittle is best explained in terms of their

(a) chemical bonds.
(b) London forces.
(c) enthalpies of vaporization.
(d) polarity.

3. ______ As light strikes the surface of a metal, the electrons in the electron sea

(a) allow the light to pass through.
(b) become attached to particular positive ions.
(c) fall to lower energy levels.
(d) absorb and re-emit the light.

4. ______ Mobile electrons in the metallic bond are responsible for

(a) luster.
(b) thermal conductivity.
(c) electrical conductivity.
(d) All of the above.

5. ______ In general, the strength of the metallic bond ____ moving from left to right on any row of the periodic table.

(a) increases
(b) decreases
(c) remains the same
(d) varies

6. ______ When a metal is drawn into a wire, the metallic bonds

(a) break easily.
(b) break with difficulty.
(c) do not break.
(d) become ionic bonds.

7. Use the concept of electron configurations to explain why the number of valence electrons in metals tends to be less than the number in most nonmetals.

____

____

____

____

8. How does the behavior of electrons in metals contribute to the metal's ability to conduct electricity and heat?

____

____

____

9. What is the relationship between the enthalpy of vaporization of a metal and the strength of the bonds that hold the metal together?

____

____

____

10. Draw two diagrams of a metallic bond. In the first diagram, draw a weak metallic bond; in the second, show a metallic bond that would be stronger. Be sure to include nuclear charge and number of electrons in your illustrations.

a. **weak bond**

b. **strong bond**

11. Complete the following table:

	Metals	Ionic Compounds
Components		
Overall charge		
Conductive in the solid state		
Melting point		
Hardness		
Malleable		
Ductile		

Name ____________________ Date ______________ Class ______________

CHAPTER 6 REVIEW

Chemical Bonding

SECTION 5

SHORT ANSWER **Answer the following questions in the space provided.**

1. Identify the major assumption of the VSEPR theory, which is used to predict the shape of atoms.

____

____

2. In water, two hydrogen atoms are bonded to one oxygen atom. Why isn't water a linear molecule?

____

____

3. What orbitals combine together to form sp^3 hybrid orbitals around a carbon atom?

____

____

4. What two factors determine whether or not a molecule is polar?

____

____

5. Arrange the following types of attractions in order of increasing strength, with 1 being the weakest and 4 the strongest.

______ hydrogen bonding

______ ionic

______ dipole-dipole

______ London dispersion

6. How are dipole-dipole attractions, London dispersion forces, and hydrogen bonding similar?

____

____

____

____

Name ______________________ Date ______________ Class ______________

7. Complete the following table:

Formula	Lewis structure	Geometry	Polar
H_2S			
CCl_4			
BF_3			
H_2O			
PCl_5			
BeF_2			
SF_6			

Name ______________________ Date ______________ Class ______________

CHAPTER 6 REVIEW

Chemical Bonding

MIXED REVIEW

SHORT ANSWER **Answer the following questions in the space provided.**

1. Name the type of energy that is a measure of strength for each of the following types of bonds:

______________________ **a.** ionic bond

______________________ **b.** covalent bond

______________________ **c.** metallic bond

2. Use the electronegativity values shown in **Figure 20,** on page 161 of the text, to determine whether each of the following bonds is nonpolar covalent, polar covalent, or ionic.

______________ **a.** H—F ______________ **d.** H—H

______________ **b.** Na—Cl ______________ **e.** H—C

______________ **c.** H—O ______________ **f.** H—N

3. How is a hydrogen bond different from an ionic or covalent bond?

____

____

____

____

____

____

____

4. H_2S and H_2O have similar structures and their central atoms belong to the same group. Yet H_2S is a gas at room temperature and H_2O is a liquid. Use bonding principles to explain why this is.

____

____

____

____

Name ________________ Date ________ Class ________

MIXED REVIEW continued

5. In what way is a polar-covalent bond similar to an ionic bond?

____

____

____

____

6. Draw a Lewis structure for each of the following formulas. Determine whether the molecule is polar or nonpolar.

________ **a.** H_2S

________ **b.** $COCl_2$

________ **c.** PCl_3

________ **d.** CH_2O

Name ______________________________

Molecular Shapes and Angles (STUDY THESE FOR TEST)

Name of shape	Bond angle	Structural formula	Name and formula

Structural formula hints:

1. Count the valence electrons FIRST. Write down this number. You **<u>must</u>** use **<u>only</u>** this amount of electrons in your structural formula.
2. Figure out which atom goes in the middle.
   a. If you have one of something, chances are it goes in the middle.
   b. If you have carbon, put it in the middle.
   c. Put the element of **lower electronegativity** in the middle.
   d. **<u>NEVER</u>** put hydrogen in the middle.
3. The rest of the atoms must join to the middle, so add them around in a symmetrical pattern using single bonds. (Think adding petals onto the center of a flower.)
4. Now count up all the electrons you have used.
5. How many more electrons are left? Count them up.
6. Make sure the outer (petals) atoms have the amount necessary to achieve the octet (if possible).
   a. **Hydrogen ONLY NEEDS TWO. NEVER give it more than two. NO LONE PAIRS around H.**
   b. Give halogens eight valence electrons.
   c. Give oxygen eight.
   d. In fact, everyone else wants eight if they are outer atoms.
   e. Count how many valence electrons you used up to this point. You can not use any more than the original number from step one.
   f. If you have additional electrons to use up, put them on the center atom as lone pairs. Don't worry if the center atom has more than eight valence electrons. There are exceptions to the octet rule.
7. Check to see if the center atom has less than eight valence electrons.
   a. **EXCEPTION:** Boron can **never** have eight. It only has three places to share. It can have only six.
   b. If the center atom has less than eight, check to see if it is possible that an outer atom can have its electrons moved around and can form double or triple bonds with the center one.
   c. Oxygen will single or double bond. Carbon can single, double, or triple bond.
   d. DON'T double bond the outer halogens. If you have a halogen in the center, it might double bond.

SHAPE and Polarity (Polar molecules are also called Dipoles)

- Lone pairs on the center atom **take up more space**. They are held only by one atom.
- All electrons in bonds and as lone pairs **REPEL** each other. This may distort the shape.
- Remember: **Valence Electrons REPEL.**
- If the bonds are polar, the molecule is NOT NECESSARILY polar. You **MUST** consider shape.
- The **polarity of a bond** is determined by the **difference in electronegativity.**
- The **polarity of a molecule** is determined by the **bond polarity** AND the **shape of the molecule.**
- A polar molecule has one end negative and the other end positive due to bond polarity AND shape.
- The polarity of a large molecule will influence its shape and function.
- The shape of a small molecule will influence its polarity.
- Polar molecules like other polar molecules.
- Non-polar molecules are usually pushed together due to the nature of polar molecules attempting to get near one another. **Like** (in polarity) **dissolves like** (in polarity). Example: Polar molecules dissolves polar molecules.

CHAPTER SEVEN

CHEMICAL FORMULAS

Name ______________________ Date ______________ Class ______________

CHAPTER 7 REVIEW

Chemical Formulas and Chemical Compounds

SECTION 1

SHORT ANSWER **Answer the following questions in the space provided.**

1. ______ In a Stock system name such as iron(III) sulfate, the Roman numeral tells us

(a) how many atoms of Fe are in one formula unit.
(b) how many sulfate ions can be attached to the iron atom.
(c) the charge on each Fe ion.
(d) the total positive charge of the formula unit.

2. ______ Changing a subscript in a correctly written chemical formula

(a) changes the number of moles represented by the formula.
(b) changes the charges on the other ions in the compound.
(c) changes the formula so that it no longer represents the compound it previously represented.
(d) has no effect on the formula.

3. The explosive TNT has the molecular formula $C_7H_5(NO_2)_3$.

______________ **a.** How many elements make up this compound?

______________ **b.** How many oxygen atoms are present in one molecule of $C_7H_5(NO_2)_3$?

______________ **c.** How many atoms in total are present in one molecule of $C_7H_5(NO_2)_3$?

______________ **d.** How many atoms are present in a sample of 2.0×10^{23} molecules of $C_7H_5(NO_2)_3$?

4. How many atoms are present in each of these formula units?

______________ **a.** $Ca(HCO_3)_2$

______________ **b.** $C_{12}H_{22}O_{11}$

______________ **c.** $Fe(ClO_2)_3$

______________ **d.** $Fe(ClO_3)_2$

5. ______________ **a.** What is the formula for the compound dinitrogen pentoxide?

______________ **b.** What is the Stock system name for the compound FeO?

______________ **c.** What is the formula for sulfurous acid?

______________ **d.** What is the name for the acid H_3PO_4?

SECTION 1 continued

6. Some binary compounds are ionic, others are covalent. The type of bond favored partially depends on the position of the elements in the periodic table. Label each of these claims as True or False; if False, specify the nature of the error.

a. Covalently bonded binary molecular compounds are typically composed of nonmetals.

____

____

b. Binary ionic compounds are composed of metals and nonmetals, typically from opposite sides of the periodic table.

____

____

7. Refer to **Table 2** on page 226 of the text and **Table 5** on page 230 of the text for examples of names and formulas for polyatomic ions and acids.

a. Derive a generalization for determining whether an acid name will end in the suffix *-ic* or *-ous*.

____

____

____

b. Derive a generalization for determining whether an acid name will begin with the prefix *hydro-* or not.

____

____

____

8. Fill in the blanks in the table below.

Compound name	Formula
Aluminum sulfide	
Cesium carbonate	
	$PbCl_2$
	$(NH_4)_3PO_4$
Hydroiodic acid	

Name ______________________ Date ______________ Class ______________

CHAPTER 7 REVIEW

Chemical Formulas and Chemical Compounds

SECTION 3

SHORT ANSWER **Answer the following questions in the space provided.**

1. Label each of the following statements as True or False:

______________ **a.** If the formula mass of one molecule is x amu, the molar mass is x g/mol.

______________ **b.** Samples of equal numbers of moles of two different chemicals must have equal masses as well.

______________ **c.** Samples of equal numbers of moles of two different molecular compounds must have equal numbers of molecules as well.

2. How many moles of each element are present in a 10.0 mol sample of $Ca(NO_3)_2$?

____

PROBLEMS **Write the answer on the line to the left. Show all your work in the space provided.**

3. Consider a sample of 10.0 g of the gaseous hydrocarbon C_3H_4 to answer the following questions.

______________ **a.** How many moles are present in this sample?

______________ **b.** How many molecules are present in the C_3H_4 sample?

______________ **c.** How many carbon atoms are present in this sample?

______________ **d.** What is the percentage composition of hydrogen in the sample?

4. One source of aluminum metal is alumina, Al_2O_3.

______________ **a.** Determine the percentage composition of Al in alumina.

______________ **b.** How many pounds of aluminum can be extracted from 2.0 tons of alumina?

5. Compound A has a molar mass of 20 g/mol, and compound B has a molar mass of 30 g/mol.

______________ **a.** What is the mass of 1.0 mol of compound A, in grams?

______________ **b.** How many moles are present in 5.0 g of compound B?

______________ **c.** How many moles of compound B are needed to have the same mass as 6.0 mol of compound A?

Name ______________________ Date ____________ Class ____________

CHAPTER 7 REVIEW

Chemical Formulas and Chemical Compounds

SECTION 4

SHORT ANSWER **Answer the following questions in the space provided.**

1. Write empirical formulas to match the following molecular formulas:

______________ **a.** $C_2H_6O_4$

______________ **b.** N_2O_5

______________ **c.** Hg_2Cl_2

______________ **d.** C_6H_{12}

2. ______________ A certain hydrocarbon has an empirical formula of CH_2 and a molar mass of 56.12 g/mol. What is its molecular formula?

3. A certain ionic compound is found to contain 0.012 mol of sodium, 0.012 mol of sulfur, and 0.018 mol of oxygen.

______________ **a.** What is its empirical formula?

______________ **b.** Is this compound a sulfate, sulfite, or neither?

PROBLEMS **Write the answer on the line to the left. Show all your work in the space provided.**

4. Water of hydration was discussed in **Sample Problem K** on pages 243–244 of the text. Strong heating will drive off the water as a vapor in hydrated copper(II) sulfate. Use the data table below to answer the following:

Mass of the empty crucible	4.00 g
Mass of the crucible plus hydrate sample	4.50 g
Mass of the system after heating	4.32 g
Mass of the system after a second heating	4.32 g

______________ **a.** Determine the mass percentage of water in the original sample.

SECTION 4 continued

_______________ **b.** The compound has the formula $CuSO_4 \cdot xH_2O$. Determine the value of x.

c. What might be the purpose of the second heating?

____

____

____

____

5. Gas X is found to be 24.0% carbon and 76.0% fluorine by mass.

_______________ **a.** Determine the empirical formula of gas X.

_______________ **b.** Given that the molar mass of gas X is 200.04 g/mol, determine its molecular formula.

6. A compound is found to contain 43.2% copper, 24.1% chlorine, and 32.7% oxygen by mass.

_______________ **a.** Determine its empirical formula.

b. What is the correct Stock system name of the compound in part **a**?

____

Name ______________________ Date ______________ Class ______________

CHAPTER 7 REVIEW

Chemical Formulas and Chemical Compounds

MIXED REVIEW

SHORT ANSWER **Answer the following questions in the space provided.**

1. Write formulas for the following compounds:

______________ **a.** copper(II) carbonate

______________ **b.** sodium sulfite

______________ **c.** ammonium phosphate

______________ **d.** tin(IV) sulfide

______________ **e.** nitrous acid

2. Write the Stock system names for the following compounds:

______________ **a.** $Mg(ClO_4)_2$

______________ **b.** $Fe(NO_3)_2$

______________ **c.** $Fe(NO_2)_3$

______________ **d.** CoO

______________ **e.** dinitrogen pentoxide

3. ______________ **a.** How many atoms are represented by the formula $Ca(HSO_4)_2$?

______________ **b.** How many moles of oxygen atoms are in a 0.50 mol sample of this compound?

SKIP **c.** Assign the oxidation number to sulfur in the HSO_4^- anion.

4. Assign the oxidation number to the element specified in each of the following:

______ **a.** hydrogen in H_2O_2

______ **b.** hydrogen in MgH_2

______ **c.** sulfur in S_8

______ **d.** carbon in $(CO_3)^{2-}$

______ **e.** chromium in $Na_2Cr_2O_7$

______ **f.** nitrogen in NO_2

SKIP

Name ______________________ Date ____________ Class ____________

MIXED REVIEW continued

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

5. ______________ Following are samples of four different compounds. Arrange them in order of increasing mass, from smallest to largest.

a. 25 g of oxygen gas

b. 1.00 mol of H_2O

c. 3×10^{23} molecules of C_2H_6

d. 2×10^{23} molecules of $C_2H_6O_2$

6. ______________ **a.** What is the formula for sodium hydroxide?

______________ **b.** What is the formula mass of sodium hydroxide?

______________ **c.** What is the mass in grams of 0.25 mol of sodium hydroxide?

7. ______________ What is the percentage composition of ethane gas, C_2H_6, to the nearest whole number?

8. ______________ Ribose is an important sugar (part of RNA),with a molar mass of 150.15 g/mol. If its empirical formula is CH_2O, what is its molecular formula?

MIXED REVIEW continued

9. Butane gas, C_4H_{10}, is often used as a fuel.

______________ **a.** What is the mass in grams of 3.00 mol of butane?

______________ **b.** How many molecules are present in that 3.00 mol sample?

______________ **c.** What is the empirical formula of the gas?

10. ______________ Naphthalene is a soft covalent solid that is often used in mothballs. Its molar mass is 128.18 g/mol and it contains 93.75% carbon and 6.25% hydrogen. Determine the molecular formula of napthalene from this information.

11. Nicotine has the formula $C_xH_yN_z$. To determine its composition, a sample is burned in excess oxygen, producing the following results:

1.0 mol of CO_2
0.70 mol of H_2O
0.20 mol of NO_2

Assume that all the atoms in nicotine are present as products.

______________ **a.** Determine the number of moles of carbon present in the products of this combustion.

MIXED REVIEW continued

______________________ **b.** Determine the number of moles of hydrogen present in the combustion products.

______________________ **c.** Determine the number of moles of nitrogen present in the combustion products.

______________________ **d.** Determine the empirical formula of nicotine based on your calculations.

______________________ **e.** In a separate experiment, the molar mass of nicotine is found to be somewhere between 150 and 180 g/mol. Calculate the molar mass of nicotine to the nearest gram.

12. When $MgCO_3(s)$ is strongly heated, it produces solid MgO as gaseous CO_2 is driven off.

______________________ **a.** What is the percentage loss in mass as this reaction occurs?

SKIP ______________________ **b.** Assign the oxidation number to each atom in $MgCO_3$.

SKIP ______________________ **c.** Does the oxidation number of carbon change upon the formation of CO_2?

Chapter 7

Chemical Formulas and Chemical Compounds

Chem 1A

Polyatomic Ions to Memorize:

NH_4^{+1}	Ammonium
OH^{-1}	Hydroxide
ClO^{-1}	Hypochlorite
ClO_2^{-1}	Chlorite
ClO_3^{-1}	Chlorate
ClO_4^{-1}	Perchlorate
NO_2^{-1}	Nitrite
NO_3^{-1}	Nitrate
SO_3^{-2}	Sulfite
SO_4^{-2}	Sulfate
HSO_4^{-1}	Hydrogen sulfate (bisulfate)
PO_3^{-3}	Phosphite
PO_4^{-3}	Phosphate
HCO_3^{-1}	Hydrogen carbonate (bicarbonate)
CO_3^{-2}	Carbonate
$C_2H_3O_2^{-1}$	Acetate
CrO_4^{-2}	Chromate
$Cr_2O_7^{-2}$	Dichromate
O_2^{-2}	Peroxide
H^{-1}	Hydride
CN^{-1}	Cyanide
SCN^{-1}	Thiocyanide
IO_3^{-1}	Iodate
MnO_4^{-1}	Permanganate

PERIODIC TABLE OF THE ELEMENTS

(based on Carbon-12 = 12.0000)

1 / 1A	2 / 2A	3 / 3B	4 / 4B	5 / 5B	6 / 6B	7 / 7B	8 / 8B	9 / 8B	10 / 8B	11 / 1B	12 / 2B	13 / 3A	14 / 4A	15 / 5A	16 / 6A	17 / 7A	18 / 8A
1 H Hydrogen 1.01																	2 He Helium 4.00
3 Li Lithium 6.94	4 Be Beryllium 9.01											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnes-ium 24.30											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphor-us 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.86	22 Ti Titanium 47.88	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Mangan-ese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge German-ium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybde-num 95.94	43 Tc Technet-ium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.90	56 Ba Barium 137.33	71 Lu Lutetium 174.97	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.85	75 Re Rhenium 186.21	76 Os Osmium 190.2	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (233)	88 Ra Radium 226.02	103 Lr Lawrenc-ium (260)	104 Rf Rutherford ium (261)	105 Db Dubnium (262)	106 Sg Seaborg-ium (263)	107 Bh Bohrium (262)	108 Hs Hassium (265)	109 Mt Meitner-ium (266)	110 Uun (269)	111 Uuu (272)	112 Uub (277)		114 Uuq (285)				

Inner Transition Metals														
Lanthanide Series	57 La Lantha-num 138.91	58 Ce Cerium 140.12	59 Pr Praseodym ium 140.94	60 Nd Neodym-ium 144.24	61 Pm Prometh-ium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.97	64 Gd Gadolin-ium 157.25	65 Tb Terbium 158.92	66 Dy Dyspros-ium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thalium 168.93	70 Yb Ytterbium 173.04
Actinide Series	89 Ac Actinium 227.03	90 Th Thorium 232.04	91 Pa Protactin-ium 231.04	92 U Uranium 238.03	93 Np Neptunium 237.05	94 Pu Plutonium (244)	95 Am Americ-ium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californ-ium (251)	99 Es Einstein-ium (252)	100 Fm Fermium (257)	101 Md Mendelv-ium (258)	102 No Nobelium (259)

Help sheet for naming acids.

An acid is a compound that gives off hydrogen ions when dissolved in water.
Binary acids (two elements together) means one is hydrogen and the other is some other element. The following are the binary acids you need to know. **Notice they begin with the prefix hydro (meaning hydrogen) and they end in -ic. The word acid is included as part of the name.**

HCl hydrochloric acid
HF hydrofluoric acid
HI hydroiodic acid
HBr hydrobromic acid

Ternary acids. Three elements are in the formula. One of them has to be hydrogen for this to be an acid. The name given to the ternary acid is related to the anions or polyatomic ions.

If the name ends in -ite the name of the acid ends in -ous.
If the name ends in -ate the name of the acid ends in -ic.

To help you remember these, let's use our memory aid again!!

	-1	**-1**	**-2**	**-3**		
Acids add	**ClO**	**NO**	**SO**	**PO**		
Hydrogens and	**2**	**2**	**3**	**3**	**ite**	**(ous)**
take away the charge	**3**	**3**	**4**	**4**	**ate**	**(ic)**

Here are the acids you can name with the above memory aid.

chlorous acid	$HClO_2$	**Sulfurous acid**	H_2SO_3
chloric acid	$HClO_3$	**Sulfuric acid**	H_2SO_4
Nitrous acid	HNO_2	**Phosphorous acid**	H_3PO_3
Nitric acid	HNO_3	**Phosphoric acid**	H_3PO_4

Note that the charge is gone because we added hydrogens. Each hydrogen has a charge of +1. Be careful with the spelling of these acids. **Look at them carefully.**

This leaves only 4 more acids to memorize!

If you know how to make the formula chlorous acid and chloric acid, remember that **hypo-** means one less oxygen and **per-** means one more oxygen. Now you know the formulas of:
HClO hypochlorous acid and $HClO_4$ perchloric acid.

The last two acids are:
H_2CO_3 carbonic acid (hydrogen plus the carbonate ion)
$HC_2H_3O_2$ acetic acid (hydrogen plus the acetate ion)

Name ____________________ Date __________ Class ________

7–1 Practice Problems

Write the correct formula for each of the compounds listed below.

1. potassium iodide

2. barium chloride

3. lithium bromide

4. sodium hypochlorite

5. iron(III) sulfate

6. chromium(III) sulfide

7. calcium carbonate

8. sodium acetate

9. cobalt(II) fluoride

10. sodium phosphide

11. tin(IV) oxide

12. gold(III) bromide

13. copper(II) iodide

14. strontium chloride

15. lithium acetate

16. magnesium hydroxide

17. nickel(II) nitrate

18. silver oxide

19. zinc chloride

20. magnesium phosphate

Name ______________________ Date __________ Class __________

7–3 Practice Problems

Write the names for each of the following ionic compounds.

1. KOH

____

2. LiI

____

3. AlF_3

____

4. $FeCl_2$

____

5. MgO

____

6. $Co(NO_3)_2$

____

7. $MgSO_4$

____

8. NH_4Cl

____

9. $CrPO_4$

____

10. $Ba(OH)_2$

____

11. PbS

____

12. Na_2CO_3

____

13. BaF_2

____

14. $Cu(NO_3)_2$

____

15. AgI

____

16. $NiSO_4$

____

17. $Zn_3(PO_4)_2$

____

18. Na_3N

____

19. Cu_2CO_3

____

20. $(NH_4)_2SO_4$

____

Name ____________________ Date __________ Class __________

7–3 *Practice Problems (continued)*

Write names for each of the following molecular substances.

21. SiO_2 ____________________

22. PCl_3 ____________________

23. SiF_4 ____________________

24. N_2O ____________________

25. SO_3 ____________________

26. N_2O_5 ____________________

27. IF_5 ____________________

28. SF_6 ____________________

29. ClO_2 ____________________

30. P_4S_3 ____________________

31. NO ____________________

32. SF_4 ____________________

33. XeF_4 ____________________

34. SbF_5 ____________________

35. NH_3 ____________________

36. SO_2 ____________________

37. H_2O ____________________

38. CS_2 ____________________

39. CI_4 ____________________

40. BCl_3 ____________________

Name:________________ Partner:__________________ Date:________________

NOMENCLATURE Practice [redacted]

Molecular Compounds, Ionic Compounds, & Acids

NAME THE FOLLOWING COMPOUNDS:

1. $BaSO_3$
2. $(NH_4)_3PO_4$
3. PBr_5
4. $MgSO_4$
5. CaO
6. H_3PO_4
7. $Na_2Cr_2O_7$
8. MgO
9. SO_3
10. $Cu(NO_3)_2$
11. HI
12. N_2O
13. MnO
14. $AgNO_3$
15. As_2O_5
16. Fe_2O_3
17. $HClO$
18. N_2O_3
19. HF
20. $H_2C_2O_4$
21. $NaHCO_3$
22. $SiBr_4$
23. $CuCl_2$
24. HNO_2
25. SnO_2
26. $BaCrO_4$

O = stock system

WRITE FORMULAS FOR THE FOLLOWING COMPOUNDS:

27. hydrobromic acid
28. chromium(III) carbonate
29. magnesium sulfide
30. iodine trichloride
31. lithium hydride
32. ammonium hydroxide
33. calcium chloride
34. hydroselenic acid
35. iron(II) nitride
36. aluminum hydroxide
37. tin(II) fluoride
38. sulfur tetrachloride
39. mercury(II) iodide
40. diphosphorus pentoxide
41. sulfurous acid
42. lead(II) nitrate
43. dihydrogen monoxide
44. sodium oxalate
45. perchloric acid
46. chlorous acid
47. silicon dioxide
48. carbonic acid
49. sodium chlorate
50. xenon hexafluoride
51. nickel(II) Nitrate
52. potassium perchlorate

$C_2O_4^{2-}$ = oxalate

Name________________________ Date______________

Part I Name the following compounds

1. CS_2 ______________________
2. FeO ______________________
3. HNO_3 ______________________
4. CCl_4 ______________________
5. $Ca_3(PO_4)_2$ ______________________
6. N_2O_4 ______________________
7. Na_2S ______________________
8. $FeCl_2$ ______________________
9. $NiCrO_4$ ______________________
10. N_2O_3 ______________________
11. H_2SO4 ______________________
12. $Zn(NO_3)_2$ ______________________
13. IF_7 ______________________
14. $AlCl_3$ ______________________
15. $NaOH$ ______________________
16. As_2S_5 ______________________
17. KNO_3 ______________________
18. $Mg(OH)_2$ ______________________
19. I_2O_5 ______________________
20. HCl ______________________

Write the formulas of the following

21. sodium nitrite ______________________
22. bromine trichloride ______________________
23. aluminum hydroxide ______________________
24. ammonium hydroxide ______________________
25. sulfur dioxide ______________________
26. hydroiodic acid ______________________
27. copper (I) oxide ______________________
28. oxygen difluoride ______________________
29. zinc oxide ______________________
30. barium sulfite ______________________
31. nitrogen trifluoride ______________________
32. silver sulfide ______________________
33. nitrous acid ______________________
34. iodine trichloride ______________________
35. copper (II) nitrate ______________________
36. magnesium sulfide ______________________
37. iodine heptafluoride ______________________
38. barium nitride ______________________
39. lead (II) sulfate ______________________
40. carbon tetraiodide ______________________

Part III Name the following compounds

41. NH_4NO_2 ______________________
42. SiC ______________________
43. $Ba(ClO_3)_2$ ______________________
44. Hg_2I_2 ______________________
45. PCl_3 ______________________
46. PbO_2 ______________________
47. K_2CO_3 ______________________
48. BF_3 ______________________
49. K_2SO_4 ______________________
50. $Hg(OH)_2$ ______________________

$Na_2SO_4 \cdot 10H_2O$ ______________________

Hydrate !!

51. Cl_2O_7 ______________________
52. Cu_2S ______________________
53. $KHSO_4$ ______________________
54. N_4Se_4 ______________________
55. $Fe_2(CrO_4)_3$ ______________________
56. $NaClO_4$ ______________________
57. PF_5 ______________________
58. $NaC_2H_3O_2$ ______________________
59. Mg_3N_2 ______________________
60. BN ______________________
61. H_3PO_4 ______________________

Name________________

Chemistry 1-2 WORKSHEET: Inorganic Nomenclature I [redacted]

1. Give the name of each of the following binary compounds. (8)

a) KCl

b) BaO

c) Rb_2S

d) Na_3P

e) AlF_3

f) Mg_3N_2

g) CaI_2

h) $RaCl_2$

2. Identify which of the following formulae and name combinations are **incorrect**. If you find a combination that is incorrect, give the correct formula (leaving the name as given in the question). (10)

a) Calcium hydride, CaH_2

b) Sodium oxide, NaO_2

c) Barium chloride, BCl_3

d) Strontium oxide SrO_2

e) Boron trifluoride, $BoFl_4$

f) Vanadium (III) chloride, VCl_3

g) Magnesium oxide, MgO_2

3. Write the name of the following ionic substances using the system that includes a Roman numeral to specify the charge on the cation. (7)

a) FeI_3

b) $MnCl_2$

c) HgO

d) Cu_2O

e) CuO

f) $SnBr_4$

g) MnO_2

4. Write the name of each of the following binary compounds of non-metals. (5)

a) N_2Br_4

b) P_2S_5

c) SeO_2

d) N_2O_5

e) SiO_2

5. Classify each of the following compounds as either *molecular* or *ionic*. **Circle** the species that would likely be *solid* at room temperature. (5)

a) AlH_3

b) FeO

c) CuI_2

d) OF_2

e) XeF_6

6. Name each of the following compounds using the periodic table to determine whether the compound is likely to be ionic (made from a metal and a non-metal) or molecular (made from two non-metals). Where appropriate use Roman numerals. (6)

a) $RaCl_2$

b) $SeCl_2$

c) PCl_3

d) Na_3P

e) CuF_2

f) V_2O_5

More practice!

Name ______________________ Date ____________

Task 4a

1. Name the following ionic compounds:

i) ______________ $NaCl$
ii) ______________ SrO
iii) ______________ AlN
iv) ______________ $BaCl_2$
v) ______________ K_2O
vi) ______________ CuO
vii) ______________ Cu_2O
viii) ______________ Iron (III) chloride

2. Name the ionic compounds below.

i) ______________ CsF
ii) ______________ $MgBr_2$
iii) ______________ MgI_2
iv) ______________ $MgCl_2$
v) ______________ Rb_2O
vi) ______________ SrF_2
vii) ______________ K_2S

3. What are the formulas for the following ionic compounds?

i) ______________ Mercury (II) oxide
ii) ______________ Iron (III) oxide
iii) ______________ Copper (I) chloride
iv) ______________ Manganese (IV) oxide
v) ______________ Aluminum sulfate
vi) ______________ Lead (II) chloride
vii) ______________ Lead (IV) chloride

4. In question 3v) above, why is aluminum not written with a Roman numeral, whereas all the other metals are?

Task 4b

1. What are the formulas for the following ionic compounds?

i) ______________ Ammonium nitrate
ii) ______________ Copper (II) bromide
iii) ______________ Copper (I) bromide
iv) ______________ Zinc hydrogen sulfate
v) ______________ Aluminum sulfate
vi) ______________ Sodium perchlorate
vii) ______________ Copper (II) iodide

2. Name the following compounds that contain polyatomic ions.

i) ______________ Na_2SO_4
ii) ______________ $Fe(NO_3)_3$
iii) ______________ $Cu(OH)_2$
iv) ______________ Na_2SO_3
v) ______________ $KMnO_4$
vi) ______________ $KHSO_4$
vii) ______________ $CaCr_2O_7$
viii) ______________ $NaHCO_3$

4. Name the following compounds.

i) ____________________ Cs_3PO_3
ii) ____________________ $LiNO_3$
iii) ____________________ $MgSO_4$
iv) ____________________ NaBr

5. Write the correct formulas for the following compounds.

i) ____________________ Nitric acid
ii) ____________________ Ammonium perchlorate
iii) ____________________ Lead (IV) oxide
iv) ____________________ Cobalt (III) fluoride
v) ____________________ Sodium hydrogen carbonate
vi) ____________________ Copper (II) hydroxide
vii) ____________________ Sulfuric acid
viii) ____________________ Vanadium (III) hydroxide
ix) ____________________ Chromium (III) sulfate

1. Name the following binary, non-metal compounds

i) ____________________ BF_3
ii) ____________________ NO
iii) ____________________ N_2O_5
iv) ____________________ CCl_4
v) ____________________ NO_2
vi) ____________________ IF_5
vii) ____________________ P_4O_6
viii) ____________________ N_2O_3
ix) ____________________ XeF_6
x) ____________________ SiO_2

2. What are the formulas for the following binary non-metal compounds?

i) ____________________ Iodine trichloride
ii) ____________________ Phosphorus pentachloride
iii) ____________________ Bromine pentafluoride
iv) ____________________ Sulfur dioxide
v) ____________________ Sulfur trioxide

3. Write formula or names for the following molecular compounds.

i) ____________________ Dinitrogen tetroxide
ii) ____________________ N_2O_5
iii) ____________________ PCl_3
iv) ____________________ Phosphorous pentachloride
v) ____________________ SF_6

Writing formula hints:

Type 1: Binary Ionic

Between a metal and a non-metal. They end with "-ide" on the second word. The metal's symbol and charge are written, followed by the non-metal symbol and charge. The charges are balanced by using subscripts to indicate the number of atoms.

Type 2: Polyatomic Ionic

They either begin with "ammonium" or end with "-ate" or "-ite" in the second word (except "hydroxide" and "cyanide"—those are polyatomic ions). A polyatomic ion is a group of atoms that together have one charge. The metal's symbol and charge are written first (or ammonium, NH_4^{+1}, the only polyatomic cation). The polyatomic anion's symbols and charge are written next. The charges are again balanced with subscripts. If a subscript is added to a polyatomic ion, use parenthesis around the ion.

Type 1 or 2 with Multivalent Metals T

They will have Roman numerals in the name. Multivalent metals are metals that have more than one possibility for the charge. The charge of the metal is indicated with Roman numerals following the metal's name. The formula is then written following the rules for either Type 1 or Type 2. **Metals found in the d and p block will need Roman numerals.** (Al^{3+}, Zn^{2+}, and Ag^{+} DO NOT have multiple ions. They do not use Roman numerals. Also, all metals in the s block NEVER use Roman numerals.)

Type 3: Binary Covalent

Between two non-metals or a metalloid and a non-metal. They use prefixes to indicate the number of atoms. "Mono-" is not used on the first element. The element symbols are written, and the prefixes indicate the subscript for each. The prefixes are: Mono-1 Di-2 Tri-3 Tetra-4 Penta-5 Hexa-6 Hepta-7 Octa-8 Nona-9 Deca-10

Type 4: Acids

The cation for an acid is H^{+}. The anion is based on the format of the name: "hydro___ic acids" end with a single element; "___ic acids" end with the "___ate" polyatomic ion; "___ous acids" end with the "___ite" polyatomic ion. Write the correct anion's symbol and charge and then balance the charges with subscripts.

Mole Road Map

Mass (g)	DIVIDE by the MOLAR MASS →	MOLE
	← MULTIPLY by the MOLAR MASS	↑ ↓
Representative Particle (atom, molecule Formula unit, ion)	DIVIDE by 6.02×10^{23} →	MOLE
	← MULTIPLY by 6.02×10^{23}	↑ ↓
Liters of any gas At STP (Standard Temperature and pressure)	DIVIDE by 22.4 →	MOLE
	← MULTIPLY by 22.4	

Name ______________________________ Date__________

Percent Composition

Percent composition gives the percent of mass that each element contributes to the compound

$$\frac{\text{mass of element}}{\text{molar mass of compound}} \times 100$$

Example#1

Find the percent composition of Na in NaCl

Step one:

Find the molar mass of the compound

NaCl 1 (23.0g/mol) + (35.5 g/mol) = 58.5 g/mol

Step two:

Divide mass of element by molar mass

$\% \text{ Na} = \frac{1(23.0)}{58.5} \times 100 = 39.3\%$ = **ANSWER!**

Example #2

Find the percent composition of each element in calcium fluoride

CaF_2

1 (40.1g/mol) + 2(19.0 g/mol) = 78.1 g/mol = molar mass

$\% \text{ composition of Ca} = \frac{1(40.1)}{78.1} \times 100 = 51.3\ \%$

$\% \text{ composition of F} = \frac{2(19.0)}{78.1} \times 100 = 48.7\%$

Practice problems:

Find the percent composition of each element in aluminum oxide

Find the percent composition of each element in magnesium bromide

Homework:

Find the percent composition of each element in potassium carbonate

Find the percent composition of each element in lead(II) phosphate

Name ______________________ Date ____________ Class ____________

10–2 Practice Problems

1. Find the mass of 0.89 mol of $CaCl_2$.

2. A bottle of $PbSO_4$ contains 158.1 g of the compound. How many moles of $PbSO_4$ are in the bottle?

3. Find the mass of 1.112 mol of HF.

4. Determine the number of moles of C_5H_{12} that are in 362.8 g of the compound.

5. Find the mass of 0.159 mol of SiO_2.

6. You are given 12.35 g of $C_4H_8O_2$. How many moles of the compound do you have?

7. Find the mass of 3.66 mol of N_2.

8. A bottle of $KMnO_4$ contains 66.38 g of the compound. How many moles of $KMnO_4$ does it contain?

9. Determine the number of atoms that are in 0.58 mol of Se.

10. How many moles of barium nitrate ($BaNO_3$) contain 6.80×10^{24} formula units?

11. Determine the number of atoms that are in 1.25 mol of O_2.

12. How many moles of magnesium bromide ($MgBr_2$) contain 5.38×10^{24} formula units?

13. Determine the number of formula units that are in 0.688 mol of $AgNO_3$.

14. How many moles of ethane (C_2H_6) contain 8.46×10^{24} formula units?

15. Determine the number of formula units that are in 1.48 mol of NaF.

16. How many formula units are in 3.5 g of NaOH?

Name ______________________ Date ____________ Class __________

10–2 Practice Problems (continued)

17. If you burned 6.10 x 10^{24} molecules of ethane (C_2H_6), what mass of ethane did you burn?

18. How many formula units are in 5.1 g of TiO_2?

19. What is the mass of 3.62 x 10^{24} molecules of methanol (CH_3OH)?

20. How many formula units are in 1.4 g of $PbCl_2$?

21. Determine the mass of 2.94 x 10^{24} molecules of decane ($C_{10}H_{22}$).

22. How many formula units are in 5.6 g of H_2S?

23. A container with a volume of 893 L contains how many moles of air at STP?

24. A chemical reaction produces 0.37 mol of N_2 gas. What volume will that gas occupy at STP?

25. A canister with a volume of 694 L contains how many moles of oxygen at STP?

26. A chemical reaction produces 13.8 mol of CO gas. What volume will that gas occupy at STP?

27. A tube with a volume of 3.68 L contains how many moles of neon gas at STP?

28. A chemical reaction produces 0.884 mol of H_2S gas. What volume will that gas occupy at STP?

29. A container with a volume of 101 L contains how many moles of argon gas at STP?

30. A chemical reaction produces 138 mol of HBr gas. What volume will that gas occupy at STP?

Name ______________________ Date ____________ Class __________

10–3 Practice Problems

1. Find the percentage composition of a compound that contains 1.94 g of carbon, 0.48 g of hydrogen, and 2.58 g of sulfur in a 5.00-g sample of the compound.

2. A sample of an unknown compound with a mass of 0.847 g has the following composition: 50.51 percent fluorine and 49.49 percent iron. When this compound is decomposed into its elements, what mass of each element would be recovered?

3. Find the percentage composition of a compound that contains 2.63 g of carbon, 0.370 g of hydrogen, and 0.580 g of oxygen in a 3.58-g sample of the compound.

4. A sample of an unknown compound with a mass of 2.876 g has the following composition: 66.07 percent carbon, 6.71 percent hydrogen, 4.06 percent nitrogen, and 23.16 percent oxygen. What is the mass of each element in this compound?

5. Find the percentage composition of a compound that contains 2.7369 g of chlorine, 0.4116 g of oxygen, and 0.7971 g of phosphorus in a 3.9460-g sample of the compound.

6. Find the percentage composition of a compound that contains 1.51 g of chromium, 1.13 g of potassium, and 1.62 g of oxygen in a 4.26-g sample of the compound.

7. A sample of a compound that has a mass of 0.432 g is analyzed. The sample is found to be made up of oxygen and fluorine only. Given that the sample contains 0.128 g of oxygen, calculate the percentage composition of the compound.

8. What is the percentage composition of a carbon-oxygen compound, given that a 95.2-g sample of the compound contains 40.8 g of carbon and 54.4 g of oxygen?

9. What is the percentage composition of a sulfur-chlorine compound, given that a 30.9-g sample of the compound contains 9.63 g of sulfur and 21.3 g of chlorine?

10. Determine the empirical formula of a compound containing 2.644 g of gold and 0.476 g of chlorine.

11. Determine the empirical formula of a compound containing 0.928 g of gallium and 0.412 g of phosphorus.

12. Determine the empirical formula of a compound containing 1.723 g of carbon, 0.289 g of hydrogen, and 0.459 g of oxygen.

Name ______________________ Date ____________ Class ____________

10–3 Practice Problems (continued)

13. Find the empirical formula of a compound, given that the compound is found to be 47.9 percent zinc and 52.1 percent chlorine by mass.

14. Find the empirical formula of a compound, given that a 48.5-g sample of the compound contains 1.75 g of carbon and 46.75 g of bromine.

15. Determine the empirical formula of a compound containing 20.23 percent aluminum and 79.77 percent chlorine.

16. Determine the empirical formula of a compound containing 24.74 percent potassium, 34.76 percent manganese, and 40.50 percent oxygen.

17. Determine the empirical formula of a compound containing 4.288 g of carbon and 5.712 g of oxygen.

18. Determine the empirical formula of a compound containing 2.16 g of aluminum, 3.85 g of sulfur, and 7.68 g of oxygen.

19. Determine the empirical formula of a compound containing 3.611 g of calcium and 6.389 g of chlorine.

20. Find the molecular formula of a compound that contains 42.56 g of palladium and 0.80 g of hydrogen. The molar mass of the compound is 216.8 g/mol.

21. Octane, a compound of hydrogen and carbon, has a molar mass of 114.26 g/mol. If the compound contains 18.17 g of hydrogen, what is its molecular formula?

22. Find the molecular formula of a compound that contains 30.45 percent nitrogen and 69.55 percent oxygen. The molar mass of the compound is 92.02 g/mol.

23. Find the molecular formula of a compound, given that a 212.1-g sample of the compound contains 42.4 g of hydrogen and 169.7 g of carbon and the molar mass is 30.0 g/mol.

24. A compound is known to have a molar mass of 391.5 g/mol. Find the molecular formula of the compound, given the results of an analysis of a 310.8-g sample that revealed that the sample contains only boron and iodine. The mass of the iodine in the sample is found to be 302.2 g.

25. Find the molecular formula of a compound that contains 56.36 g of oxygen and 43.64 g of phosphorus. The molar mass of the compound is 283.9 g/mol.

Here is the rhyme again!

Percent to mass
Mass to mole
Divide by small
Multiply 'til whole

Here's the example problem: A compound is analyzed and found to contain 68.54% carbon, 8.63% hydrogen, and 22.83% oxygen. The molecular weight of this compound is known to be approximately 140 g/mol. What is the empirical formula? What is the molecular formula?

1) Percent to mass. Assume 100 grams of the substance is present, therefore its composition is:
carbon: 68.54 grams
hydrogen: 8.63 grams
oxygen: 22.83 grams

(2) Mass to moles. Divide each mass by the proper atomic weight.

carbon: 68.54 / 12.011 = 5.71 mol
hydrogen: 8.63 / 1.008 = 8.56 mol
oxygen: 22.83 / 16.00 = 1.43 mol

(3) Divide by small:

carbon: 5.71 ÷ 1.43 = 3.99
hydrogen: 8.56 ÷ 1.43 = 5.99
oxygen: 1.43 ÷ 1.43 = 1.00

(4) Multiply 'til whole. Not needed since all values came out whole.

The **empirical formula** of the compound is C_4H_6O.

Next we need to determine the molecular formula, knowing the empirical formula and the molecular weight.

Here's how:

1) Calculate the "empirical formula weight." This is not a standard chemical term, but the ChemTeam believes it is understandable.

C_4H_6O gives an "EFW" of 70.092.

2) Divide the molecular weight by the "EFW."

140 ÷ 70 = 2

3) Multiply the subscripts of the empirical formula by the factor just computed.
C_4H_6O times 2 gives a formula of $C_8H_{12}O_2$.

Writing formula hints:

Type 1: Binary Ionic

Between a metal and a non-metal. They end with "-ide" on the second word. The metal's symbol and charge are written, followed by the non-metal symbol and charge. The charges are balanced by using subscripts to indicate the number of atoms.

Type 2: Polyatomic Ionic

They either begin with "ammonium" or end with "-ate" or "-ite" in the second word (except "hydroxide" and "cyanide"—those are polyatomic ions). A polyatomic ion is a group of atoms that together have one charge. The metal's symbol and charge are written first (or ammonium, NH_4^{+1}, the only polyatomic cation). The polyatomic anion's symbols and charge are written next. The charges are again balanced with subscripts. If a subscript is added to a polyatomic ion, use parenthesis around the ion.

Type 1 or 2 with Multivalent Metals T

They will have Roman numerals in the name. Multivalent metals are metals that have more than one possibility for the charge. The charge of the metal is indicated with Roman numerals following the metal's name. The formula is then written following the rules for either Type 1 or Type 2. **Metals found in the d and p block will need Roman numerals.** (Al^{3+}, Zn^{2+}, and Ag^{+} DO NOT have multiple ions. They do not use Roman numerals. Also, all metals in the s block NEVER use Roman numerals.)

Type 3: Binary Covalent

Between two non-metals or a metalloid and a non-metal. They use prefixes to indicate the number of atoms. "Mono-" is not used on the first element. The element symbols are written, and the prefixes indicate the subscript for each. The prefixes are: Mono-1 Di-2 Tri-3 Tetra-4 Penta-5 Hexa-6 Hepta-7 Octa-8 Nona-9 Deca-10

Type 4: Acids

The cation for an acid is H^{+}. The anion is based on the format of the name: "hydro___ic acids" end with a single element; "___ic acids" end with the "___ate" polyatomic ion; "___ous acids" end with the "___ite" polyatomic ion. Write the correct anion's symbol and charge and then balance the charges with subscripts.

Reference
Chapter 7
Chem 1A

REFERENCE

FORMULAS AND NOMENCLATURE OF IONIC AND COVALENT COMPOUNDS

Adapted from McMurry/Fay, section 2.10, p. 56-63 and the 1411 Lab Manual, p. 27-31.

TYPES OF COMPOUNDS

Ionic compounds are compounds composed of *ions*, charged particles that form when an atom (or group of atoms) gains or loses electrons. (A *cation* is a positively charged ion; an *anion* is a negatively charged ion.) **Covalent** or **molecular compounds** form when elements share electrons in a covalent bond to form *molecules*. Molecular compounds are electrically neutral.

Metal + Nonmetal ⟶ ionic compound (usually)
Metal + Polyatomic ion ⟶ ionic compound (usually)
Nonmetal + Nonmetal ⟶ covalent compound (usually)
Hydrogen + Nonmetal ⟶ covalent compound (usually)

TYPES OF IONS

Main-Group Metals (Groups IA, IIA, and IIIA)

Group IA, IIA, and IIIA metals tend to form *cations* by losing all of their outermost (valence) electrons. *The charge on the cation is the same as the group number.* The cation is given the same name as the neutral metal atom.

Ions of Some Main-Group Metals (Groups IA – IIIA)

Group	Element	Cation	Ion name	Group	Element	Cation	Ion name
IA	H	H^+	hydrogen ion	IIA	Mg	Mg^{2+}	magnesium ion
	Li	Li^+	lithium ion		Ca	Ca^{2+}	calcium ion
	Na	Na^+	sodium ion		Sr	Sr^{2+}	strontium ion
	K	K^+	potassium ion		Ba	Ba^{2+}	barium ion
	Cs	Cs^+	cesium ion	IIIA	Al	Al^{3+}	aluminum ion

Transition (B-group) and Post-Transition (Group IVA and VA) Metals

These elements usually form ionic compounds; many of them can form more than one cation. (The charges of the transition metals must be memorized; Group IV and V metal cations tend to be either the group number, or the group number minus two.)

Many of these ions have common or trivial names (*-ic* endings go with the higher charge, *-ous* endings go with the lower charge). The systematic names (also known as the *Stock system*) for these ions are derived by naming the metal first, followed in parentheses by the charge written in Roman numerals. For the metals below that typically form only one charge, it is not usually necessary to specify the charge in the compound name.

The mercury I cation is a special case; it consists of two Hg^+ ions joined together, and so is always found as Hg_2^{2+}. (Hence, mercury(I) chloride is Hg_2Cl_2, while mercury (II) chloride is $HgCl_2$.)

Polyatomic Ions

Polyatomic ions are ions that are composed of two or more atoms that are linked by covalent bonds, but that still have a net deficiency or surplus of electrons, resulting in an overall charge on the group. A metal plus a polyatomic ion yields an ionic compound.

Formulas and Names of Some Polyatomic Ions

NH_4^+	ammonium	CO_3^{2-}	carbonate
H_3O^+	hydronium	HCO_3^-	hydrogen carbonate (bicarbonate)
OH^-	hydroxide	OCN^-	cyanate
CN^-	cyanide	SCN^-	thiocyanate
O_2^{2-}	peroxide	$S_2O_3^{2-}$	thiosulfate
N_3^-	azide	CrO_4^{2-}	chromate
NO_2^-	nitrite	$Cr_2O_7^{2-}$	dichromate
NO_3^-	nitrate	SO_4^{2-}	sulfate
ClO^-	hypochlorite	SO_3^{2-}	sulfite
ClO_2^-	chlorite	HSO_4^-	hydrogen sulfate (bisulfate)
ClO_3^-	chlorate	PO_4^{3-}	phosphate
ClO_4^-	perchlorate	HPO_4^{2-}	monohydrogen phosphate
MnO_4^-	permanganate	$H_2PO_4^-$	dihydrogen phosphate
$C_2H_3O_2^-$	acetate (OAc^-)		
$C_2O_4^{2-}$	oxalate	HSO_3^-	hydrogen sulfite (bisulfite)

There are some regularities in the names of these polyatomic ions.

a. *Thio-* implies replacing an oxygen with a sulfur:

SO_4^{2-} = sulfate $S_2O_3^{2-}$ = thiosulfate
OCN^- = cyanate SCN^- = thiocyanate

b. Replacing the first element with another element from the same group gives a polyatomic ion with the same charge, and a similar name:

Group VIIA	Group VIA	Group VA	Group IVA	Group III A
ClO_3^- chlorate	SO_4^{2-} sulfate	PO_4^{3-} phosphate	CO_3^{2-} carbonate	BO_3^{3-} Borate
BrO_3^- bromate	SeO_4^{2-} selenate	AsO_4^{3-} arsenate	SiO_3^{2-} silicate	
IO_3^- iodate	TeO_4^{2-} tellurate			

c. Some nonmetals form a series of polyatomic ions with oxygen (all having the same charge): ClO^-, hypochlorite; ClO_2^-, chlorite; ClO_3^-, chlorate; ClO_4^-, perchlorate. The general rule for such series is:

XO_n^{y-}	*stem + -ate*	SO_4^{2-}	sulfate
XO_{n-1}^{y-}	*stem + -ite*	SO_3^{2-}	sulfite
XO_{n-2}^{y-}	*hypo- + stem + -ite*	SO_2^{2-}	hyposulfite
XO_{n+1}^{y-}	*per- + stem + -ate*	SO_5^{2-}	persulfate
X^{y-}	*stem + -ide* (the monatomic ion)	S^{2-}	sulfide

Note that in some cases, the *-ate* form has three oxygens, and in some cases four oxygens. (These forms must be memorized.)

BrO^-	Hypobromite
BrO_2^-	bromite
BrO_3^-	bromate
BrO_4^-	perbromate

Ions of Some Transition Metals and Post-Transition Metals (Groups IVA and VA)

Metal	Ion	Systematic name	Common name
Cadmium	Cd^{2+}	cadmium ion	
Chromium	Cr^{2+}	chromium(II) ion	chromous ion
	Cr^{3+}	chromium(III) ion	chromic ion
Cobalt	Co^{2+}	cobalt(II) ion	cobaltous ion
	Co^{3+}	cobalt(III) ion	cobaltic ion
Copper	Cu^{+}	copper(I) ion	cuprous ion
	Cu^{2+}	copper(II) ion	cupric ion
Gold	Au^{3+}	gold(III) ion	
Iron	Fe^{2+}	iron(II) ion	ferrous ion
	Fe^{3+}	iron(III) ion	ferric ion
Manganese	Mn^{2+}	manganese(II) ion	manganous ion
	Mn^{3+}	manganese(III) ion	manganic ion
Mercury	Hg_2^{2+}	mercury(I) ion	mercurous ion
	Hg^{2+}	mercury(II) ion	mercuric ion
Nickel	Ni^{2+}	nickel(II) ion	
Silver	Ag^{+}	silver ion	
Zinc	Zn^{2+}	zinc ion	
Tin	Sn^{2+}	tin(II) ion	stannous ion
	Sn^{4+}	tin(IV) ion	stannic ion
Lead	Pb^{2+}	lead(II) ion	plumbous ion
	Pb^{4+}	lead(IV) ion	plumbic ion
Bismuth	Bi^{3+}	bismuth(III) ion	
	Bi^{5+}	bismuth(V) ion	

Main-Group Nonmetals (Groups IVA, VA, VIA, and VIIA)

Group IVA, VA, VIA, and VIIA nonmetals tend to form *anions* by gaining enough electrons to fill their valence shell with eight electrons. *The charge on the anion is the group number minus eight.* The anion is named by taking the element stem name and adding the ending *-ide*.

Ions of Some Nonmetals (Groups IVA - VIIA)

Group	Element	Anion	Ion name	Group	Element	Anion	Ion name
IVA	C	C^{4-}	carbide ion	VIA	Se	Se^{2-}	selenide ion
	Si	Si^{4-}	silicide ion		Te	Te^{2-}	telluride ion
VA	N	N^{3-}	nitride ion	VIIA	F	F^{-}	fluoride ion
	P	P^{3-}	phosphide ion		Cl	Cl^{-}	chloride ion
	As	As^{3-}	arsenide ion		Br	Br^{-}	bromide ion
VIA	O	O^{2-}	oxide ion		I	I^{-}	iodide ion
	S	S^{2-}	sulfide ion	IA	H	H^{-}	hydride ion

CATIONS (+ve)		ANIONS (-ve)	
Aluminum	Al^{3+}	Bromide	Br^-
Ammonium	NH_4^+	Bromate	BrO_3^-
Arsenic (III)	As^{3+}	Bromite	BrO_2^-
Arsenic (V)	As^{5+}	Carbonate	CO_3^{2-}
Barium	Ba^{2+}	Chlorate	ClO_3^-
Bismuth (III)	Bi^{3+}	Chlorite	ClO_2^-
Bismuth (V)	Bi^{5+}	Chloride	Cl^-
Cadmium	Cd^{2+}	Chromate	CrO_4^{2-}
Calcium	Ca^{2+}	Cyanide	CN^-
Chromium (II)	Cr^{2+}	Dichromate	$Cr_2O_7^{2-}$
Chromium (III)	Cr^{3+}	Fluoride	F^-
Cobalt (II)	Co^{2+}	Hydride	H^-
Cobalt (III)	Co^{3+}	Hydrogen carbonate (bicarbonate)	HCO_3^-
Copper (I) (Cuprous)	Cu^+	Hydrogen phosphate	HPO_4^{2-}
Copper (II) (Cupric)	Cu^{2+}	Hydrogen sulfate (bisulfate)	HSO_4^-
Hydrogen	H^+	Hydrogen sulfite (bisulfite)	HSO_3^-
Hydronium	H_3O^+	Hydroxide	OH^-
Iron (II) (Ferrous)	Fe^{2+}	Hypobromite	BrO^-
Iron (III) (Ferric)	Fe^{3+}	Hypochlorite	ClO^-
Lead (II)	Pb^{2+}	Hypoiodite	IO^-
Lead (IV)	Pb^{4+}	Iodate	IO_3^-
Lithium	Li^+	Iodite	IO_2^-
Magnesium	Mg^{2+}	Iodide	I^-
Manganese (II)	Mn^{2+}	Nitrate	NO_3^-
Manganese (IV)	Mn^{4+}	Nitride	N^{3-}
Mercury(I) (Mercurous)	Hg_2^{2+}	Nitrite	NO_2^-
Mercury(II) (Mercuric)	Hg^{2+}	Oxide	O^{2-}
Nickel	Ni^{2+}	Permanganate (manganate (VII))	MnO_4^-
Potassium	K^+	Perbromate	BrO_4^-
Silver	Ag^+	Perchlorate	ClO_4^-
Sodium	Na^+	Periodate	IO_4^-
Strontium	Sr^{2+}	Phosphate	PO_4^{3-}
Tin(II) (Stannous)	Sn^{2+}	Phosphide	P^{3-}
Tin(IV) (Stannic)	Sn^{4+}	Phosphite	PO_3^{3-}
Zinc	Zn^{2+}	Sulfate	SO_4^{2-}
		Sulfite	SO_3^{2-}
		Thiosulfate	$S_2O_3^{2-}$
		Sulfide	S^{2-}

N.B. Most transition metal ions include a number written as a Roman numeral after the name. Since most transition metals have varying charges this number identifies the charge on the metal on.

Writing Formulas of Ionic Compounds

1. The positive ion is given first, followed by the monatomic or polyatomic anion.
2. The subscripts in the formula must produce an electrically neutral formula unit. (That is, the total positive charge must equal the total negative charge.)
3. The subscripts should be the smallest set of whole numbers possible.
4. If there is only one of a polyatomic ion in the formula, do not place parentheses around it; e.g., $NaNO_3$, not $Na(NO_3)$. If there is more than one of a polyatomic ion in the formula, put the ion in parentheses, and place the subscript after the parentheses; e.g., $Ca(OH)_2$, $Ba_3(PO_4)_2$, etc. [Remember the Prime Directive in writing formulas: $Ca(OH)_2 \neq CaOH_2$!]

Na^+	Cl^-	$NaCl$
Ca^{2+}	Br^-	$CaBr_2$
Na^+	S^{2-}	Na_2S
Mg^{2+}	O^{2-}	MgO
Fe^{3+}	O^{2-}	Fe_2O_3
Na^+	SO_4^{2-}	Na_2SO_4
Mg	NO_3^-	$Mg(NO_3)_2$
NH_4^+	SO_4^{2-}	$(NH_4)_2SO_4$

Nomenclature of Ionic and Covalent Compounds

1. **Binary Ionic Compounds Containing a Metal and a Nonmetal.** A *binary compound* is a compound formed from *two different elements*. There may or may not be more than one of each element. A *diatomic compound* (or diatomic molecule) contains two atoms, which may or may not be the same.

Cl_2	Not binary (only one type of atom), but diatomic (two atoms).
BrCl	Binary and diatomic. (Two atoms, and they're different elements.)
H_2O	Binary, since there are only two types of atoms.
CH_4	Binary, since there are only two types of atoms.
$CHCl_3$	Not binary or diatomic.

Metals combine with nonmetals to give ionic compounds. When naming binary ionic compounds, name the cation first (specifying the charge, if necessary), then the nonmetal anion (element stem + *-ide*). Do NOT use prefixes to indicate how many of each element is present; this information is implied in the name of the compound.

NaCl	Sodium chloride
$AlBr_3$	Aluminum bromide
Ca_3P_2	Calcium phosphide
SrI_2	Strontium iodide
$FeCl_2$	Iron(II) chloride or ferrous chloride

2. **Ionic Compounds Containing a Metal and a Polyatomic Ion.** Metals combine with polyatomic ions to give ionic compounds. Name the cation first (specifying the charge, if necessary), then the polyatomic ion as listed in the table above. Do NOT use prefixes to

the halogen is placed first. If both elements are in the same group, the one with the higher period number is named first.] The first element in the formula is given the neutral element name, and the second one is named by replacing the ending of the neutral element name with *-ide*. A prefix is used in front of each element name to indicate how many of that element is present:

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

If there is only one of the first element in the formula, the *mono-* prefix is dropped.

SO_2	sulfur dioxide	NO_2	nitrogen dioxide
SO_3	sulfur trioxide	N_2O_4	dinitrogen tetroxide
N_2O	dinitrogen monoxide	N_2O_5	dinitrogen pentoxide
NO	nitrogen monoxide		

5. **Hydrocarbons.** *Hydrocarbons* contain only carbon and hydrogen, and are the simplest type of organic compound. *Alkanes* contain only carbon-carbon single bonds, and are the simplest of the hydrocarbons. The simplest of the alkanes are the straight-chain alkanes, in which all of the carbon atoms are linked together in a line, with no branches. (They don't get simpler than that!) Alkanes have the general formula C_nH_{2n+2}, and are the constituents of several important fuels, such as natural gas and gasoline.

   Organic chemistry has a completely different set of rules for nomenclature; straight-chain alkanes are named using a prefix plus the suffix *–ane*. (Notice that after C_4, the prefixes are the same as those listed above for binary covalent compounds.)

CH_4	methane	C_6H_{14}	hexane
C_2H_6	ethane	C_7H_{16}	heptane
C_3H_8	propane	C_8H_{18}	octane
C_4H_{10}	butane	C_9H_{20}	nonane
C_5H_{12}	pentane	$C_{10}H_{22}$	decane

Molecular Masses from Chemical Formulas

The **molecular mass** (or **molecular weight**) of a compound is obtained by adding up the atomic masses of all of the atoms present within a unit of the substance. For ionic compounds, the term *formula mass* or *formula weight* is used instead, since there aren't really any molecules present.

For example, the molecular weight of water would be obtained by the following process:

$$\begin{aligned}\text{Molecular mass of } H_2O &= (2 \times \text{atomic mass of H}) + (1 \times \text{atomic mass of O}) \\ &= (2 \times 1.008) + (1 \times 16.00) \text{ amu} \\ &= 18.02 \text{ amu}\end{aligned}$$

WORKSHEET- FORMULAS **NAME** ____________________

WRITE THE CHEMICAL FORMULA FOR EACH OF THE FOLLOWING:

1) Barium sulfide 1) __________

2) Manganese (III) iodide 2) __________

3) Ammonium hydrogen phosphate 3)__________

4) Carbon disulfide 4)__________

5)Lead (II) sulfate 5)__________

6) Magnesium carbonate 6)__________

7) Potassium permanganate 7)__________

8) Silver bicarbonate 8) __________

9) Bismuth (III) bromide 9) __________

10) Tetranitrogen tetrasulfide 10) __________

11) Ferrous perchlorate 11) __________

12) Chromium (III) chlorite 12) __________

13)Tin (II) thiosulfate 13) __________

14) Cuprous sulfite 14) __________

15) Sodium bisulfate 15) __________

16) Carbon tetrachloride 16) __________

17) Sodium acetate 17) __________

18) Ferric dihydrogen phosphate 18) __________

19) Chromium (II) phosphate 19) __________

20) Mercuric perchlorate 20) __________

21) Nickel (II) borate 21) __________

22) Cadmium thiocyanate 22) __________

23) Ammonium sulfide 23) __________

24) Bismuth (III) bisulfite 24) __________

25) Strontium chlorate 25) __________

26) Aluminum bisulfide 26) __________

27) Diphosphorus trioxide 27) __________

28) Zinc hydroxide 28) __________

29) Silver chromate 29) __________

30) Copper (II) acetate 30) __________

31) Cobaltous iodide 31) __________

32) Cuprous dichromate 32) __________

33) Sodium peroxide 33) __________

34) Dinitrogen trioxide 34) __________

35) Dichlorine heptoxide 35) __________

36) Cobaltic nitrite 36) __________

37) Barium cyanide 37) __________

38) Hypochlorous acid 38) __________

39) Sulfurous acid 39) __________

40) Hydrobromic acid 40) __________

41) Nitric acid 41) __________

42) Periodic acid 42) __________

43) Bromous acid 43) __________

44) Iodic acid 44) __________

45) Hydrosulfuric acid 45)__________

46) Perbromic acid 46) __________

47) Hydrofluoric acid 47) __________

48) hypobromous acid 48) __________

49) Mercurous chloride 49) __________

50) Ferric cyanate 50) __________

(OVER)

GIVE THE NAME FOR EACH OF THE FOLLOWING :

ANSWERS

1) BaS 1) ____________________
2) MnI_3 2) ____________________
3) $(NH_4)_2HPO_4$ 3) ____________________
4) CS_2 4) ____________________
5) $PbSO_4$ 5) ____________________
6) $MgCO_3$ 6) ____________________
7) $KMnO_4$ 7) ____________________
8) $AgHCO_3$ 8) ____________________
9) $BiBr_3$ 9) ____________________
10) N_4S_4 10) ____________________
11) $Fe(ClO_4)_2$ 11) ____________________
12) $Cr(ClO_2)_3$ 12) ____________________
13) SnS_2O_3 13) ____________________
14) Cu_2SO_3 14) ____________________
15) $NaHSO_4$ 15) ____________________
16) CCl_4 16) ____________________
17) $NaC_2H_3O_2$ 17) ____________________
18) $Fe(H_2PO_4)_3$ 18) ____________________
19) $Cr_3(PO_4)_2$ 19) ____________________
20) $Hg(ClO_4)_2$ 20) ____________________
21) $Ni_3(BO_3)_2$ 21) ____________________
22) $Cd(SCN)_2$ 22) ____________________
23) $(NH_4)_2S$ 23) ____________________
24) $Bi(HSO_3)_3$ 24) ____________________
25) $Sr(ClO_3)_2$ 25) ____________________
26) $Al(HS)_3$ 26) ____________________
27) P_2O_3 27) ____________________
28) $Zn(OH)_2$ 28) ____________________
29) Ag_2CrO_4 29) ____________________
30) $Cu(C_2H_3O_2)_2$ 30) ____________________
31) CoI_2 31) ____________________
32) $Cu_2Cr_2O_7$ 32) ____________________
33) Na_2O_2 33) ____________________
34) N_2O_3 34) ____________________
35) Cl_2O_7 35) ____________________
36) $Co(NO_2)_3$ 36) ____________________
37) $Ba(CN)_2$ 37) ____________________
38) HClO 38) ____________________
39) H_2SO_3 39) ____________________
40) HBr 40) ____________________
41) HNO_3 41) ____________________
42) HIO_4 42) ____________________
43) $HBrO_2$ 43) ____________________
44) HIO_3 44) ____________________
45) H_2S 45) ____________________
46) $HBrO_4$ 46) ____________________
47) HF 47) ____________________
48) HBrO 48) ____________________
49) Hg_2Cl_2 49) ____________________
50) $Fe(CNO)_3$ 50) ____________________

REVIEW ANSWERS

1. a. ions formed from a single atom
   b. Examples include Na^+, Mg^{2+}, and Cl^-.
2. The nitrate ion, NO_3^-, has one more oxygen atom than the nitrite ion, NO_2^- does.
3. a. K^+ b. Ca^{2+} c. S^{2-} d. Cl^- e. Ba^{2+} f. Br^-
4. a. Na^+, 1+ b. Al^{3+}, 3+ c. Cl^-, 1− d. N^{3-}, 3− e. Fe^{2+}, 2+ f. Fe^{3+}, 3+
5. a. potassium ion (or potassium cation)
   b. magnesium ion (or magnesium cation)
   c. aluminum ion (or aluminum cation)
   d. chloride ion (or chloride anion)
   e. oxide ion (or oxide anion)
   f. calcium ion (or calcium cation)
6. a. NaI b. CaS c. $ZnCl_2$ d. BaF_2 e. Li_2O
7. a. potassium chloride
   b. calcium bromide
   c. lithium oxide
   d. magnesium chloride
8. a. CrF_2, chromium(II) fluoride
   b. NiO, nickel(II) oxide
   c. Fe_2O_3, iron(III) oxide
9. The less electronegative element is written first.
10. a. carbon dioxide
    b. carbon tetrachloride
    c. phosphorus pentachloride
    d. selenium hexafluoride
    e. diarsenic pentoxide
11. a. CBr_4 b. SiO_2 c. P_4O_{10} d. As_2S_3
12. Binary acids, such as HCl and HBr, contain only two elements, usually hydrogen and a halogen. Oxyacids, such as H_2SO_4 and HNO_3, contain hydrogen, oxygen, and a third element.

13. a. an ionic compound composed of a cation and the anion from an acid
    b. Examples include NaCl and $MgSO_4$.
14. a. hydrofluoric acid
    b. hydrobromic acid
    c. nitric acid
    d. sulfuric acid
    e. phosphoric acid
15. a. H_2SO_3 b. $HClO_3$ c. HCl d. HClO e. $HClO_4$ f. H_2CO_3 g. CH_3COOH
16. a. NaF b. CaO c. K_2S d. $MgCl_2$ e. $AlBr_3$ f. Li_3N g. FeO
17. a. ammonium ion
    b. chlorate ion
    c. hydroxide ion
    d. sulfate ion
    e. nitrate ion
    f. carbonate ion
    g. phosphate ion
    h. acetate ion
    i. hydrogen carbonate ion (or bicarbonate ion)
    j. chromate ion
18. a. NH_4^+, 1+
    b. CH_3COO^-, 1−
    c. OH^-, 1−
    d. CO_3^{2-}, 2−
    e. SO_4^{2-}, 2−
    f. PO_4^{3-}, 3−
    g. Cu^{2+}, 2+
    h. Sn^{2+}, 2+
    i. Fe^{3+}, 3+
    j. Cu^+, 1+
    k. Hg_2^{2+}, 2+
    l. Hg^{2+}, 2+
19. a. iron(II) ion b. iron(III) ion c. lead(II) ion d. lead(IV) ion e. tin(II) ion f. tin(IV) ion
20. a. carbon(IV) bromide
    b. silicon(IV) oxide
    c. phosphorus(V) oxide
    d. arsenic(III) sulfide

21. a. PI_3 b. SCl_2 c. CS_2 d. N_2O_5
22. a. numbers assigned to bonded atoms in molecular compounds or polyatomic ions to indicate the general distribution of electrons
    b. They aid in naming compounds, writing formulas, balancing chemical equations, and studying chemical reactions.
23. a. sodium chloride
    b. potassium fluoride
    c. calcium sulfide
    d. cobalt(II) nitrate
    e. iron(III) phosphate
    f. mercury(I) sulfate
    g. mercury(II) phosphate
24. a. +1, −1 b. +3, −1 c. +4, −2 d. +1, −1 e. +5, −2 f. +1, +5, −2
25. a. +5, −2 b. +7, −2 c. +5, −2 d. +6, −2 e. +4, −2
26. a. the sum of the average atomic masses of all atoms represented in the formula of a molecule, formula unit, or ion
    b. amu
27. the mass in grams of 1 mol of the basic particles of a compound—it is numerically equal to formula mass
28. a. 180.18 amu
    b. 158.18 amu
    c. 18.05 amu
    d. 83.45 amu
29. a. 1 mol K^+, 1 mol NO_3^-; 1 mol N, 3 mol O
    b. 2 mol Na^+, 1 mol SO_4^{2-}; 1 mol S, 4 mol O
    c. 1 mol Ca^{2+}, 2 mol OH^-; 1 mol O, 1 mol H
    d. 2 mol NH_4^+, 1 mol SO_3^{2-}; 1 mol N, 4 mol H, 1 mol S, 3 mol O
    e. 3 mol Ca^{2+}, 2 mol PO_4^{3-}; 1 mol P, 4 mol O
    f. 2 mol Al^{3+}, 3 mol CrO_4^{2-}; 1 mol Cr, 4 mol O

30. a. 101.11 g/mol
 b. 142.05 g/mol
 c. 74.10 g/mol
 d. 116.17 g/mol
 e. 310.18 g/mol
 f. 401.96 g/mol

31. a. 0.250 mol c. 0.3627 mol
 b. 2.752 mol

32. a. 39.34% Na, 60.66% Cl
 b. 63.50% Ag, 8.25% N, 28.26% O
 c. 41.68% Mg, 54.86% O, 3.46% H

33. 36.07% H_2O

34. percentage composition, mass composition, and composition by moles

35. *x*(empirical formula) = molecular formula, where *x* is a multiple of the subscripts in the molecular formula

36. $AgNO_3$

37. C_2H_6O

38. $C_4H_8O_4$

39. C_3H_6

40. $C_6H_8O_7$

41. a. lithium bromide
 b. tin(II) nitrate
 c. iron(II) chloride
 d. magnesium oxide
 e. potassium hydroxide
 f. iron(III) oxide
 g. silver nitrate
 h. iron(II) hydroxide
 i. chromium(II) fluoride

42. a. 58.44 g c. 259.4 g
 b. 36.04 g d. 163 g

43. a. 207.29 amu, 207.29 g/mol
 b. 264.36 amu, 264.36 g/mol
 c. 654.98 amu, 654.98 g/mol
 d. 89.57 amu, 89.57 g/mol

44. a. AlF_3 d. CoS
 b. MgO e. $SrBr_2$
 c. V_2O_5 f. SO_3

45. 1 Fe, 3 C, 5 H, 7 O; 8.62%

46. a. nitrous acid; +1, +3, −2
 b. sulfurous acid; +1, +4, −2
 c. carbonic acid; +1, +4, −2
 d. hydriodic acid; +1, −1

47. a. 30.88% Na, 47.62% Cl, 21.49% O
 b. 2.46% H, 39.07% S, 58.47% O
 c. 48.63% C, 8.18% H, 43.19% O
 d. 11.28% Be, 88.72% Cl

48. a. magnesium iodide
 b. sodium fluoride
 c. carbon disulfide
 d. dinitrogen tetroxide
 e. sulfur dioxide
 f. phosphorus tribromide
 g. calcium chloride
 h. silver iodide

49. a. +4, −2 e. +1, −1
 b. −3, +1 f. +5, −2
 c. +7, −2 g. +2, −1
 d. +2, −2

50. $C_5H_{10}O_5NNa$

51. The formula is SO_3, so the molecule contains one sulfur atom and three oxygen atoms. The oxidation numbers are +6 and −2, respectively.

52. mass of nickel = 1.05 g; mass of nickel oxide = 1.34 g; mass of oxygen = 0.29 g; empirical formula: NiO

53. a. KO_2, RbO_2, CsO_2
 b. 1+
 c. 2− in oxides, 1− in peroxides, 1− for the two oxygen atoms together in superoxides

54. a. MgO, BeO, CaO
 b. SrO_2, BaO_2
 c. Mg_3N_2, Be_3N_2, Ca_3N_2
 d. −1

55. a. Cd = +2, Zn = +2, Pb = +2
 b. Fe = +3, Mn = +7, Co = +2
 c. Cu = +2

56. a. NO, NO_2, N_2O_3, N_2O_5, P_4O_{10}, As_2O_5, Sb_2O_5, Bi_2O_5
 b. for NO, N = +2; for NO_2, N = +4; for N_2O_3, N = +3; for N_2O_5, N = +5

57. Answers will vary.

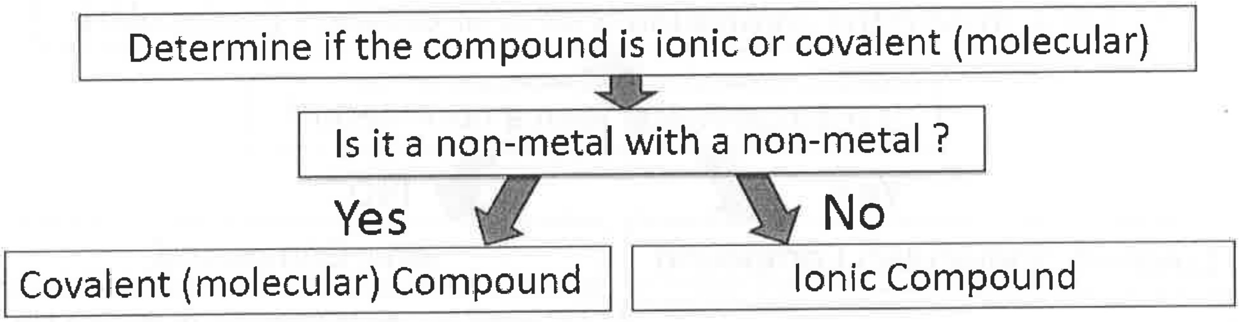
58. a. sodium bicarbonate, $NaHCO_3$
 b. magnesium hydroxide, $Mg(OH)_2$
 c. magnesium sulfate heptahydrate, $MgSO_4 \cdot 7H_2O$
 d. calcium carbonate, $CaCO_3$
 e. sodium hydroxide, NaOH
 f. methanol, CH_3OH

59. Answers will vary.

60. NH_3 has a higher percentage of nitrogen than NH_4NO_3. Answers will vary.

Naming Compounds Tutorial and Worksheet

Since we use different methods in naming covalent (molecular) compounds and ionic compounds, the **first step** in naming or writing the formula of a compound is to **determine which of the 2 compound classes it belongs.** This can be done as follows:



The only exception we will see to the above flow chart is when we see the polyatomic ion **ammonium (NH_4^+)** combined with any anion; in those cases the compound is **ionic** even though the compound is composed of non-metals only.

Once it is determined that the compound is **ionic** or **covalent**, the student can be asked to do either of the following:

1) Given the **name** of the compound, write the **formula.**

<u>Or</u>

2) Given the **formula** of the compound, write the **name.**

In this tutorial we will review the process for achieving these 2 objectives and practice with some worksheet problems. First, we will review and practice how to write formulas for compounds when given the compound's name. Second, we will review and practice how to write the name of a compound when given the compound's formula.

Review of Writing Formulas for Compounds

Given the **Name**, Writing the **Formula**:

Determine if the compound is ionic or covalent (molecular)

Is it a non-metal with a non-metal ?

Yes

Covalent (molecular) Compound

1) Write the symbol/formula of the first element in the compound's name, then the symbol/formula of the second ion in the compound's name.
2) Indicate how many of each element the molecule contains using subscripts after the atomic symbol.
   - The numbers of atoms are given in the molecule's name in Greek prefixes
   - NOTE: If there is no Greek prefix in front of the first element in the name, that means the number is 1.

No

Ionic Compound

1) Write the symbol/formula of the first ion in the compound's name, then the symbol/formula of the second ion in the compound's name.
2) Indicate the ratio of each ion in the compound using subscripts after each ion.
   - The ratio of the ions are deduced by balancing the charges of the ions.

IMPORTANT: When there is more than one of a polyatomic ion in the formula unit we use parenthesis. Example $Mg(NO_3)_2$

Writing the Formulas of Ionic Compounds

Example: Write the formula for **calcium bromide**.

1) Write the symbol/formula of the first ion in the compound's name, then the symbol/formula of the second ion in the compound's name.

   $\mathrm{Ca\ \ Br}$

2) Indicate the ratio of each ion in the compound using subscripts after each ion.
   - This step involves filling in the subscripts boxes as we did in the lecture:

     $\mathrm{Ca}_{\square}\,\mathrm{Br}_{\square}$

   - The ratio of the ions is deduced by **balancing the charges** of the ions.
     - This is done so that the **total charge** in the crystal, when large numbers of cations and anions combine, is **equal to zero**.
     - We find the ion's charge from its position on the periodic table <u>**or**</u> we look it up in a table in the case of polyatomic ions.
     - Transition metal with varying charges will be written in the compound name in Roman numerals.

   - First, temporarily write the charge of each ion above the ion's symbol.

     $\overset{2+}{\mathrm{Ca}}_{\square}\,\overset{1-}{\mathrm{Br}}_{\square}$

   - Next, place numbers in the subscripts such that the total charge of the compound is zero. Note that in this example, we need **two** bromide ions, each has a charge of (1-) to cancel the (2+) charge of the calcium ion:
     - 2(-1) + (+2) = 0 zero total charge.

     $$\overset{2+}{\mathrm{Ca}}_{\boxed{1}}\,\overset{1-}{\mathrm{Br}}_{\boxed{2}} \quad = \quad \mathrm{CaBr_2}$$

     - We saw a shortcut way to do this called the Criss-Cross Method (see your chapter 3 notes)

     $$\overset{2+}{\mathrm{Ca}}_{\boxed{1}}\,\overset{1-}{\mathrm{Br}}_{\boxed{2}} \quad = \quad \mathrm{CaBr_2}$$

     - Note, we do not leave the charges written above the symbols in the completed formula.

IMPORTANT: When there is more than one of a polyatomic ion in the formula, we use parenthesis.

- Not applicable in this example since there are no polyatomic ions in calcium bromide.

Examples: Writing the Formulas of Ionic Compounds

Write the formula for **magnesium nitrate**.

1) Write the symbol/formula of the first ion in the compound's name, then the symbol/formula of the second ion in the compound's name.
    - When you see a polyatomic ion (nitrate), look up the formula and charge in the table of polyatomic ions.

$$Mg \;\; NO_3$$

2) Indicate the ratio of each ion in the compound using subscripts after each ion.
    a. This step involves filling in the subscripts boxes as we did in the lecture:

$$Mg_{\square}NO_{3\square}$$

- The ratio of the ions is deduced by **balancing the charges** of the ions.
    - This is done so that the **total charge** in the crystal, when large numbers of cations and anions combine, is **equal to zero**.
    - We find the ion's charge from its position on the periodic table <u>**or**</u> we look it up in a table in the case of polyatomic ions.
    - Transition metal with varying charges will be written in the compound name in Roman numerals.

- First, temporarily write the charge of each ion above the ion's symbol.

$$\overset{2+}{Mg}_{\square}\overset{1-}{NO_3}_{\square}$$

- Next, place numbers in the subscripts such that the total charge of the compound is zero. Note that in this example, we need **two** nitrate ions, each has a charge of (1-) to cancel the (2+) charge of the magnesium ion:
    - $2(-1) + (+2) = 0$ zero total charge.

$$\overset{2+}{Mg}_{\boxed{1}}\;\overset{1-}{NO_3}_{\boxed{2}}$$

- We saw a shortcut way to do this called the Criss-Cross Method (see your chapter 3 notes)

$$\overset{2+}{Mg}_{\boxed{1}}\;\overset{1-}{NO_3}_{\boxed{2}}$$

*IMPORTANT: When there is more than one of a polyatomic ion in the formula unit we use parenthesis. There are **2 ions** of nitrate in magnesium nitrate*

$$Mg_{\boxed{1}}\;NO_{3\boxed{2}} \;=\; Mg(NO_3)_2$$

In compound where there is just **one formula unit** of a polyatomic ion, no parenthesis are needed. An example of this is **sodium nitrate: $NaNO_3$**

Examples: Writing the Formulas of Ionic Compounds

Write the formula for **iron(II) phosphate**.

1) Write the symbol/formula of the first ion in the compound's name, then the symbol/formula of the second ion in the compound's name.
    - When you see a polyatomic ion (nitrate), look up the formula and charge in the table of polyatomic ions.

$$Fe \;\; PO_4$$

2) Indicate the ratio of each ion in the compound using subscripts after each ion.
    b. This step involves filling in the subscripts boxes as we did in the lecture:

$$Fe_{\square} \; PO_{4\square}$$

- The ratio of the ions is deduced by **balancing the charges** of the ions.
    - This is done so that the **total charge** in the crystal, when large numbers of cations and anions combine, is **equal to zero**.
    - We find the ion's charge from its position on the periodic table <u>**or**</u> we look it up in a table in the case of polyatomic ions.
    - **Transition metal with varying charges will be written in the compound name in Roman numerals.**
        - In this example, now we know the charge on **the Fe ion is 2+**

- First, temporarily write the charge of each ion above the ion's symbol.

$$\overset{2+}{Fe}_{\square} \; \overset{3-}{PO}_{4\square}$$

- Next, place numbers in the subscripts such that the total charge of the compound is zero. Note that in this example, we need **two** phosphate ions, each has a charge of (3-) and three Fe^{2+} ions to balance the charge:
    - $2(-3) \; + \; 3(-2) = 0$ zero total charge.

$$\overset{2+}{Fe}_{\boxed{3}} \; \overset{3-}{PO}_{4\boxed{2}}$$

- We saw a shortcut way to do this called the Criss-Cross Method (see your chapter 3 notes)

$$\overset{2+}{Fe}_{\boxed{3}} \; \overset{3-}{PO}_{4\boxed{2}}$$

IMPORTANT: When there is more than one of a polyatomic ion in the formula unit we use parenthesis. There are ***2 ions*** *of phosphate in iron(II)phosphate.*

$$Fe_{\boxed{3}} \; PO_{4\boxed{2}} \; = \; Fe_3(PO_4)_2$$

Examples: Writing the Formulas of Ionic Compounds

Write the formula for **barium sulfide**.

1) Write the symbol/formula of the first ion in the compound's name, then the symbol/formula of the second ion in the compound's name.

   Ba S

2) Indicate the ration of each ion in the compound using subscripts after each ion.
   - This step involves filling in the subscripts boxes as we did in the lecture:

     $Ba_{\boxed{}}S_{\boxed{}}$

   - The ratio of the ions is deduced by **balancing the charges** of the ions.
     - This is done so that the **total charge** in the crystal, when large numbers of cations and anions combine, is **equal to zero.**
     - We find the ion's charge from its position on the periodic table <u>**or**</u> we look it up in a table in the case of polyatomic ions.
     - Transition metal with varying charges will be written in the compound name in Roman numerals.

   - First, temporarily write the charge of each ion above the ion's symbol.

     $\overset{2+}{Ba}_{\boxed{}}\overset{2-}{S}_{\boxed{}}$

   - Next, place numbers in the subscripts such that the total charge of the compound is zero. Note that in this example, we need **one** sulfide ion, with a charge of (2-) to cancel the (2+) charge of the barium ion:
     - (-2) + (+2) = 0 zero total charge.

     $\overset{2+}{Ba}_{\boxed{1}}\overset{2-}{S}_{\boxed{1}} = BaS$

     - We saw a shortcut way to do this called the Criss-Cross Method (see your chapter 3 notes)

       $\overset{2+}{Ba}_{\boxed{2}}\overset{2-}{S}_{\boxed{2}} = BaS$

     - Note, the subscripts in ionic compound represent the ratio in which large numbers of anions and cations combine to form the ionic compounds. Since we want the **lowest ratio**: we use 1:1, since 2:2 = 1:1

       $\cancel{Ba_2S_2} = BaS$

Write the formula for the following ionic compounds: (see next page for key)

sodium bicarbonate ________________

sodium fluoride ________________

iron (III) chloride ________________

sodium carbonate ________________

copper (II) sulfate ________________

magnesium hydroxide ________________

barium nitrate ________________

lithium sulfate ________________

magnesium chloride ________________

silver nitrate ________________

aluminum sulfate ________________

calcium hydroxide ________________

calcium sulfate ________________

mercury (II) nitrate ________________

lead (IV) nitrate ________________

magnesium iodide ________________

sodium nitride ________________

Practice Problems KEY

sodium bicarbonate $NaHCO_3$

sodium fluoride NaF

iron (III) chloride $FeCl_3$

sodium carbonate Na_2CO_3

copper (II) sulfate $CuSO_4$

magnesium hydroxide $Mg(OH)_2$

barium nitrate $Ba(NO_3)_2$

lithium sulfate Li_2SO_4

magnesium chloride $MgCl_2$

silver nitrate $AgNO_3$

aluminum sulfate $Al_2(SO_4)_3$

calcium hydroxide $Ca(OH)_2$

calcium sulfate $CaSO_4$

mercury (II) nitrate $Hg(NO_3)_2$

lead (IV) nitrate $Pb(NO_3)_4$

magnesium iodide MgI_2

sodium nitride Na_3N

Writing the Formulas of Covalent Compounds

1) Write the symbol/formula of the first element in the compound's name, then the symbol/formula of the second ion in the compound's name.
2) Indicate how many of each element the molecule contains using subscripts after the atomic symbol.
    - The numbers of atoms are given in the molecule's name in Greek prefixes
    - NOTE: If there is no Greek prefix in front of the first element in the name, that means the number is 1.

Example: Write the formula of **dinitrogen tetrafluoride**.

1) Write the symbol/formula of the first element in the compound's name, then the symbol/formula of the second ion in the compound's name.

$$N \ F$$

2) Indicate how many of each element the molecule contains using subscripts after the atomic symbol.

$$N_{\square}F_{\square}$$

- The numbers of atoms are given in the molecule's name in Greek prefixes.
    - **di**nitrogen **tetra**fluoride
    - see your chapter 3 notes for a list of the Greek prefixes

$$N_2F_4$$

- **NOTE**: If there is no Greek prefix in front of the first element in the name, then the number is 1.
    - Example carbon tetrachloride = CCl_4

Example: Write the formula of **carbon disulfide**.

1) Write the symbol/formula of the first element in the compound's name, then the symbol/formula of the second ion in the compound's name.

$$C \ S$$

2) Indicate how many of each element the molecule contains using subscripts after the atomic symbol.

$$C_{\square}S_{\square}$$

- The numbers of atoms are given in the molecule's name in Greek prefixes.
    - carbon **di**sulfide
    - see your chapter 3 notes for a list of the Greek prefixes

$$C_1S_2 \quad = CS_2$$

- **NOTE**: If there is no Greek prefix in front of the first element in the name, then the number is 1.

Write the formulas for the following covalent compounds:

See next page for KEY

a. disulfur tetrafluoride ____________________

b. carbon trioxide ____________________

c. nitrogen pentoxide ____________________

d. nitrogen tribromide ____________________

e. dinitrogen heptachloride ____________________

f. carbon tetrachloride____________________

g. hydrogen monochloride____________________

h. trihydrogen monophosphide ____________________

i. dihydrogen monoxide ____________________

KEY

a. disulfur tetrafluoride S_2F_4

b. carbon trioxide CO_3

c. nitrogen pentoxide NO_5

d. nitrogen tribromide NBr_3

e. dinitrogen heptachloride N_2Cl_7

f. carbon tetrachloride CCl_4

g. hydrogen monochloride HCl

h. trihydrogen monophosphide H_3P

i. dihydrogen monoxide H_2O

Review of Writing Formulas for Compounds

Given the **Formula**, Writing the **Name:**

Determine if the compound is ionic or covalent (molecular)

Is it a non-metal with a non-metal ?

Yes

No

Covalent (molecular) Compound

1) List the name of the first element in the formula.
2) List the second element and add the –ide suffix.
3) Use Greek prefixes to indicate the number of each atom in the formula.
   - •Exception: do not use mono- for the first element in the name.
   - •The *o* or *a* at the end of the Greek prefix is usually dropped when the element name begins with a vowel

Ionic Compound

1) Write the cation name first, then name the anion.
   - monoatomic anions use the "ide" suffix

2) If the cation is one of the transition metals with various charges, write the charge using parenthesis and Roman numerals after the metal name.

IMPORTANT: When there is more than one of a polyatomic ion in the formula unit we use parenthesis. Example $Mg(NO_3)_2$

Writing the Names of Ionic Compounds

Example: Write the name for **$CaBr_2$**

1) Write the cation name first, then name the anion.
    - monoatomic anions use the "ide" suffix

calcium bromide

2) If the cation is one of the transition metals with various charges, write the charge using parenthesis and Roman numerals after the metal name.
    - Not necessary here, there is not a transition metal present

Example: Write the name for **$Mg(NO_3)_2$**

1) Write the cation name first, then name the anion.
    - monoatomic anions use the "ide" suffix
        - Here we notice that the anion is a **polyatomic ion**. Get the name from the polyatomic ion table (in your notes or textbook). *You will be given a copy of the polyatomic ion table on your exams.*
        - **Do not** change the suffix to "ide" with polyatomic ions:

magnesium nitrate

2) If the cation is one of the transition metals with various charges, write the charge using parenthesis and Roman numerals after the metal name.
    - Not necessary here, there is not a transition metal present

Writing the Names of Ionic Compounds

Example: Write the name for **CuF_2**

1) Write the cation name first, then name the anion.
   - monoatomic anions use the "ide" suffix

copper fluoride

2) If the cation is one of the ***transition metals*** with various charges, write the **charge using parenthesis and Roman numerals** after the metal name.

copper(?) fluoride

- We must figure out what the charge is on the copper, we can deduce the charge on the transition metal cations from the charge on the anions
  - Recall that the total charge for any compound must equal zero.
  - Since there are two bromides, each with a charge of (1-) and there is only one copper, we can conclude that the charge on the copper must be (2+).
    - You can think of this as the reverse-criss-cross! See chapter 3 notes for more details.

copper(II) fluoride

write the **charge in parenthesis in Roman numerals after the cation name**

Write the names of the following compounds:

See next page for key

$NaCl$ ______________________________

$Fe_2(CO_3)_3$ ______________________________

$Cu(OH)_2$ ______________________________

$(NH_4)_2SO_4$ ______________________________

$LiNO_3$ ______________________________

$BaSO_4$ ______________________________

$Mg(NO_3)_2$ ______________________________

$AgCl$ ______________________________

$Al(OH)_3$ ______________________________

$CaSO_4$ ______________________________

FeS ______________________________

$PbCl_2$ ______________________________

NaI ______________________________

$MgCO_3$ ______________________________

KEY

NaCl sodium chloride

$Fe_2(CO_3)_3$ iron(III) carbonate

$Cu(OH)_2$ copper(II) hydroxide

$(NH_4)_2SO_4$ ammonium sulfate

$LiNO_3$ lithium nitrate

$BaSO_4$ barium sulfate

$Mg(NO_3)_2$ magnesium nitrate

AgCl silver chloride

- (note: silver is one of the transition metals that only occurs as a (1+) ion)

$Al(OH)_3$ aluminum hydroxide

$CaSO_4$ calcium sulfate

FeS Iron(II) sulfide

$PbCl_2$ lead(II) chloride

NaI sodium iodide

$MgCO_3$ magnesium carbonate

Writing the Names of Covalent Compounds

1) List the name of the first element in the formula.

2) List the second element and add the –ide suffix.

3) Use Greek prefixes to indicate the number of each atom in the formula.
- Exception: do not use mono- for the first element in the name.
- The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel

Example: Write the name for $\mathbf{N_2S_4}$

1) List the name of the first element in the formula.

nitrogen

2) List the second element and add the –ide suffix.

nitrogen sulfide

3) Use Greek prefixes to indicate the number of each atom in the formula.
- See your textbook or lecture notes for a table of the Greek prefixes.

____nitrogen _____sulfide

dinitrogen tetrasulfide

- Exception: do not use mono- for the first element in the name.
  - Not applicable in this example
- The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel
  - Not applicable in this example

Example: Write the name for $\mathbf{SO_3}$

1) List the name of the first element in the formula.

sulfur

2) List the second element and add the –ide suffix.

sulfur oxide

3) Use Greek prefixes to indicate the number of each atom in the formula.

____ sulfur _____ oxide

sulfur trioxide

- Exception: do not use **mono-** for the ***first*** element in the name.
  - NOTE, we did not write **mono**sulfur because of this rule!
- The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel
  - Not applicable in this example

Example: Write the name for $\mathbf{SO_3}$

1) List the name of the first element in the formula.

sulfur

2) List the second element and add the –ide suffix.

sulfur oxide

3) Use Greek prefixes to indicate the number of each atom in the formula.

____ sulfur _____ oxide

sulfur trioxide

- Exception: do not use **mono-** for the ***first*** element in the name.
    - NOTE, we did not write **mono**sulfur because of this rule!
- The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel
    - Not applicable in this example

Example: Write the name for **CO**

1) List the name of the first element in the formula.

carbon

2) List the second element and add the –ide suffix.

carbon oxide

3) Use Greek prefixes to indicate the number of each atom in the formula.

____ carbon _____ oxide

carbon monoxide

- Exception: do not use **mono-** for the ***first*** element in the name.
    - NOTE, we did not write **monocarbon** because of this rule!
- The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel
    - NOTE, we did not write **mon<u>o</u>oxygen** because of this rule!

Write the names of the following compounds:

See next page for key

a. Br_2I_4 ______________________________

b. P_5F_8 ______________________________

c. NO_5 ______________________________

- Remember: The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel

d. NBr_3 ______________________________

e. N_2O_5 ______________________________

f. $BrCl_3$ ______________________________

g. H_2S ______________________________

h. N_2O ______________________________

KEY

a. Br_2I_4 dibromine tetriodide

b. P_5F_8 pentaphosphorus octafluoride

c. NO_5 nitrogen pentoxide
   - The *o* or *a* at the end of the Greek pre-fix is usually dropped when the element name begins with a vowel
     - NOTE, we did not write **pent<u>ao</u>xygen** because of this rule!

d. NBr_3 nitrogen tribromide

e. N_2O_5 dinitrogen pentoxide

f. $BrCl_3$ bromine trichloride

g. H_2S dihydrogen monosulfide

h. N_2O dinitrogen monoxide

EMPIRICAL AND MOLECULAR FORMULAS

Example: EMPIRICAL FORMULA:

Suppose a compound is analyzed to contain 48.8 g of cadmium, 20.8 g of carbon, 2.62 g of hygrogen, and 27.8 g of oxygen. Determine the empirical formula of this compound.

To determine the empirical formula of a compound for which the amounts of each element are given you convert the amounts to moles using the elements atomic weights from the periodic table as follows:

$48.8\text{ g Cd} \left| \frac{1\text{ mole Cd}}{112.4\text{ g Cd}} \right. = 0.434\text{ moles Cd}$

$20.8\text{ g C} \left| \frac{1\text{ mole C}}{12.01\text{ g C}} \right. = 1.67\text{ moles C}$

$2.62\text{ g H} \left| \frac{1\text{ mole H}}{1.01\text{ g H}} \right. = 2.62\text{ moles H}$

$27.8\text{ g O} \left| \frac{1\text{ mole O}}{16\text{ g O}} \right. = 1.74\text{ moles O}$

This gives you the mole ratio of the elements in the formula and theoretically the formula for the compound could be written as:

$$Cd_{0.434}C_{1.67}H_{2.62}O_{1.74}$$

But we know that atoms are not fractional in nature so we must find the smallest whole number ratio for the elements making up the compound. To do this we divide each of the numbers by the smallest number in the set. In our case the smallest number is 0.434 moles. *

$48.8\text{ g Cd} \left| \frac{1\text{ mole Cd}}{112.4\text{ g Cd}} \right. = 0.434\text{ moles Cd}/0.434\text{ moles} = 1$

$20.8\text{ g C} \left| \frac{1\text{ mole C}}{12.01\text{ g C}} \right. = 1.73\text{ moles C}/0.434\text{ moles} = 4$

$2.62\text{ g H} \left| \frac{1\text{ mole H}}{1.01\text{ g H}} \right. = 2.62\text{ moles H}/0.434\text{ moles} = 6$

$27.8\text{ g O} \left| \frac{1\text{ mole O}}{16\text{ g O}} \right. = 1.74\text{ moles O}/0.434\text{ moles} = 4$

* *After division if this step produces any number ending in .5 multiply all numbers by 2 to obtain small whole numbers*

* *After division if this step produces any number ending in .33 multiply all numbers by 3 to obtain small whole numbers*

These whole numbers become the subscripts for the empirical formula: $CdC_4H_6O_4$. You don't write the (1) for Cd---the symbol already indicates (1). From experience it is probably safe to say the formula is: $Cd(C_2H_3O_2)_2$. Cadmium Acetate. This is the Empirical or Simple Formula for the compound. It is just the ratio of the amounts of each type of element in the compound.

EXAMPLE: MOLECULAR FORMULA

In order to determine the Actual or Molecular Formula, you need to know the molecular mass of the compound.

Suppose a compound whose molecular mass is 695 is analyzed to contain 26.7% phosphorus, 12.1% nitrogen, and 61.2% chlorine.

You first have to find the Empirical Formula the same way we did problem#1. We can state the % given to us for each element as grams if we base our calcualtions on 100 g of compound. Steps have been combined to give the following

$$\frac{26.7 \text{ g P}}{} \left| \frac{1 \text{ mole P}}{30.97 \text{ g P}} \right. = 0.862 \text{ moles P}/0.862\text{moles} = 1$$

$$\frac{12.1 \text{ g N}}{} \left| \frac{1 \text{ mole N}}{14 \text{ g N}} \right. = 0.864 \text{ moles N}/0.862 \text{ moles} = 1$$

$$\frac{61.2 \text{ g Cl}}{} \left| \frac{1 \text{ mole Cl}}{35.45 \text{ g Cl}} \right. = 1.726 \text{ mols Cl}/0.862/\text{moles} = 2$$

Therefore the Empirical formual is $PNCl_2$

To determine the molecular formula you set the EF mass taken X times and set it equal to the Molecular Weight of the compound. Like this:

$$(PNCl_2)\ X = 695$$

Now substitute the atomic weight of the elements into the equation to get this:

$$[30.97 + 14 + 2(35.45)]\ X = 695$$

$$[115.87]\ X = 695$$

$$X = 695/115.87$$

$$X = 6$$

CHEMISTRY **E.F. AND M. F. Worksheet**

To get the Molecular Formula we must multiply each of the subscripts in the Empirical Formual by 6 to give us this:

$$P_6N_6Cl_{12}$$

As a check, you can find the formula mass as usual from the sum of the atomic weights:

$P = 6 \times 30.97$	$= 185.82$
$N = 6 \times 14$	$= 84$
$Cl = 12 \times 35.45$	$= 425.4$
Total	$= 695.22$

You can see from this example that the MF is a multiple of the EF and represents the actual numbe of each type of atom that makes up the compound.

PRACTICE PROBLEMS:

1. What is the empirical formula of a substance composed 56.6% potassium, 8.68% carbon, and 34.7% oxygen.

2. What is the empirical formula of a compound composed of 3.26 gof arsenic and 1.04 g of oxygen?

3. An unknown compound is analyzed and found to consist of 24.3 % carbon, 4.1 % hydrogen, and 71.6 % chlorine. If the molecular mass of the compound is 98.8, what is the molecular formula of the compound?

4. What is the empirical formula of a substance composed 49.89% strontium, 13,67% carbon, and 36.44% oxygen?

5. What is the empirical formula of a compound composed of 10.12 g of aluminum and 17.93 g of sulfur?

6. An unknown compound is analyzed and found to consist of 49.0 % carbon, 2.7 % hydrogen, and 48.2 % chlorine. If the molecular mass of the compound is 150, what is the molecular formula of the compound?

7. Find the molecular formula for a compound with percentage composition 85.6 % C, 14.4 % H, and molecular mass 42.1.

8. What is the molecular formula of a substance with empirical formula $TlC_2H_2O_3$ and molecular mass 557?

9. Hydroquinone is an organic compound commonly used as a photographic developer. It has a molecular weight of 110 g/mole and a composition of 65.45% C, 5.45 % H, and 29.09 % O. Calculate the molecular formula of hydroquinone.

Empirical and Molecular Formula Worksheet

SHOW WORK ON A SEPARATE SHEET OF PAPER.

Write the empirical formula for the following compounds.

1) C_6H_6
2) C_8H_{18}
3) WO_2
4) $C_2H_6O_2$
5) $X_{39}Y_{13}$

6) A compound with an empirical formula of C_2OH_4 and a molar mass of 88 grams per mole. What is the molecular formula of this compound?

7) A compound with an empirical formula of C_4H_4O and a molar mass of 136 grams per mole. What is the molecular formula of this compound?

8) A compound with an empirical formula of CFBrO and a molar mass of 254.7 grams per mole. What is the molecular formula of this compound?

9) A compound with an empirical formula of C_2H_8N and a molar mass of 46 grams per mole. What is the molecular formula of this compound?

10) A well-known reagent in analytical chemistry, dimethylglyoxime, has the empirical formula C_2H_4NO. If its molar mass is 116.1 g/mol, what is the molecular formula of the compound?

12. Nitrogen and oxygen form an extensive series of oxides with the general formula N_xO_y. One of them is a blue solid that comes apart, reversibly, in the gas phase. It contains 36.84% N. What is the empirical formula of this oxide?

13. A sample of indium chloride weighing 0.5000 g is found to contain 0.2404 g of chlorine. What is the empirical formula of the indium compound?

14. An unknown compound was found to have a percent composition as follows: 47.0 % potassium, 14.5 % carbon, and 38.5 % oxygen. What is its empirical formula? If the true molar mass of the compound is 166.22 g/mol, what is its molecular formula?

15. Rubbing alcohol was found to contain 60.0 % carbon, 13.4 % hydrogen, and the remaining mass was due to oxygen. What is the empirical formula of rubbing alcohol?

Empirical and Molecular Formula Worksheet ANSWER KEY

Write the empirical formula for the following compounds.

1) C_6H_6 **C_3H_3**
6) C_8H_{18} **C_4H_9**
7) WO_2 **WO_2**
8) $C_2H_6O_2$ **CH_3O**
9) $X_{39}Y_{13}$ **X_3Y**

6) A compound with an empirical formula of C_2OH_4 and a molar mass of 88 grams per mole. What is the molecular formula of this compound? **$C_4O_2H_8$**

7) A compound with an empirical formula of C_4H_4O and a molar mass of 136 grams per mole. What is the molecular formula of this compound? **$C_8H_8O_2$**

8) A compound with an empirical formula of CFBrO and a molar mass of 254.7 grams per mole. What is the molecular formula of this compound? **$C_2F_2Br_2O_2$**

9) A compound with an empirical formula of C_2H_8N and a molar mass of 46 grams per mole. What is the molecular formula of this compound? **C_2H_8N**

10) A well-known reagent in analytical chemistry, dimethylglyoxime, has the empirical formula C_2H_4NO. If its molar mass is 116.1 g/mol, what is the molecular formula of the compound? **$C_4H_8N_2O_2$**

12. A certain blue solid contains 36.84% N. What is the empirical formula of this compound? **The ratios are $N_{1.00}O_{1.50}$. Since 1.50 is not close to a whole number, we multiply *both* subscripts by 2. The empirical formula is thus N_2O_3. (The name is dinitrogen trioxide.)**

13. A sample of indium chloride weighing 0.5000 g is found to contain 0.2404 g of chlorine. What is the empirical formula of the indium compound? **$InCl_3$**

14. An unknown compound was found to have a percent composition as follows: 47.0 % potassium, 14.5 % carbon, and 38.5 % oxygen. What is its empirical formula? If the true molar mass of the compound is 166.22 g/mol, what is its molecular formula? **$K_2C_2O_4$**

15. Rubbing alcohol was found to contain 60.0 % carbon, 13.4 % hydrogen, and the remaining mass was due to oxygen. What is the empirical formula of rubbing alcohol? **C_3H_8O**

PERIODIC TABLE OF THE ELEMENTS

(based on Carbon-12 = 12.0000)

Period	1/ 1A	2/ 2A	3/ 3B	4/ 4B	5/ 5B	6/ 6B	7/ 7B	8/ 8B	9/ 8B	10/ 8B	11/ 1B	12/ 2B	13/ 3A	14/ 4A	15/ 5A	16/ 6A	17/ 7A	18/ 8A
1	**1 H** Hydrogen 1.01																	**2 He** Helium 4.00
2	**3 Li** Lithium 6.94	**4 Be** Beryllium 9.01											**5 B** Boron 10.81	**6 C** Carbon 12.01	**7 N** Nitrogen 14.01	**8 O** Oxygen 16.00	**9 F** Fluorine 19.00	**10 Ne** Neon 20.18
3	**11 Na** Sodium 22.99	**12 Mg** Magnes-ium 24.30											**13 Al** Aluminum 26.98	**14 Si** Silicon 28.09	**15 P** Phospho-rus 30.97	**16 S** Sulfur 32.07	**17 Cl** Chlorine 35.45	**18 Ar** Argon 39.95
4	**19 K** Potassium 39.10	**20 Ca** Calcium 40.08	**21 Sc** Scandium 44.96	**22 Ti** Titanium 47.88	**23 V** Vanadium 50.94	**24 Cr** Chromium 52.00	**25 Mn** Mangan-ese 54.94	**26 Fe** Iron 55.85	**27 Co** Cobalt 58.93	**28 Ni** Nickel 58.69	**29 Cu** Copper 63.55	**30 Zn** Zinc 65.39	**31 Ga** Gallium 69.72	**32 Ge** German-ium 72.61	**33 As** Arsenic 74.92	**34 Se** Selenium 78.96	**35 Br** Bromine 79.90	**36 Kr** Krypton 83.80
5	**37 Rb** Rubidium 85.47	**38 Sr** Strontium 87.62	**39 Y** Yttrium 88.91	**40 Zr** Zirconium 91.22	**41 Nb** Niobium 92.91	**42 Mo** Molyb-denum 95.94	**43 Tc** Technet-ium (98)	**44 Ru** Ruthen-ium 101.07	**45 Rh** Rhodium 102.91	**46 Pd** Palladium 106.42	**47 Ag** Silver 107.87	**48 Cd** Cadmium 112.41	**49 In** Indium 114.82	**50 Sn** Tin 118.71	**51 Sb** Antimony 121.76	**52 Te** Tellurium 127.60	**53 I** Iodine 126.90	**54 Xe** Xenon 131.29
6	**55 Cs** Cesium 132.90	**56 Ba** Barium 137.33	**71 Lu** Lutetium 174.97	**72 Hf** Hafnium 178.49	**73 Ta** Tantalum 180.95	**74 W** Tungsten 183.85	**75 Re** Rhenium 186.21	**76 Os** Osmium 190.2	**77 Ir** Iridium 192.22	**78 Pt** Platinum 195.08	**79 Au** Gold 196.97	**80 Hg** Mercury 200.59	**81 Tl** Thallium 204.38	**82 Pb** Lead 207.2	**83 Bi** Bismuth 208.98	**84 Po** Polonium (209)	**85 At** Astatine (210)	**86 Rn** Radon (222)
7	**87 Fr** Francium (223)	**88 Ra** Radium (226)	**103 Lr** Lawrenc-ium (262)	**104 Rf** Ruther-fordium (261)	**105 Db** Dubnium (262)	**106 Sg** Seaborg-ium (266)	**107 Bh** Bohrium (264)	**108 Hs** Hassium (270)	**109 Mt** Meitner-ium (276)	**110 Ds** Darmstad-tium (281)	**111 Rg** Roentgen-ium (280)	**112 Cn** Copernic-ium (285)	**113 Uut** # (284)	**114 Uuq** # (285)	**115 Uup** # (288)	**116 Uuh** # (292)	**117 Uus** # (293)	**118 Uuo** # (294)

Lanthanide Series	**57 La** Lantha-num 138.91	**58 Ce** Cerium 140.12	**59 Pr** Praseod-ymium 140.91	**60 Nd** Neodym-ium 144.24	**61 Pm** Prometh-ium (145)	**62 Sm** Samarium 150.36	**63 Eu** Europium 151.96	**64 Gd** Gadolin-ium 157.25	**65 Tb** Terbium 158.92	**66 Dy** Dyspros-ium 162.50	**67 Ho** Holmium 164.93	**68 Er** Erbium 167.26	**69 Tm** Thulium 168.93	**70 Yb** Ytterbium 173.04
Actinide Series	**89 Ac** Actinium (227)	**90 Th** Thorium 232.04	**91 Pa** Protactin-ium 231.04	**92 U** Uranium 238.03	**93 Np** Neptun-ium (237)	**94 Pu** Plutonium (244)	**95 Am** Americ-ium (243)	**96 Cm** Curium (247)	**97 Bk** Berkelium (247)	**98 Cf** Californ-ium (251)	**99 Es** Einstein-ium (252)	**100 Fm** Fermium (257)	**101 Md** Mendelv-ium (258)	**102 No** Nobelium (259)

Revised 4/8/2010

Notes; *proposed name, #has not been confirmed