

## 3

## Chemical Foundations

## Chapter Objectives (Sections 3.4–3.7):

- Learn the basic concepts behind chemical bonds.
- Learn how to balance chemical equations.
- Learn how to write formulas and name chemical compounds.

Mr. Kevin A. Boudreaux

Angelo State University  
www.angelo.edu/faculty/kboudreaux

1

## Ionic and Molecular Compounds

2

## Chemical Compounds

- Most substances that we encounter are *compounds*, not elements.
- A **chemical compound** is a pure substance formed from the combination of two or more different elements. The properties of the compound may be completely unlike those of the elements that form it.
- The **chemical formula** for a compound lists the symbols of the individual elements followed by subscripts which indicate the number of atoms of that element.
  - If no subscript is given, it is understood to be “1.”
  - e.g., NaCl, H<sub>2</sub>O, H<sub>2</sub>O<sub>2</sub>, C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>, C<sub>34</sub>H<sub>32</sub>O<sub>4</sub>N<sub>4</sub>Fe.

3

## Types of Bonding

- When two atoms collide during a reaction, it is the electrons that make the actual contact, since they occupy a much greater volume than the nucleus.
- Thus, it is the electrons that form the connections, or chemical bonds, that join atoms together to form compounds.
- Elements combine to form compounds in two different ways:
  - **transferring electrons** from atoms of one element to another results in **ionic bonds**.
  - **sharing electrons** between atoms of different elements results in **covalent bonds**.

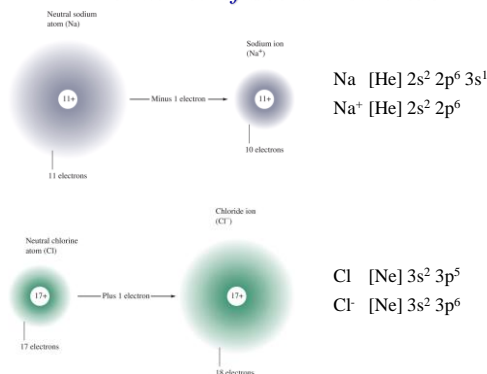
4

## Ionic Compounds

- **Ionic bonds** form when one atom **transfers** one or more electrons to another atom, producing ions.
- **Ionic compounds** are compounds that are held together by **ionic bonds** between positively-charged cations and negatively-charged anions.
- Ionic compounds generally result when a metal combines with a nonmetal:
  - **Metal + Nonmetal** → ionic compound
  - **Metal + Polyatomic ion** → ionic compound
  - In general, metals tend to give up electrons, while nonmetals accept electrons.

5

## Formation of Sodium Chloride



6

### Formation of Sodium Chloride

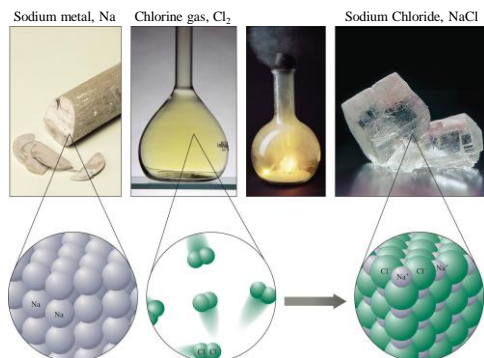
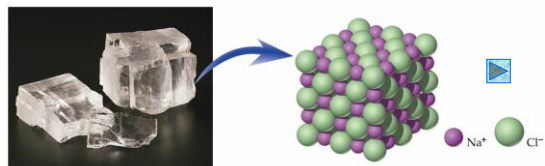


Figure 3.6

7

### Ionic Solids and Crystals

- The ionic bond is the strong attraction between the cations and the anions. Unlike molecules, *the cation and anion are not physically joined together.*
- Thus, there is *no molecule* of NaCl; ionic compounds instead form **ionic solids**, which contain equal amounts of positive and negative charge surrounding each other in a regular array called a **crystal**.



8

### Formulas of Ionic Compounds

- The smallest unit of an ionic compound is the **formula unit**, the smallest electrically neutral collection of ions.
- Monatomic ions** are cations or anions derived from a single atom, such as Cl<sup>-</sup>, O<sup>2-</sup>, Na<sup>+</sup>, and Mg<sup>2+</sup>.
- Polyatomic ions** are combinations of atoms that possess an overall charge, such as CO<sub>3</sub><sup>2-</sup>, SO<sub>4</sub><sup>2-</sup>, NO<sub>3</sub><sup>-</sup>, CN<sup>-</sup>, NH<sub>4</sub><sup>+</sup>, C<sub>2</sub>H<sub>3</sub>O<sub>2</sub><sup>-</sup>, etc.
- Ionic compounds are electrically neutral* — there must be the same amount of positive charge as there is negative charge.

9

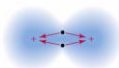
### Examples: Writing Ionic Formulas

- Write the formula for the ionic compound formed between the following pairs of elements and provide a name for the compound.
  - Al and F
  - Na and S
  - Ba and S
  - Mg and P
  - Ca and Cl
  - Na and P

10

### Covalent Bonds and Molecules

- Covalent bonds** form when two or more nonmetals **share** their electrons. The electrons are at their lowest potential energy when they are *between* the two nuclei that are being joined.



- Each atom in the bond “holds on” to the shared electrons, and the atoms are thus physically tied together.

11

### Molecular Compound

- When two or more atoms are joined by covalent bonds, the resulting structure is called a **molecule** (or **molecular compound**).
- Molecular compounds are usually formed from combinations of nonmetals:
  - Nonmetal + Nonmetal** → molecular compound
  - Hydrogen + Nonmetal** → molecular compound

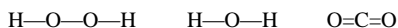
12

## Types of Chemical Formulas

- An **empirical formula** gives the *relative* number of atoms of each element in a compound; i.e., the *smallest* whole number ratio that is possible.
- A **molecular formula** gives the *actual* number of atoms of each element in a molecule of a compound.

	Molecular formula	Empirical formula
Hydrogen peroxide	H <sub>2</sub> O <sub>2</sub>	HO
Water	H <sub>2</sub> O	H <sub>2</sub> O
Glucose	C <sub>6</sub> H <sub>12</sub> O <sub>6</sub>	CH <sub>2</sub> O

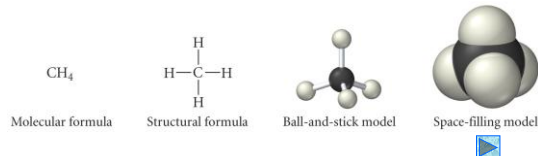
- A **structural formula** uses lines to represent covalent bonds, and shows how the atoms in a molecule are joined together:



13

## Molecular Models

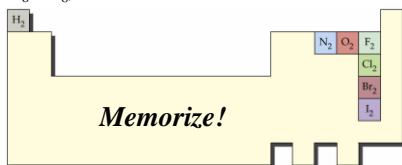
- A **ball-and-stick model** represents atoms as balls and covalent bonds as sticks; they are normally color-coded to specific elements, and show the three-dimensional relationships between atoms in a molecule.
- A **space-filling model** shows the room taken up by the electron clouds in the molecule, and shows how the molecule might appear if it were scaled to a visible size.



14

## Atomic and Molecular Elements

- Atomic elements** are found in nature in units of single atoms. Most elements are atomic elements.
- Molecular elements** are found as molecules with two or more of the same atom joined together.
  - Many nonmetals are found in their elemental form as **diatomic** molecules (H<sub>2</sub>, N<sub>2</sub>, O<sub>2</sub>, F<sub>2</sub>, Cl<sub>2</sub>, Br<sub>2</sub>, I<sub>2</sub>); some as polyatomic molecules (P<sub>4</sub>, As<sub>4</sub>, S<sub>8</sub>, Se<sub>8</sub>)



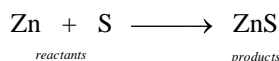
15

16

# Chemical Equations

## Chemical Reactions and Chemical Equations

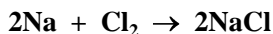
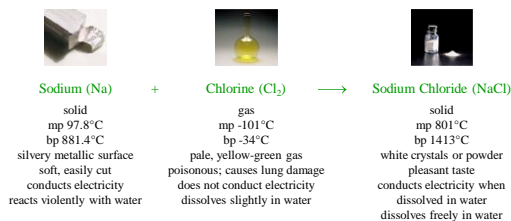
- A **chemical reaction** occurs when atoms of different elements combine and create a new chemical compound, with properties which may be completely unlike those of its constituent elements.
- A chemical reaction is written in a standard format called a **chemical equation**. The **reactants** (starting materials) are written on the left, and the **products** on the right, with an arrow in between to indicate a transformation.
- Equations are the “sentences” of chemistry, just as formulas are the “words” and atomic symbols are the “letters.”



1

2

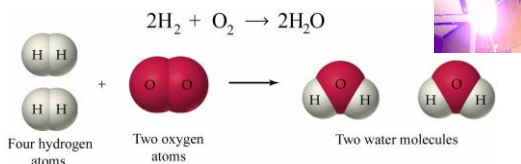
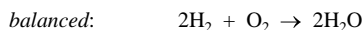
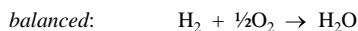
## A Chemical Reaction Illustrated



3

## Balancing Chemical Reactions

- A chemical equation must be **balanced**: the kinds and numbers of atoms must be the same on both sides of the reaction arrow (conservation of mass).



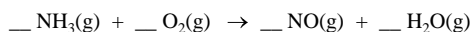
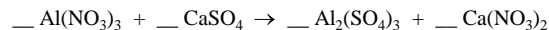
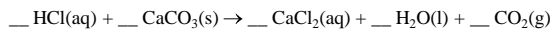
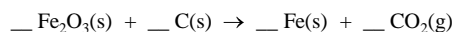
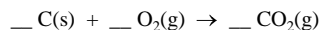
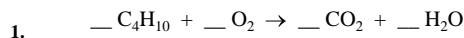
4

## Balancing Chemical Reactions

- Equations are balanced by placing a **stoichiometric coefficient** in front of each species, indicating how many units of each compound participate in the reaction.
  - If no coefficient is present, it is assumed to be 1.
  - Usually, we use the smallest whole-number ratios for the coefficients.
  - Never balance equations by changing subscripts!** This changes the identity of the species involved in the reaction!
  - In general, it's a good idea to balance the atoms in the most complex substances first, and the atoms in the simpler substances last.

5

## Examples: Balancing Equations



6



# Naming Chemical Compounds

## Main-Group Metals

- Group 1A, 2A, and 3A 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, with the word “ion” added to the end.

Group	Ion	Ion name	Group	Ion	Ion name
1A	H <sup>+</sup>	hydrogen ion	2A	Mg <sup>2+</sup>	magnesium ion
	Li <sup>+</sup>	lithium ion		Ca <sup>2+</sup>	calcium ion
	Na <sup>+</sup>	sodium ion		Sr <sup>2+</sup>	strontium ion
	K <sup>+</sup>	potassium ion	Ba <sup>2+</sup>	barium ion	
	Cs <sup>+</sup>	cesium ion	3A	Al <sup>3+</sup>	aluminum ion

1

2

## Transition and Post-Transition 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 4A and 5A metal cations tend to be either the group number, or the group number minus two.)
- Common or trivial names: *-ic* endings go with the higher charge, *-ous* endings go with the lower charge.
  - Fe<sup>2+</sup> ferrous ion, Fe<sup>3+</sup> ferric ion
- Systematic names (*Stock system*): name the metal first, followed in parentheses by the charge in Roman numerals.
  - Fe<sup>2+</sup> iron(II) ion, Fe<sup>3+</sup> iron(III) ion

3

## Transition and Post-Transition Metals

Ion	Systematic name	Common name
Cr <sup>2+</sup>	chromium(II) ion	chromous ion
Cr <sup>3+</sup>	chromium(III) ion	chromic ion
Mn <sup>2+</sup>	manganese(II) ion	manganous ion
Mn <sup>3+</sup>	manganese(III) ion	manganic ion
Fe <sup>2+</sup>	iron(II) ion	ferrous ion
Fe <sup>3+</sup>	iron(III) ion	ferric ion
Co <sup>2+</sup>	cobalt(II) ion	cobaltous ion
Co <sup>3+</sup>	cobalt(III) ion	cobaltic ion
Ni <sup>2+</sup>	nickel(II) ion	
Cu <sup>+</sup>	copper(I) ion	cuprous ion
Cu <sup>2+</sup>	copper(II) ion	cupric ion
Zn <sup>2+</sup>	zinc ion	
Ag <sup>+</sup>	silver ion	
Cd <sup>2+</sup>	cadmium ion	

4

## Transition and Post-Transition Metals

Ion	Systematic name	Common name
Au <sup>3+</sup>	gold(III) ion	
Hg <sub>2</sub> <sup>2+</sup>	mercury(I) ion	mercurous ion
Hg <sup>2+</sup>	mercury(II) ion	mercuric ion
Sn <sup>2+</sup>	tin(II) ion	stannous ion
Sn <sup>4+</sup>	tin(IV) ion	stannic ion
Pb <sup>2+</sup>	lead(II) ion	plumbous ion
Pb <sup>4+</sup>	lead(IV) ion	plumbic ion
Bi <sup>3+</sup>	bismuth(III) ion	
Bi <sup>5+</sup>	bismuth(V) ion	

5

## Main-Group Nonmetals

- Group 4A - 7A nonmetals form *anions* by gaining enough electrons to fill their valence shell (eight electrons). The charge on the anion is the group number minus eight.
- The anion is named by taking the element stem and adding the ending *-ide*.

Group	Ion	Ion name	Group	Ion	Ion name
4A	C <sup>4-</sup>	carbide ion	6A	Se <sup>2-</sup>	selenide ion
	Si <sup>4-</sup>	silicide ion		Te <sup>2-</sup>	telluride ion
5A	N <sup>3-</sup>	nitride ion	7A	F <sup>-</sup>	fluoride ion
	P <sup>3-</sup>	phosphide ion		Cl <sup>-</sup>	chloride ion
	As <sup>3-</sup>	arsenide ion		Br <sup>-</sup>	bromide ion
6A	O <sup>2-</sup>	oxide ion		I <sup>-</sup>	iodide ion
	S <sup>2-</sup>	sulfide ion	1A	H <sup>-</sup>	hydride ion

6

## Common Cations and Anions

Elements To Memorize																																																					
Atomic Number																																																					
Symbol																																																					
Name																																																					
1	2															10	11	12																																			
H	He															Ne	Ar	Kr	Xe	Rn																																	
3	4															9	10	11	12	13	14	15	16	17	18																												
Li	Be															B	C	N	O	F	Ne	Ar	Kr	Xe	Rn																												
5	6															7	8	9	10	11	12	13	14	15	16	17	18																										
Na	Mg															Al	Si	P	S	Cl	Ar	Kr	Xe	Rn																													
19	20															19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36																				
K	Ca															Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	Xe	Rn																				
37	38															37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54																				
Rb	Sr															Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	Rn																					
55	56															55	56	57	58	59	60	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78	79	80												
Cs	Ba															La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn								
87	88															87	88	89	90	91	92	93	94	95	96	97	98	99	100	101	102	103	104	105	106	107	108	109	110	111	112	113	114	115	116	117	118						
Ra	Ac															101	102	103	104	105	106	107	108	109	110	111	112	113	114	115	116	117	118	119	120	121	122	123	124	125	126	127	128	129	130	131	132	133	134	135	136	137	138
Lanthanoids																																																					
Actinoids																																																					

7

## Polyatomic Ions

- Polyatomic ions** are ions composed of groups of covalently bonded atoms which have an overall charge.

$\text{NH}_4^+$	ammonium	$\text{OCN}^-$	cyanate
$\text{H}_3\text{O}^+$	hydronium	$\text{MnO}_4^-$	permanganate
$\text{OH}^-$	hydroxide	$\text{C}_2\text{H}_3\text{O}_2^-$	acetate ( $\text{OAc}^-$ , $\text{CH}_3\text{CO}_2^-$ )
$\text{CN}^-$	cyanide	$\text{CO}_3^{2-}$	carbonate
$\text{O}_2^{2-}$	peroxide	$\text{HCO}_3^-$	hydrogen carbonate, bicarbonate
$\text{N}_3^-$	azide	$\text{SO}_4^{2-}$	sulfate
$\text{NO}_3^-$	nitrate	$\text{SO}_3^{2-}$	sulfite
$\text{NO}_2^-$	nitrite	$\text{S}_2\text{O}_3^{2-}$	thiosulfate
$\text{ClO}_3^-$	chlorate	$\text{C}_2\text{O}_4^{2-}$	oxalate
$\text{ClO}_2^-$	chlorite	$\text{CrO}_4^{2-}$	chromate
$\text{ClO}^-$	hypochlorite	$\text{Cr}_2\text{O}_7^{2-}$	dichromate
$\text{ClO}_4^-$	perchlorate	$\text{PO}_4^{3-}$	phosphate

8

## Polyatomic Ions — Regularities in Names

- There are some regularities in the names of these polyatomic ions:
- Thio-** implies replacing an oxygen with a sulfur:
 

$\text{SO}_4^{2-}$	sulfate	$\text{OCN}^-$	cyanate
$\text{S}_2\text{O}_3^{2-}$	thiosulfate	$\text{SCN}^-$	thiocyanate
- Replacing the first element with another element from the same group gives a polyatomic ion with the same charge, and a similar name:

Group 7A	Group 6A	Group 5A	Group 4A
$\text{ClO}_3^-$	chlorate	$\text{SO}_4^{2-}$	sulfate
$\text{BrO}_3^-$	bromate	$\text{PO}_4^{3-}$	phosphate
$\text{IO}_3^-$	iodate	$\text{CO}_3^{2-}$	carbonate
		$\text{SeO}_4^{2-}$	selenate
		$\text{AsO}_4^{3-}$	arsenate
		$\text{SiO}_3^{2-}$	silicate
		$\text{TeO}_4^{2-}$	tellurate

9

## Polyatomic Ions — Oxoanions

- Some nonmetals form a series of **oxoanions** having different numbers of oxygens (all with the same charge). The general rule for such series is shown below. (Note that in some cases, the *-ate* form has three oxygens, and in some cases four oxygens. These forms must be memorized.)

$\text{XO}_n^{y-}$	<i>stem + ate</i>	$\text{ClO}_3^-$	chlorate
$\text{XO}_{n-1}^{y-}$	<i>stem + ite</i>	$\text{ClO}_2^-$	chlorite
$\text{XO}_{n-2}^{y-}$	<i>hypo + stem + ite</i>	$\text{ClO}^-$	hypochlorite
$\text{XO}_{n+1}^{y-}$	<i>per + stem + ate</i>	$\text{ClO}_4^-$	perchlorate
$\text{X}^{y-}$	<i>stem + ide</i> (the monatomic ion)	$\text{Cl}^-$	chloride

10

## Polyatomic Ions — Ions Containing Hydrogens

- Acid salts** are ionic compounds that still contain an acidic hydrogen, such as  $\text{NaHSO}_4$ . In naming these salts, specify the number of acidic hydrogens still in the salt.
- The prefix *bi-* implies an acidic hydrogen.

$\text{CO}_3^{2-}$	carbonate
$\text{HCO}_3^-$	hydrogen carbonate, bicarbonate
$\text{SO}_4^{2-}$	sulfate
$\text{HSO}_4^-$	hydrogen sulfate, bisulfate
$\text{PO}_4^{3-}$	phosphate
$\text{HPO}_4^{2-}$	monohydrogen phosphate
$\text{H}_2\text{PO}_4^-$	dihydrogen phosphate

11

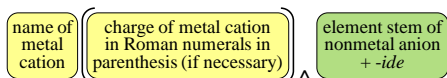
## Writing Formulas of Ionic Compounds

- The positive ion is given first, followed by the monatomic or polyatomic anion.
- The subscripts in the formula must produce an electrically neutral formula unit.
- The subscripts should be the smallest set of whole numbers possible.
- If there is only one of a polyatomic ion in the formula, do not place parentheses around it. If there is more than one of a polyatomic ion, put the ion in parentheses, and place the subscript after the parentheses.
- Remember the Prime Directive in writing formulas:  $\text{Ca}(\text{OH})_2 \neq \text{CaOH}_2!$

12

### Nomenclature of Ionic Compounds: Metal + Nonmetal

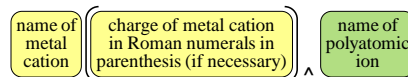
- A *binary compound* is a compound formed from *two different elements*. A *diatomic compound* (or diatomic molecule) contains *two atoms*, which may or may not be the same.
- Metals combine with nonmetals to form ionic compounds. Name the cation first (specify the charge, if necessary), then the nonmetal anion (element stem + *-ide*).
- Do NOT use counting prefixes! This information is implied in the name of the compound.



13

### Nomenclature of Ionic Compounds: Metal + Polyatomic Ion

- Metals combine with polyatomic ions to give ionic compounds. Name the cation first (specify the charge, if necessary), then the polyatomic ion as listed in the previous table.
- Once again, do NOT use counting prefixes!



14

### Nomenclature of Ionic Compounds: Examples

Na <sup>+</sup>	Cl <sup>-</sup>	→	NaCl	sodium chloride
Na <sup>+</sup>	S <sup>2-</sup>	→	Na <sub>2</sub> S	sodium sulfide
Na <sup>+</sup>	P <sup>3-</sup>	→	Na <sub>3</sub> P	sodium phosphide
Ca <sup>2+</sup>	Cl <sup>-</sup>	→	CaCl <sub>2</sub>	calcium chloride
Ca <sup>2+</sup>	S <sup>2-</sup>	→	CaS	calcium sulfide
Fe <sup>2+</sup>	Cl <sup>-</sup>	→	FeCl <sub>2</sub>	iron(II) chloride
Fe <sup>3+</sup>	Cl <sup>-</sup>	→	FeCl <sub>3</sub>	iron(III) chloride
Na <sup>+</sup>	SO <sub>4</sub> <sup>2-</sup>	→	Na <sub>2</sub> SO <sub>4</sub>	sodium sulfate
Ca <sup>2+</sup>	CO <sub>3</sub> <sup>2-</sup>	→	CaCO <sub>3</sub>	calcium carbonate
Cr <sup>2+</sup>	NO <sub>3</sub> <sup>-</sup>	→	Cr(NO <sub>3</sub> ) <sub>2</sub>	chromium(II) nitrate
Al <sup>3+</sup>	O <sup>2-</sup>	→	Al <sub>2</sub> O <sub>3</sub>	aluminum oxide

15

### Nomenclature of Ionic Compounds: Hydrates

- Hydrates** are ionic compounds which also contain a specific number of water molecules associated with each formula unit. The water molecules are called *waters of hydration*.
- The formula for the ionic compound is followed by a raised dot and #H<sub>2</sub>O — e.g., MgSO<sub>4</sub>·7H<sub>2</sub>O.
- They are named as ionic compounds, followed by a counting prefix and the word “hydrate”

MgSO <sub>4</sub> ·7H <sub>2</sub> O	magnesium sulfate heptahydrate (Epsom salts)
CaSO <sub>4</sub> ·½H <sub>2</sub> O	calcium sulfate hemihydrate
BaCl <sub>2</sub> ·6H <sub>2</sub> O	barium chloride hexahydrate
CuSO <sub>4</sub> ·5H <sub>2</sub> O	copper(II) sulfate pentahydrate

16

### Nomenclature of Binary Molecular Compounds

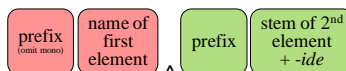
- Two nonmetals combine to form a *molecular* or *covalent compound* (i.e., one that is held together by covalent bonds, not ionic bonds).
- In many cases, two elements can combine in several ways to make completely different compounds (e.g., CO and CO<sub>2</sub>). It is necessary to specify how many of each element is present within the compound.
- In writing formulas, the more cation-like element (the one further to the left on the periodic table) is placed first, then the more anion-like element (the one further to the right on the periodic table).
- Important exception:** halogens are written before oxygen. For two elements in the same group, the one with the higher period number is placed first.

17

### Nomenclature of Binary Molecular Compounds

- The first element in the formula is given the element name, and the second one is named by replacing the ending of the element name with *-ide*.
- A numerical prefix is used in front of each element name to indicate how many of that element is present. (If there is only one of the first element in the formula, the *mono-* prefix is dropped.)

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



18

*Nomenclature of Binary Molecular Compounds*

- NO nitrogen monoxide  
 NO<sub>2</sub> nitrogen dioxide  
 N<sub>2</sub>O dinitrogen monoxide  
 N<sub>2</sub>O<sub>3</sub> dinitrogen trioxide  
 N<sub>2</sub>O<sub>4</sub> dinitrogen tetroxide  
 N<sub>2</sub>O<sub>5</sub> dinitrogen pentoxide
- Some molecular compounds are known by common or trivial names:
    - H<sub>2</sub>O water
    - NH<sub>3</sub> ammonia

19

*Examples: Formulas and Nomenclature*

- Write the formula for the ionic compound formed between the following pairs of species and provide a name for the compound.
  - Na and sulfate
  - Ammonium and nitrate
  - Ammonium and sulfate
  - Zn and Cl
  - Mercury(I) and nitrite
  - Mercury(II) and sulfite
  - Chromium and S

20

*Examples: Formulas and Nomenclature*

- Name the following compounds.
  - Ca(NO<sub>3</sub>)<sub>2</sub>
  - BaCO<sub>3</sub>
  - SO<sub>3</sub>
  - SnCl<sub>4</sub>
  - Fe<sub>2</sub>(CO<sub>3</sub>)<sub>3</sub>
  - AlPO<sub>4</sub>
  - N<sub>2</sub>O

21

*Examples: Formulas and Nomenclature*

- Name the following compounds.
  - CrO
  - Mn<sub>2</sub>O<sub>3</sub>
  - NO<sub>2</sub>
  - NaNO<sub>2</sub>
  - PBr<sub>3</sub>
  - KHSO<sub>4</sub>
  - LiH<sub>2</sub>PO<sub>4</sub>

22

*Examples: Formulas and Nomenclature*

- Write formulas for the following compounds.
  - sodium nitrite
  - lithium hydroxide
  - barium chlorate
  - potassium perchlorate
  - diphosphorus pentoxide
  - magnesium phosphate
  - iron(II) carbonate

23

*Examples: Formulas and Nomenclature*

- Write formulas for the following compounds.
  - calcium bicarbonate
  - manganese(III) carbonate
  - potassium hypochlorite
  - silver chromate
  - nickel acetate
  - barium peroxide
  - titanium(IV) oxide

24

*Nomenclature of Acids*

- Acids are compounds in which the “cation” is  $H^+$ . These are often given special “acid names” derived by omitting the word “hydrogen,” adding the word “acid” at the end, and changing the compound suffix as shown below:

	Compound name	Acid name
oxyacids	<i>stem + ate</i>	<i>stem + ic acid</i>
	<i>stem + ite</i>	<i>stem + ous acid</i>
binary acids	<i>stem + ide</i>	<i>hydro + stem + ic acid</i>
	HClO <sub>3</sub> hydrogen chlorate	chloric acid
	H <sub>2</sub> SO <sub>4</sub> hydrogen sulfate	sulfuric acid
	HClO <sub>2</sub> hydrogen chlorite	chlorous acid
	HCl hydrogen chloride	hydrochloric acid

25

*Examples: Acid Nomenclature*

6. Write formulas or names for the following acids.

- a. HCl \_\_\_\_\_
- b. HClO<sub>2</sub> \_\_\_\_\_
- c. H<sub>2</sub>SO<sub>3</sub> \_\_\_\_\_
- d. H<sub>3</sub>PO<sub>4</sub> \_\_\_\_\_
- e. hydrofluoric acid \_\_\_\_\_
- f. periodic acid \_\_\_\_\_
- g. chloric acid \_\_\_\_\_
- h. phosphorous acid \_\_\_\_\_

26

*Examples: Formulas and Nomenclature*

7. Which of the following formulas and/or names is written *incorrectly*?
- a. NaSO<sub>4</sub>
- b. Na<sub>2</sub>Cl
- c. MgNO<sub>3</sub>
- d. magnesium dichloride
- e. iron(III) phosphate, Fe<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub>
- f. tin(IV) sulfate, Sn(SO<sub>4</sub>)<sub>2</sub>
- g. nitrogen chloride, NCl<sub>3</sub>
- h. HClO<sub>2</sub>, hypobromous acid

27

The End

28

# 9 Stoichiometry

## Chapter Objectives (Sections 9.1–9.5):

- Learn how to use percent compositions to find empirical and molecular formulas.
- Learn how to use the mole concept to relate amounts of chemicals to each other (stoichiometry).
- Learn how to find the theoretical yield in limiting reactant problems.

Mr. Kevin A. Boudreaux

Angelo State University  
www.angelo.edu/faculty/kboudreaux

1

2

# Percent Composition and Empirical Formulas

## Percent Composition and Mass Percentage

- The **percent composition** of a compound is a list of the elements present in a substance listed by *mass percent*. Knowing the percent composition is often a first step to determining the formula of an unknown compound.
- The **mass percentage (mass %)** of an element in the compound is the portion of the compound's mass contributed by that element, expressed as a percentage:

$$\text{Mass \% of element X} = \frac{\text{atoms of X in formula} \times \text{molar mass of X}}{\text{molar mass of compound}} \times 100$$

3

## Percent Composition and Mass Percentage

- What is the mass percentage of Cl in the chlorofluorocarbon  $\text{CCl}_2\text{F}_2$  (Freon-12)?

$$\begin{aligned} \text{Mass \% of Cl} &= \frac{2 \times \text{atomic mass of Cl}}{\text{molar mass of CCl}_2\text{F}_2} \times 100 \\ &= \frac{2 \times 35.453 \text{ g/mol}}{120.91 \text{ g/mol}} \times 100 \\ &= 58.64\% \end{aligned}$$

4

### Examples: Mass Percentage

1. Glucose, or blood sugar, has the molecular formula  $\text{C}_6\text{H}_{12}\text{O}_6$ .
  - a. What is the percent composition of glucose? (similar to Example 9.1)
  - b. How many grams of carbon are in 39.0 g of glucose (the amount of sugar in a typical soft drink)?

**Answer:** a) 40.00% C, 6.714% H, 53.29% O  
b) 15.6 g C

5

### Examples: Mass Percentage

2. The U.S. Food and Drug Administration (FDA) recommends that you consume less than 2.4 g of sodium per day. What mass of sodium chloride in grams can you consume and still be within the FDA guidelines? (similar to Example 3.14)

**Answer:** 6.1 g NaCl

6

### Empirical Formula from Mass Percentage

- We can use the percent composition of a substance to find its *empirical* and *molecular formula*.
- If by some process we determine the percent composition of an unknown compound, we can convert this into a gram ratio by assuming that we have 100 g of the compound, and then to a mole ratio by using the atomic weights:

Sample: 84.1% C, 15.9% H

Assume 100 g of sample:

$$84.1 \text{ g C} \times \frac{1 \text{ mol C}}{12.01115 \text{ g C}} = 7.00 \text{ mol C}$$

$$15.9 \text{ g H} \times \frac{1 \text{ mol H}}{1.00797 \text{ g H}} = 15.8 \text{ mol H}$$

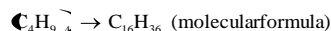
7

### Molecular Formula from Empirical Formula

- If we know the molar mass of the compound, we can obtain the **molecular formula** by dividing the weight of the empirical formula into the molar mass; this will determine the number of empirical formula units in the molecule.

Suppose the molecular weight of the substance is found to be 228.48 g/mol:

$$\frac{\text{molecular weight}}{\text{empirical formula weight}} \rightarrow \frac{228.48 \text{ g/mol}}{57.12 \text{ g/mol}} = 4.000$$



9

### Examples: Empirical & Molecular Formulas

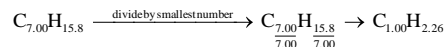
- Black iron oxide is an ore containing iron and oxygen that occurs in magnetite. A 2.4480 g sample of the ore is found to contain 1.7714 g of iron. Calculate the empirical formula of black iron oxide.

Answer: Fe<sub>3</sub>O<sub>4</sub>

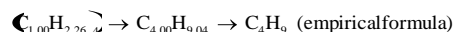
11

### Empirical Formula from Mass Percentage

- Since atoms combine in the same ratio that moles do, we divide all of the numbers of moles by the smallest number to put everything into lowest terms:



- If the mole ratio is not all whole numbers, we multiply through by the smallest integer which will turn all of the numbers into integers. These numbers are the subscripts of the elements in the **empirical formula**.



8

### Examples: Empirical & Molecular Formulas

- Vitamin C (ascorbic acid) contains 40.92% C, 4.58% H, and 54.50% O by mass. What is the empirical formula of ascorbic acid? (sim. to Example 9.2)

Answer: C<sub>3</sub>H<sub>4</sub>O<sub>3</sub>

10

### Examples: Empirical & Molecular Formulas

- Styrofoam is a polymer made from the monomer styrene. Elemental analysis of styrene shows its percent composition to be 92.26% C and 7.75% H. Its molecular mass is found to be 104.15 g/mol. What are the empirical and molecular formulas of styrene? (similar to example 9.3)

Answer: empirical = CH, molecular = C<sub>8</sub>H<sub>8</sub>

12

**Examples: Empirical & Molecular Formulas**

6. Butanedione is a main component in the smell and taste of butter and cheese. The empirical formula of butanedione is  $C_2H_3O$  and its molar mass is 86.09 g/mol. What is its molecular formula?

Answer:  $C_4H_6O_2$

13

**Elemental / Combustion Analysis**

- One common way of obtaining a chemical formula is by performing a **combustion analysis** (a specific type of *elemental analysis*).
- In this technique, an unknown sample is burned in pure  $O_2$  (a combustion reaction), which converts all of the carbon atoms in the sample into  $CO_2$  and all of the hydrogen atoms into  $H_2O$ .

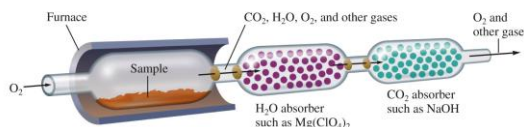
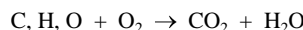
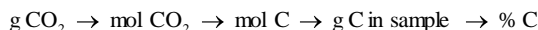


Figure 9.1

14

**Elemental / Combustion Analysis**

- The masses of  $CO_2$  and  $H_2O$  are measured after the process is complete, and from this data, the amount of carbon and hydrogen in the original sample can be determined.
- Elements besides C and H must be determined by other methods; O is usually found by difference.



$$\% O = 100\% - (\% C + \% H)$$

15

**Examples: Combustion Analysis**

7. A sample of an unknown compound with a mass of 0.5438 g is burned in a combustion analysis. The mass of  $CO_2$  produced was 1.039 g and the mass of  $H_2O$  was 0.6369 g. What is the empirical formula of the compound? (sim. to Examples 3.19 & 3.20)

Answer:  $C_2H_6O$

16

**Molecules and Isomers**

- Even knowing the empirical or molecular formulas of a compound does not necessarily tell us what that compound actually is.
- We've already seen that the empirical formula only tells us about the *relative* numbers of atoms present within the formula unit or molecule.
- Many different compounds can have the same empirical formula. For instance, there are dozens of different compounds that have the empirical formula  $CH_2O$ .

– Notice that in on the following slide, there is no relationship between the structure and how many 'CH<sub>2</sub>O' units the molecule contains.

17

**Some Compounds with Empirical Formula  $CH_2O$** 

- Composition by mass 40.0% C, 6.71% H, 53.3% O

Name	Molecular Formula	No. of 'CH <sub>2</sub> O' Units	Molar Mass (g/mol)	Function
Formaldehyde	CH <sub>2</sub> O	1	30.03	Disinfectant; biological preservative
Acetic acid	C <sub>2</sub> H <sub>4</sub> O <sub>2</sub>	2	60.05	Vinegar (5% solution); acetate polymers
Lactic acid	C <sub>3</sub> H <sub>6</sub> O <sub>3</sub>	3	90.08	Found in sour milk and sourdough bread; forms in muscles during exercise
Erythrose	C <sub>4</sub> H <sub>8</sub> O <sub>4</sub>	4	120.10	Forms during sugar metabolism
Ribose	C <sub>5</sub> H <sub>10</sub> O <sub>5</sub>	5	150.13	Component of ribonucleic acid (RNA); found in vitamin B <sub>5</sub>
Glucose	C <sub>6</sub> H <sub>12</sub> O <sub>6</sub>	6	180.16	Major nutrient for energy in cells

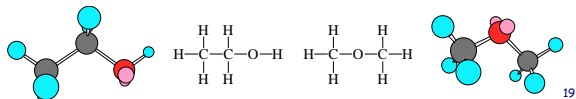


18

### Structural Isomers

- Even compounds that have the *same* molecular formula can have the atoms connected in a *different order* — these are **structural isomers**.

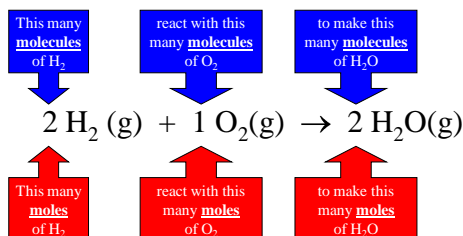
	Ethanol	Dimethyl ether
<b>Molecular Formula</b>	C <sub>2</sub> H <sub>6</sub> O	C <sub>2</sub> H <sub>6</sub> O
<b>Molar Mass (g/mol)</b>	46.07	46.07
<b>Appearance</b>	Colorless liquid	Colorless gas
<b>Melting point</b>	-117 C	-139 C
<b>Boiling point</b>	78.5 C	-25 C
<b>Density (at 20°C)</b>	0.789 g/mL	0.00195 g/mL
<b>Function</b>	Intoxicant	Refrigerant



# Stoichiometry

## From Molecules to Moles to Grams

- Since moles combine in the same ratio that atoms or molecules do, **the coefficients in a balanced chemical reaction specify the relative amounts in moles of each of the substances involved in the reaction.**



1

2

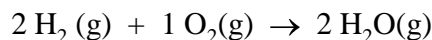
## Stoichiometry: Chemical Arithmetic

Greek: *stoicheion* + *metron*  
element or part + measure

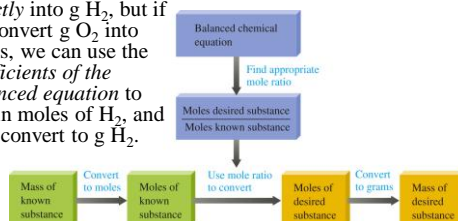
- Stoichiometry** is the study of the numerical relationships in chemical formulas and reactions.
  - Knowing the stoichiometry of a formula allows us to relate moles and grams for *particular* reactants or products (e.g., that 1 mole of H<sub>2</sub>O weighs 18.02 g).
  - Knowing the stoichiometry of a reaction allows us to relate amounts of different substances *to each other*, using the **mole ratios** in the balanced equation, and allows us to predict how much of the products will be formed or how much of the reactants will be needed.

3

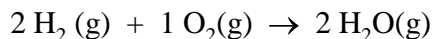
## Reaction Stoichiometry: An Example



- Suppose we have 32.0 g of O<sub>2</sub>. How many grams of H<sub>2</sub> will be needed for this reaction? How many grams of H<sub>2</sub>O will be produced?
- We can't convert g O<sub>2</sub> directly into g H<sub>2</sub>, but if we convert g O<sub>2</sub> into moles, we can use the *coefficients of the balanced equation* to obtain moles of H<sub>2</sub>, and then convert to g H<sub>2</sub>.



## Reaction Stoichiometry: An Example



Convert g O<sub>2</sub> to mol O<sub>2</sub>:

$$32.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 1.00 \text{ mol O}_2$$

Convert mol O<sub>2</sub> to mol H<sub>2</sub>:

$$1.00 \text{ mol O}_2 \times \frac{2 \text{ mol H}_2}{1 \text{ mol O}_2} = 2.00 \text{ mol H}_2$$

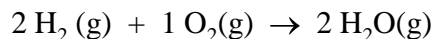
↖ coefficient of what we're interested in  
↗ coefficient of what we're canceling out

Convert mol H<sub>2</sub> to g H<sub>2</sub>:

$$2.00 \text{ mol H}_2 \times \frac{2.02 \text{ g H}_2}{1 \text{ mol H}_2} = 4.04 \text{ g H}_2$$

5

## Reaction Stoichiometry: An Example



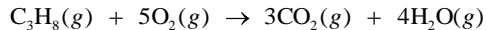
Or we can put everything together:

$$32.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol H}_2}{1 \text{ mol O}_2} \times \frac{2.02 \text{ g H}_2}{1 \text{ mol H}_2} = 4.04 \text{ g H}_2$$

How many grams of H<sub>2</sub>O will be formed?

$$32.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 36.0 \text{ g H}_2\text{O}$$

6

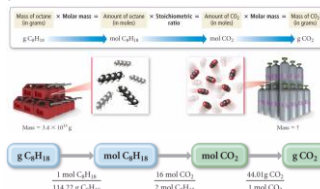
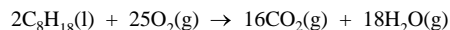
**Examples: Reaction Stoichiometry**

- How many moles of  $\text{CO}_2$  can we make from 2.0 moles of  $\text{C}_3\text{H}_8$ ?
- How many moles of  $\text{H}_2\text{O}$  can we make from 2.0 moles of  $\text{C}_3\text{H}_8$ ?
- How many moles of  $\text{O}_2$  are needed to react with 2.0 moles of  $\text{C}_3\text{H}_8$ ?
- How many moles of  $\text{CO}_2$  can be produced from 3.5 mol  $\text{O}_2$ ?
- How many grams of  $\text{CO}_2$  are produced from 50.0 g of  $\text{C}_3\text{H}_8$ ?

7

**Examples: Reaction Stoichiometry**

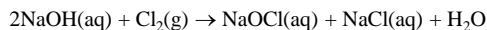
- In 2004, the world burned  $3.0 \times 10^{10}$  barrels of petroleum, roughly equivalent to  $3.4 \times 10^{15}$  g of gasoline ( $\text{C}_8\text{H}_{18}$ ). How much  $\text{CO}_2$  is released into the atmosphere from the combustion of this much gasoline?

Answer:  $1.0 \times 10^{16}$  g  $\text{CO}_2$ 

8

**Examples: Reaction Stoichiometry**

- Aqueous sodium hypochlorite ( $\text{NaOCl}$ ), best known as household bleach, is prepared by reaction of sodium hydroxide with chlorine:



How many grams of  $\text{NaOH}$  are needed to react with 25.0 g of  $\text{Cl}_2$ ?

Answer: 28.2 g  $\text{NaOH}$ 

9

**Examples: Molecule Stoichiometry**

- How many grams of  $\text{Cl}$  atoms are needed to combine with 24.4 g of  $\text{Si}$  atoms to make silicon tetrachloride,  $\text{SiCl}_4$ ?

Answer: 123 g  $\text{Cl}$ 

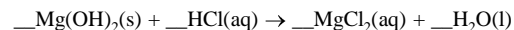
10

**Examples: Molecule Stoichiometry**

- Baking soda,  $\text{NaHCO}_3$ , is often used as an antacid. It neutralizes excess hydrochloric acid,  $\text{HCl}$ , secreted by the stomach:



Milk of magnesia is an aqueous suspension of magnesium hydroxide,  $\text{Mg}(\text{OH})_2$ , which is also used as an antacid:



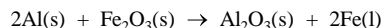
Which is the more effective antacid per gram,  $\text{NaHCO}_3$  or  $\text{Mg}(\text{OH})_2$ ? [Example 9.8]

Answer:  $\text{Mg}(\text{OH})_2$  is the more effective antacid.

11

**Examples: Reaction Stoichiometry**

- One of the most spectacular reactions of aluminum, the thermite reaction, is with iron(III) oxide,  $\text{Fe}_2\text{O}_3$ , by which metallic iron is made. So much heat is generated that the iron forms in the liquid state. The equation is



A certain welding operation, used over and over, requires that each time at least 86.0 g of  $\text{Fe}$  be produced. (a) What is the minimum mass in grams of  $\text{Fe}_2\text{O}_3$  that must be used for each operation? (b) How many grams of aluminum are also needed?

Answer: (a) 123 g  $\text{Fe}_2\text{O}_3$ ; (b) 41.5 g  $\text{Al}$ 

12

### Yields of Chemical Reactions

- In the examples we've seen, we have assumed that all of the reactions "go to completion" — that is, that *all* reactant molecules are converted into products. In real life, some product is almost always lost due to small amounts of contamination present in the glassware, impurities in the reactants, incomplete reactions, reactants evaporating into the air, too much heating, insufficient heating, klutzes, gremlins, evil spirits, evil co-workers, or *side reactions* that form other products.
- The **theoretical yield** is the amount that would be obtained if the reaction goes to completion (i.e., the *maximum amount* that could be made).
- The **actual yield** of a reaction is the amount that is actually obtained. (You could've guessed that.)

13

### Yields of Chemical Reactions

- The **percent yield (% yield)** is the actual yield expressed as a percentage of the theoretical yield:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

- The reactant we run out of first, which limits the yield of the entire reaction, is the **limiting reactant** (or **limiting reagent**).
- The **excess reactant** is any reactant that is present in a larger amount than what is required to react completely with the limiting reactant.

14

### Examples: Percent Yield

7. Methyl *tert*-butyl ether (MTBE, C<sub>5</sub>H<sub>12</sub>O), a substance used as an octane booster in gasoline, can be made by reaction of isobutylene (C<sub>4</sub>H<sub>8</sub>), with methanol (CH<sub>3</sub>OH). What is the percent yield of the reaction if 32.8 g of MTBE is obtained from reaction of 26.3 g of isobutylene with sufficient methanol?



**Answer: 79.4%**

15

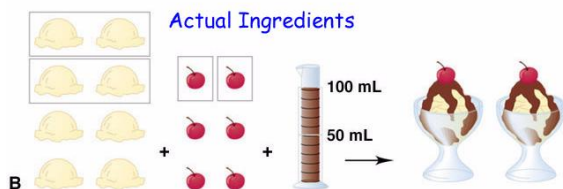
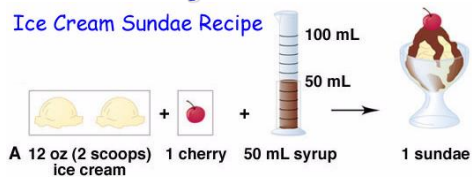
16

# Limiting Reactants

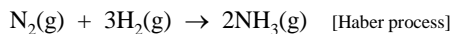
1

## Limiting Reactants and Sundaes

### Ice Cream Sundae Recipe



## Limiting Reactants



- Suppose we mix 1.00 mol of  $\text{N}_2$  and 5.00 mol of  $\text{H}_2$ . What is the maximum amount of  $\text{NH}_3$  that can be produced? How much  $\text{H}_2$  will be left over?
- Now suppose we mix 2.15 mol of  $\text{N}_2$  and 6.15 mol of  $\text{H}_2$ . What is the theoretical yield of  $\text{NH}_3$ ?

Assuming the  $\text{N}_2$  reacts completely, how much  $\text{NH}_3$  can be made?

$$2.15 \text{ mol N}_2 \times \frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2} = 4.30 \text{ mol NH}_3$$

Assuming the  $\text{H}_2$  reacts completely, how much  $\text{NH}_3$  can be made?

$$6.15 \text{ mol H}_2 \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} = 4.10 \text{ mol NH}_3$$

$\text{H}_2$  is the limiting reactant; the theoretical yield of  $\text{NH}_3$  is 4.10 mol 5

## Limiting Reactants

- When we are given a reaction between two or more reactants, one may be completely consumed before the other(s). The reaction must **stop** at this point, leaving us with the remaining reactants *in excess*.
- The amount of this reactant, then, determines the maximum amount of the product(s) that can form, and is known as the **limiting reactant**.
- For example, suppose we were making standard 4-door cars, and we had the following (incomplete) list of “ingredients.” How many cars could we make?

4 engines	4 drivers' seats
4 steering wheels	4 rear-view mirrors
15 doors	8 windshield wipers
8 headlights	11 wheels

2

## Limiting Reactants and Pizza

**Pizza recipe:**

**1 crust + 5 oz. tomato sauce + 2 cups cheese → 1 pizza**

If we have 4 crusts, 10 cups of cheese, and 15 oz. of tomato sauce, how many pizzas can we make?

$$4 \text{ crusts} \times \frac{1 \text{ pizza}}{1 \text{ crust}} = 4 \text{ pizzas}$$

$$10 \text{ cups cheese} \times \frac{1 \text{ pizza}}{2 \text{ cups cheese}} = 5 \text{ pizzas}$$

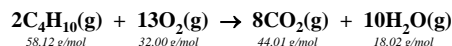
$$15 \text{ ounces tomato sauce} \times \frac{1 \text{ pizza}}{5 \text{ ounces tomato sauce}} = 3 \text{ pizzas}$$

Tomato sauce is the **limiting reagent**, and the **theoretical yield** is 3 pizzas.

4

## Examples: Limiting Reactants

- Butane,  $\text{C}_4\text{H}_{10}$ , undergoes combustion with oxygen,  $\text{O}_2$ , to form carbon dioxide and water:



If 100. g of  $\text{C}_4\text{H}_{10}$  and 100. g of  $\text{O}_2$  are mixed,

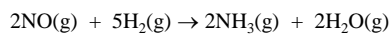
- Which of the two reactants is the limiting reagent, and how many grams of  $\text{CO}_2$  will be formed?
- How many grams of  $\text{H}_2\text{O}$  will be formed?
- How many grams of excess reagent are left over?
- If the actual yield of  $\text{CO}_2$  had been 75.0 g, what would be the percent yield of the reaction?

**Answer:** (a)  $\text{O}_2$  limiting; 84.6 g  $\text{CO}_2$ ; (b) 43.3 g  $\text{H}_2\text{O}$ ; (c) 72 g  $\text{C}_4\text{H}_{10}$ ; (d) 88.6%

6

**Examples: Limiting Reactants**

2. Ammonia,  $\text{NH}_3$ , can be synthesized by the following reaction:



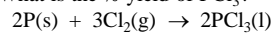
Starting with 86.3 g NO and 25.6 g  $\text{H}_2$ , find the theoretical yield of ammonia in grams. (Ex. 4.3)

**Answer: NO limiting; 49.0 g  $\text{NH}_3$ .**

7

**Examples: Limiting Reactants**

3. In a synthesis of phosphorus trichloride, a chemist mixed 12.0 g P with 35.0 g  $\text{Cl}_2$ ; she obtained 42.4 g of  $\text{PCl}_3$ . What is the % yield of  $\text{PCl}_3$ ?



**Answer: 93.8%**

8