Chem 115 POGIL Worksheet - Week 3
Compounds, Naming, Reaction Equations, and Formula Weights

Why?

Compounds are generally classified as molecular, ionic, or (more rarely) network. Knowing the classification allows us to name the compound correctly and to understand the microscopic organization of it. Describing the fundamental compound unit as either a molecule or a formula unit allows us to determine the mass of that unit. Knowing these fundamental molecular or ionic unit masses allows us to predict mass changes that occur on the macroscopic scale as a result of chemical reactions.

Learning Objective

- Understand the difference between molecular and ionic compounds
- Understand the distinction between molecular formula and empirical formula
- Know the I.U.P.A.C. rules for naming simple inorganic compounds
- Understand the meaning of a balanced chemical equation
- Understand the relationships between formula weight, molecular weight, and percent composition

Success Criteria

- Be able to predict whether a compound is molecular or ionic from the positions of its elements in the periodic table
- Be able to predict the empirical formulas of simple ionic compounds
- Be able to name simple compounds on the basis of their being molecular or ionic
- Be able to balance simple reaction equations by inspection

Prerequisite

- Have read sections 2.6 – 2.9 and 3.1 – 3.5 in your text

Information

Molecules are combinations of atoms tightly bound together to form a chemically identifiable unit. Many elements and compounds — but not all — are composed of molecules. Molecules of elements are **homonuclear**, because they are composed of only one kind of atom. Molecules of compounds are **heteronuclear**, because they are composed of two or more different kinds of atoms. For a molecular substance (element or compound), the composition of the molecules is indicated by a **molecular formula**, which shows the kinds and numbers of each atom in the molecule. The following common elements are composed of molecules with the compositions indicated by their molecular formulas: $H_2(g)$, $F_2(g)$, $Cl_2(g)$, $Br_2(l)$, $I_2(s)$, $O_2(g)$, $S_8(s)$, $N_2(g)$, $P_4(s)$. The notations *g*, *l*, *s* stand for gas, liquid, and solid, respectively, and represent the states of these elements at room temperature. *Know the molecular formulas and room-temperature states of these common elements.* Not all compounds contain discrete, identifiable molecules, but for those that do a molecular formula indicates the actual numbers and kinds of atoms comprising the molecular unit. When formulas of compounds are determined by chemical analysis, the information obtained usually gives only the simplest whole number ratios among the elements, which is expressed as an **empirical formula**. The following examples show the relationship between the molecular formula and empirical formula for some compounds.

Molecular Formula	Empirical Formula
H_2O	H_2O
H_2O_2	HO
CH_4	CH_4
C_2H_4	CH_2
C_6H_{12}	CH_2
NO_2	NO_2
N_2O_4	NO_2

Note that sometimes the molecular and empirical formulas are the same, and sometimes they are different. But in every case, the molecular formula is a whole number multiple of the empirical formula. Note, too, that different compounds, composed of different kinds of molecules, may have the same empirical formula.

Ionic compounds are made up of electrically equivalent numbers of cations and anions. Ionic compounds are typically crystalline solids with a structure consisting of an orderly three-dimensional array of ions called a **crystal lattice**. The composition of an ionic compound can be represented by an empirical formula, but there is no corresponding molecular formula because there are no molecules of the compound. It is sometimes convenient to refer to the **formula unit** of an ionic compound, a minimum collection of ions that corresponds to the empirical formula. However, the formula unit has no chemical existence in the way that a molecule does.

Compounds consisting of nonmetal elements are usually molecular; e.g., H_2O, C_2H_6, N_2O_4, NH_3. Binary (two-element) compounds formed between metals and nonmetals are usually ionic; e.g., $NaCl$, CaF_2, Al_2O_3, Na_3N. The following generalizations help determine the charges on the ions in ionic compounds.

1. Metals form cations; nonmetals form anions.

2. Main-group metals tend to form cations with charges equal to the last digit of their group number (IUPAC convention)).

   Examples: Na^+ (Group 1), Mg^{2+} (Group 2), Al^{3+} (Group 13)

3. Nonmetals tend to form anions with charges equal to their group number (IUPAC convention) minus 18.

   Examples: F^- (Group 17, 17 - 18 = -1), O^{2-} (Group 16, 16 - 18 = -2), N^{3-} (Group 15, 15 - 18 = -3)

4. Transition metals and some heavier main group elements can form more than one kind of cation.

   Examples: Cu^{+} & Cu^{2+}; Fe^{2+} & Fe^{3+}; Co^{2+} & Co^{3+}; Cr^{2+} & Cr^{3+}; Tl^{+} & Tl^{3+}

5. Ionic charges greater than ±3 are not real. Compounds in which an element might be assigned such high charge are probably molecular (or less commonly, network solids).

Some ionic compounds are formed from **polyatomic ions**, which are molecular ions with specific charges. Either the cation or anion or both can be polyatomic ions. For example, Na_2SO_4 consists of two monatomic sodium ions for every one sulfate ion, which has the formula and charge SO_4^{2-}. *You need to memorize the names, formulas, and charges of the common ions shown in the table at the end of this work sheet.* There is no easy way to figure these out.

If you know the charges on the ions comprising an ionic compound you can predict the empirical formula by taking the smallest number of cations and anions that would add to a charge of zero. For example, the ionic compound formed between Al^{3+} cations and SO_4^{2-} anions would have the empirical formula $Al_2(SO_4)_3$, because (2)(3+) + (3)(2–) = 0. Note that the sulfate ion is shown in parentheses to indicate that three whole SO_4^{2-} polyatomic anions are present for every two Al^{3+} cations. Also note that the charges on the ions are not included in the empirical formula $Al_2(SO_4)_3$.

Network solids are elements or compounds in which all the atoms are bound together in a limitless three-dimensional structure. The two forms of carbon, diamond and graphite, are elemental examples of network solids, each with its own structure. Quartz (silicon dioxide) is also a network solid. Because there are no discrete molecules, the formula for a network solid can only be represented as an empirical formula. Thus, graphite is represented as C(*s*), and quartz is represented as $SiO_2(s)$. Distinct from ionic compounds, network solids do not contain ions.

Key Questions

1. Do all compounds contain molecules?

2. What is the difference between a molecular and empirical formula?

3. What kinds of compounds (molecular, ionic, network) can be represented with a molecular formula?

4. What kinds of compounds (molecular, ionic, network) can be represented with an empirical formula?

Exercises

5. Using the periodic table, predict the chemical formula of the ionic compound formed by the following pairs of elements:
   Ga and F   Ca and O   Na and N   Al and O

6. Complete the following table by filling in the formula for the ionic compound formed by each pair of cations and anions, as shown for the first pair.

Ion	K^+	NH_4^+	Mg^{2+}	Fe^{3+}
S^{2-}	K_2S			
NO_3^-				
SO_4^{2-}				
PO_4^{3-}				

7. Predict whether each the following compounds is molecular or ionic: PF_5, NaI, SCl_2, B_2H_6, $LiNO_3$, NOCl, $CoCO_3$, NF_3

Information

Chemical nomenclature is a system for naming compounds. Today we follow rules set down by the International Union for Pure and Applied Chemistry (IUPAC). We will be concerned with the rules for naming simple inorganic compounds, as opposed to organic compounds. **Organic compounds** contain both C and H, and may contain one more other elements, such as N, O, S, F, Cl, Br, I, P. Any other compound that does not have *both* C and H is an **inorganic compound**. The rules for naming simple inorganic compounds are based on identifying the compound as either molecular or ionic. Therefore, it is very important that you be able to recognize the difference in order to name the compound correctly. The rules for naming simple inorganic compounds, which you need to learn, are described in section 2.8 of your text. There is a key distinction between the naming system for ionic compounds and that for molecular compounds; viz., *numerical prefixes (mono-, di-, tri-, tetra, etc.) are not used with ionic compounds, because the numbers of each element in the empirical formula can be deduced by the charges on the component ions.* By contrast, numerical prefixes *are* used with molecular compounds, because there is no simple way to deduce the numbers of atoms of each element in the molecular (or empirical) formula. Acids have their own set of naming rules (cf. sec. 2.8), which you must also know. You also need to be able to recognize and name a few organic compounds, such as simple alkanes and alcohols (see sec. 2.9).

Key Questions

8. In the chemical formula and name, which element is given first, a metal or nonmetal?

9. What suffix (ending) is added to the root of the name of the nonmetal in naming an ionic compound?

10. When a metal ion can form more than one kind of cation, how is the charge on the cation indicated in the name of a compound?

Exercises

11. Name the following molecular compounds: SCl_2, N_2O_4, P_4O_{10}, PF_5

12. Some molecular compounds are not named systematically, but rather retain their traditional names. Name the following compounds that retain their traditional names: H_2O, NH_3, H_2O_2, H_2S

13. Name the following ionic compounds: $AlCl_3$, Li_3PO_4, $Ba(ClO_4)_2$, $Cu(NO_3)_2$, $Fe_2(SO_4)_3$, $Ca(C_2H_3O_2)_2$, $Cr_2(CO_3)_3$, K_2CrO_4, $(NH_4)_2SO_4$

14. Give the name or chemical formula, as appropriate, for each of the following acids: $HClO_4$, HBr, H_3PO_4, hypochlorous acid, iodic acid, sulfurous acid

15. Name the following simple organic compounds: CH_4, C_2H_6, CH_3OH, C_2H_5OH

Information

A chemical equation is a written expression of a chemical reaction; e.g.,

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

Reactants are written on the left, and **products** are written on the right. In a **balanced equation** the total numbers of atoms of each kind on both sides are the same. To achieve a balance, we write **coefficients** in front of each chemical species, although the number 1 is never written as a coefficient. The equation $2\ H_2 + O_2 \rightarrow 2\ H_2O$ can be taken to say, "For every two molecules of H_2, one molecule of O_2 reacts to form two molecules of H_2O." Note that the language of this statement implies ratio relationships between reactants and products. We will use these implied ratio relationships shortly to carry out calculations based on balanced chemical equations. It is important at this stage to realize that a balanced equation in general *does not* tell us anything about how the reaction proceeds on a molecular level (the reaction's *mechanism*). It is a mistake to think that the balanced equation $2\ H_2 + O_2 \rightarrow 2\ H_2O$ necessarily means that two molecules of H_2 react directly with one molecule of O_2 in forming two molecules of H_2O. Indeed, the mechanism by which this reaction occurs is less direct.

A **skeletal equation** is a reaction equation that has not yet been balanced. It is just a statement of what the reactants and products are. We will often use a skeletal equation as a starting point for writing a balanced chemical equation. Initially, we will balance equations by inspection (i.e., not by some systematic approach). When doing this, it is important to realize that subscripts are part of the molecular or empirical formula for each species, and changing a subscript would change the identity of the species. Therefore, *subscripts are never altered in constructing a balanced chemical equation*. Here are some tips for taking a skeletal reaction equation and turning it into a balanced chemical equation:

1. Do not add new species or change the formulas of given species in the skeletal equation.
2. First balance elements that occur in only one reactant and one product (if such exist).

3. Fractional coefficients may be useful in achieving a balance (e.g., when a diatomic reactant yields an odd number of atoms in product molecules), but in general fractions should be cleared by multiplying both sides of the equation by a suitable constant.
4. If polyatomic ions or fragments of such occur as both reactants and products, treat them as discrete units (i.e., don't break them up).
5. Check element by element to be sure that the same numbers of each element occur on both sides of the equation (i.e., make sure the equation is really balanced).

Example
Given the skeletal equation $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$, write the balanced equation.

Step 1. We look over the given reactants ($C_2H_6 + O_2$) and products ($CO_2 + H_2O$) and see that reactant oxygen ends up in two products. Therefore, we will *not* start by attempting to balance the oxygen. By contrast, we see that all of the carbon in C_2H_6 ends up only in CO_2, and all of the hydrogen ends up only in H_2O. Therefore, we can start by attempting to balance carbon and hydrogen separately, leading to the following intermediate result:

$$C_2H_6 + O_2 \rightarrow 2\ CO_2 + 3\ H_2O$$

Step 2. By balancing the carbon and hydrogen, we have created seven oxygen atoms on the right, four from CO_2 and three from H_2O. We need to find a coefficient for O_2 that gives us seven oxygen atoms. But the oxygen exists as diatomic molecules, and we cannot break these up when writing the reaction to give the odd number we need to achieve a balance. However, if we place the coefficient 7/2 in front of O_2, that will give us seven oxygen atoms to balance the seven oxygen atoms in $2\ CO_2 + 3\ H_2O$. This gives us the following balanced equation:

$$C_2H_6 + 7/2\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O$$

Step 3. In some contexts (e.g., thermodynamics) we may be quite content to use this balanced equation in this form. But in most cases (all of them for now), we will want the balanced equation to be written with lowest whole number coefficients throughout. If we were to multiply the equation in Step 2 by the factor 2, we would clear the fraction and obtain the following result:

$$2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$$

Step 4. We check to be sure that we have actually written a balanced equation by counting the numbers of atoms of each element on the left and right. We find

✔ 4 C from $2\ C_2H_6$ = 4 C from $4\ CO_2$
✔ 12 H from $2\ C_2H_6$ = 12 H from $6\ H_2O$
✔ 14 O from $7\ O_2$ = 14 O from $4\ CO_2 + 6\ H_2O$

The equation is balanced!

Exercise

16. Balance the following skeletal equations, using lowest whole-number coefficients:

$N_2O_5 + H_2O \rightarrow HNO_3$

$Ca(OH)_2 + H_3PO_4 \rightarrow Ca_3(PO_4)_2 + H_2O$

$C_5H_{10}O_2 + O_2 \rightarrow CO_2 + H_2O$

Information

The mass of the molecules in a molecular compound, called the **molecular weight** (m.w.), can be calculated by simply adding up the masses of all the atoms comprising the molecules.

$$\text{m.w. } N_2O_4 = (2)(14.01 \text{ u}) + (4)(16.00 \text{ u}) = 92.02 \text{ u}$$

Any compound, whether molecular or ionic, can be described by an empirical formula. In fact, for ionic or network compounds, an empirical formula is the only kind of chemical formula we can write. Although the empirical formula does not correspond to any real chemical entity, such as a molecule, we often find it useful to define the mass of this hypothetical unit, called the **formula weight** (f.w.).

$$\text{molecular formula } N_2O_4 \Rightarrow \text{empirical formula } NO_2$$
$$\text{f.w. } NO_2 = 14.01 + (2)(16.00 \text{ u}) = 46.01 \text{ u}$$
$$\text{m.w. } N_2O_5 = 2 \times \text{f.w. } N_2O_5 = 2 \times 46.01 \text{ u} = 92.02 \text{ u}$$

Percent composition gives the element by element percentages of the masses of all elements in a compound. This can be calculated for each element, using either a molecular or empirical formula, by taking the ratio of the mass of the element in the compound to the molecular weight or formula weight, and then multiplying by 100%. For N_2O_4, for example, we could use the empirical formula NO_2 and calculate the percentages of nitrogen and oxygen as follows:

$$\%\text{ N} = \frac{14.01}{46.01} \times 100\% = 30.45\%$$
$$\%\text{ O} = \frac{(2)(16.00)}{46.01} \times 100\% = 69.55\%$$

Of course, the sum of percentages for all elements in the compound adds to 100%.

Key Questions

17. Why is it incorrect to talk about the molecular weight of NaCl?

18. Would the sum of the masses of all atoms in the chemical formula C_2H_5OH be a molecular weight or a formula weight?

19. Is there a difference between the molecular weight and formula weight of the molecular compound N_2O_5?

Exercises

20. Calculate the molecular weight and formula weight of glucose, $C_6H_{12}O_6$.

21. Calculate the percent composition of glucose.

Periodic Table of the Elements

1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8	8B 9	10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1 H 1.008																	2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc [98]	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po [209]	85 At [210]	86 Rn [222]
87 Fr [223]	88 Ra [226]	89 Ac [227]	104 Rf [261]	105 Db [262]	106 Sg [263]	107 Bh [264]	108 Hs [265]	109 Mt [268]	110 Uun [269]	111 Uuu [272]	112 Uub [277]						

58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm [145]	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np [237]	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]	103 Lr [262]

Some Common Polyatomic Ions to Know

Cations		
Ammonium ion	NH_4^+	
Hydronium ion	H_3O^+	(In solution; no common compounds)
Mercury(I)	Hg_2^{2+}	(A diatomic molecule of two Hg^+ ions with a +2 charge)
Anions		
Acetate ion	$C_2H_3O_2^-$	(Also written $CH_3CO_2^-$ or CH_3COO^-)
Carbonate ion	CO_3^{2-}	
Hydrogen carbonate ion	HCO_3^-	(Formerly called bicarbonate)
Hypochlorite ion	ClO^-	
Chlorite ion	ClO_2^-	
Chlorate ion	ClO_3^-	
Perchlorate ion	ClO_4^-	
Chromate ion	CrO_4^{2-}	
Dichromate ion	$Cr_2O_7^{2-}$	
Cyanate ion	OCN^-	
Thiocyanate ion	SCN^-	
Cyanide ion	CN^-	
Hydroxide ion	OH^-	
Nitrite ion	NO_2^-	
Nitrate ion	NO_3^-	
Oxalate ion	$C_2O_4^{2-}$	
Permanganate ion	MnO_4^-	
Peroxide	O_2^{2-}	
Phosphate ion	PO_4^{3-}	
Hydrogen phosphate ion	HPO_4^{2-}	(Also called monohydrogen phosphate)
Dihydrogen phosphate ion	$H_2PO_4^-$	
Sulfite ion	SO_3^{2-}	
Hydrogen sulfite ion	HSO_3^-	(Formerly called bisulfite)
Sulfate ion	SO_4^{2-}	
Hydrogen sulfate ion	HSO_4^-	(Formerly called bisulfate)
Thiosulfate ion	$S_2O_3^{2-}$	