Lecture Presentation

Nomenclature and Stoichiometry

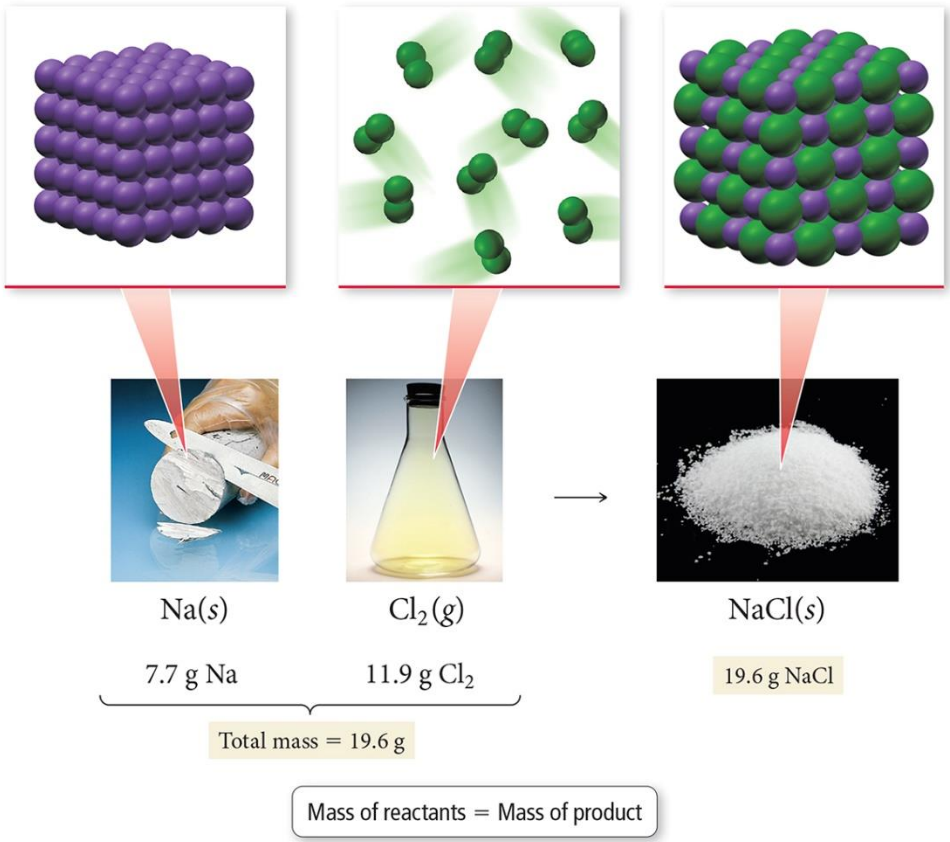
The Law of Conservation of Mass

- Antoine Lavoisier formulated the **law of conservation of mass**, which states that ***in a chemical reaction, matter is neither created nor destroyed.***

- Hence, when a chemical reaction occurs, the total mass of the substances involved in the reaction does not change.

The Law of Conservation of Mass

- This law is consistent with the idea that matter is composed of small, indestructible particles.



The Neutrons

- The mass of a neutron is similar to that of a proton.
- However, a neutron has no electrical charge.
  – The helium atom is four times as massive as the hydrogen atom because it contains two protons and *two neutrons.*
- Hydrogen, on the other hand, contains only one proton and no neutrons.

Classification of Elements

- Elements in the periodic table are classified as the following:
  - Metals
  - Nonmetals
  - Metalloids

Metals

- **Metals**, on the lower-left side and middle of the periodic table, share some common properties.
  - Good conductors of heat and electricity
  - Can be pounded into flat sheets (malleability)
  - Can be drawn into wires (ductility)
  - Often shiny
  - Tend to lose electrons when they undergo chemical changes
- Chromium, copper, strontium, and lead are typical metals.

Nonmetals

- **Nonmetals** lie on the upper-right side of the periodic table.
- There are a total of 17 nonmetals:
  - Five are solids at room temperature (C, P, S, Se, and I).
  - One is a liquid at room temperature (Br).
  - Eleven are gases at room temperature (H, He, N, O, F, Ne, Cl, Ar, Kr, Xe, and Rn).

Nonmetals

- Nonmetals as a whole tend to have these properties:
  - Poor conductors of heat and electricity
  - Not ductile and not malleable
  - Gain electrons when they undergo chemical changes
- Oxygen, carbon, sulfur, bromine, and iodine are nonmetals.

Metalloids

- Metalloids are sometimes called *semimetals*.
- They are elements that lie along the zigzag diagonal line that divides metals and nonmetals.
- They exhibit mixed properties.
- Several metalloids are also classified as **semiconductors** because of their intermediate (and highly temperature-dependent) electrical conductivity.

Ions and the Periodic Table

- A main-group metal tends to lose electrons, forming a cation with the same number of electrons as the nearest noble gas.

- A main-group nonmetal tends to gain electrons, forming an anion with the same number of electrons as the nearest noble gas.

Ions and the Periodic Table

- In general, the alkali metals (group 1A) have a tendency to lose one electron and form 1+ ions.
- The alkaline earth metals (group 2A) tend to lose two electrons and form 2+ ions.
- The halogens (group 7A) tend to gain one electron and form 1– ions.
- The oxygen family nonmetals (group 6A) tend to gain two electrons and form 2– ions.

Ions and the Periodic Table

- For the main-group elements that form cations with a predictable charge, the charge is equal to the group number.
- For main-group elements that form anions with a predictable charge, the charge is equal to the group number minus eight.
- Transition elements may form various different ions with different charges.

Ions and the Periodic Table

Elements That Form Ions with Predictable Charges

1A	2A	Transition metals	3A	4A	5A	6A	7A	8A
H^+							H^-	Noble Gases
Li^+					N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}		Al^{3+}			S^{2-}	Cl^-	
K^+	Ca^{2+}					Se^{2-}	Br^-	
Rb^+	Sr^{2+}					Te^{2-}	I^-	
Cs^+	Ba^{2+}							

Atomic Mass: The Average Mass of an Element's Atoms

- Atomic mass is sometimes called *atomic weight* or *standard atomic weight.*
- The atomic mass of each element is directly beneath the element's symbol in the periodic table.
- It represents the average mass of the isotopes that compose that element, *weighted according to the natural abundance of each isotope*.

Example: Atomic Mass

- Naturally occurring chlorine consists of 75.77% chlorine-35 atoms (mass 34.97 amu) and 24.23% chlorine-37 atoms (mass 36.97 amu). Calculate its atomic mass.

- Solution:
  – Convert the percent abundance to decimal form and multiply each with its isotopic mass.

    Cl-37 = 0.2423(36.97 amu) = 8.9578 amu

    Cl-35 = 0.7577(34.97 amu) = 26.4968 amu

    Atomic Mass Cl = 8.9578 + 26.4968 = 35.45 amu

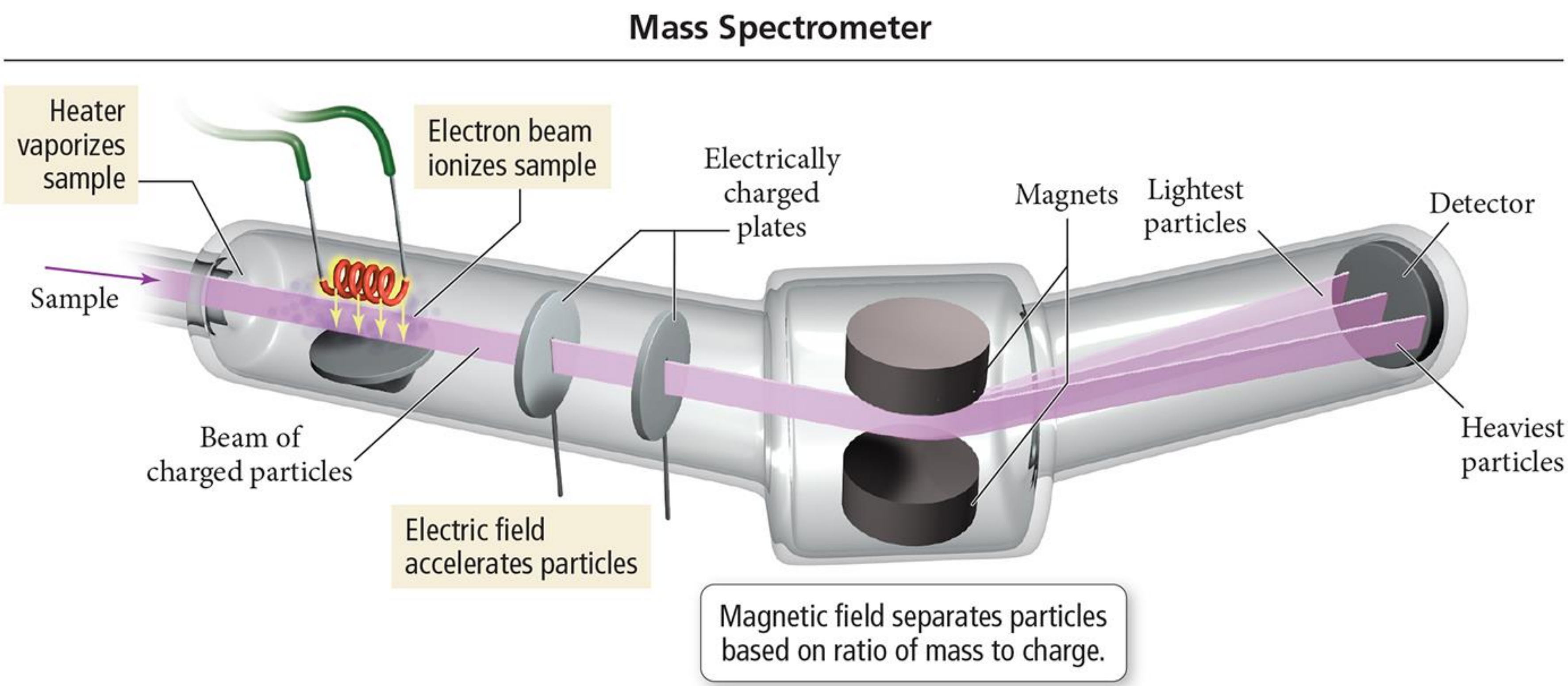
Atomic Mass

17
Cl
35.45
chlorine

- In general, we calculate the atomic mass with the following equation:

$$\begin{aligned}\text{Atomic mass} &= \sum_n (\text{fraction of isotope } n) \times (\text{mass of isotope } n) \\ &= (\text{fraction of isotope 1} \times \text{mass of isotope 1}) \\ &+ (\text{fraction of isotope 2} \times \text{mass of isotope 2}) \\ &+ (\text{fraction of isotope 3} \times \text{mass of isotope 3}) + \ldots\end{aligned}$$

Mass Spectrometry



Molar Mass: Counting Atoms by Weighing Them

- As chemists, we often need to know the number of atoms in a sample of a given mass. Why? *Because chemical processes happen between particles*.

- Therefore, if we want to know the number of atoms in anything of ordinary size, we count them by weighing.

The Mole: A Chemist's "Dozen"

- When we count large numbers of objects, we often use units such as
  - 1 dozen objects = 12 objects.
  - 1 gross objects = 144 objects.
- The chemist's "dozen" is the **mole** (abbreviated mol). A mole is the measure of material containing 6.02214×10^{23} particles:

  1 mole = 6.02214×10^{23} particles
- This number is **Avogadro's number**.

The Mole

- First thing to understand about the mole is that it can specify Avogadro’s number of anything.
- For example, 1 mol of marbles corresponds to 6.02214×10^{23} marbles.
- 1 mol of sand grains corresponds to 6.02214×10^{23} sand grains.
- *One mole of anything is* 6.02214×10^{23} *units of that thing.*

The Mole

- The second, and more fundamental, thing to understand about the mole is how it gets its specific value.

- **The value of the mole is equal to the number of atoms in exactly 12 grams of pure C-12**.

- **12 g C = 1 mol C atoms = 6.022×10^{23} C atoms**

Converting between Number of Moles and Number of Atoms

- Converting between number of moles and number of atoms is similar to converting between dozens of eggs and number of eggs.
- For atoms, you use the conversion factor 1 mol atoms = 6.022×10^{23} atoms.
- The conversion factors take the following forms:

$$\frac{1 \text{ mol atoms}}{6.022 \times 10^{23} \text{ atoms}} \quad \text{or} \quad \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol atoms}}$$

Converting between Mass and Amount (Number of Moles)

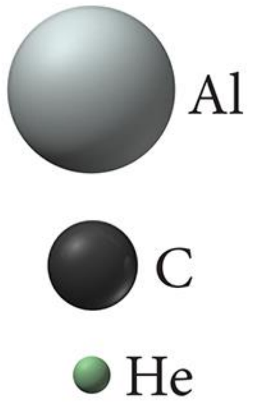
- To count atoms by weighing them, we need one other conversion factor—the mass of 1 mol of atoms.
- The mass of 1 mol of atoms of an element is the **molar mass**.
- **An element's molar mass in grams per mole is numerically equal to the element's atomic mass in atomic mass units (amu).**

Converting between Mass and Moles

$$26.98\text{ g aluminum} = 1\text{ mol aluminum} = 6.022 \times 10^{23}\text{ Al atoms}$$

$$12.01\text{ g carbon} = 1\text{ mol carbon} = 6.022 \times 10^{23}\text{ C atoms}$$

$$4.003\text{ g helium} = 1\text{ mol helium} = 6.022 \times 10^{23}\text{ He atoms}$$



- The lighter the atom, the less mass in 1 mol of atoms.

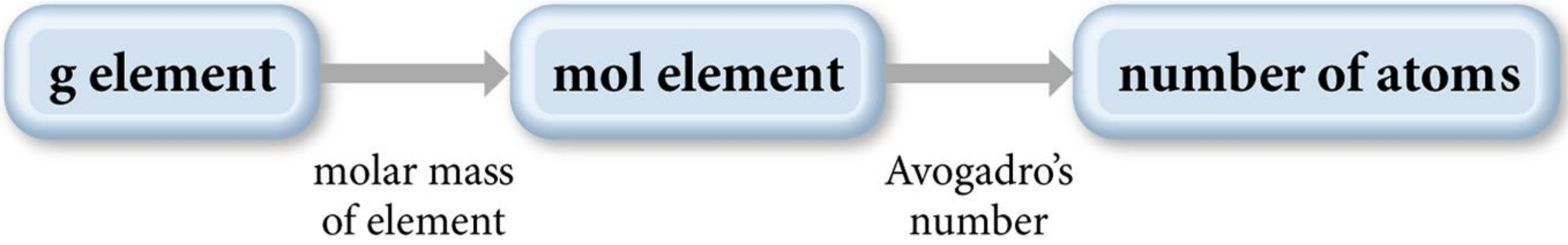
Converting between Mass and Moles

- The molar mass of any element is the conversion factor between the mass (in grams) of that element and the amount (in moles) of that element. For carbon,

$$12.01 \text{ g C} = 1 \text{ mol C} \quad \text{or} \quad \frac{12.01 \text{ g C}}{\text{mol C}} \quad \text{or} \quad \frac{1 \text{ mol C}}{12.01 \text{ g C}}$$

Conceptual Plan

- We now have all the tools to count the number of atoms in a sample of an element by weighing it.
  - First, we obtain the mass of the sample.
  - Then, we convert it to the amount in moles using the element's molar mass.
  - Finally, we convert it to the number of atoms using Avogadro's number.
- The conceptual plan for these kinds of calculations takes the following form:



Representing Compounds: Chemical Formulas and Molecular Models

- A compound's **chemical formula** indicates the elements present in the compound and the relative number of atoms or ions of each.
  - Water is represented as H_2O.
  - Sodium Chloride is represented as NaCl.
  - Carbon dioxide is represented as CO_2.
  - Carbon tetrachloride is represented as CCl_4.

Types of Chemical Formulas

- Chemical formulas can generally be categorized into three different types:
  - empirical formulas,
  - molecular formulas, and
  - structural formulas.

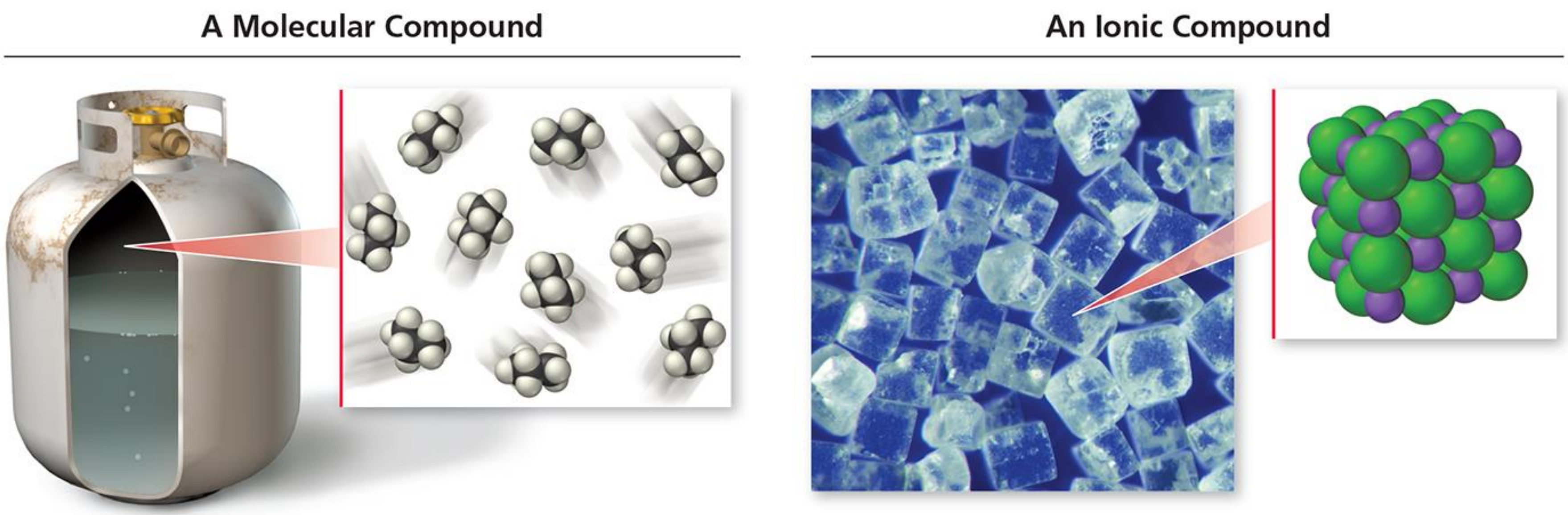
Molecular Compounds

- **Molecular compounds** are usually composed of two or more covalently bonded nonmetals.
- The basic units of molecular compounds are **molecules** composed of the constituent atoms.
  - Water is composed of H_2O molecules.
  - Dry ice is composed of CO_2 molecules.
  - Propane (often used as a fuel for grills) is composed of C_3H_8 molecules.

Ionic Compounds

- **Ionic compounds** are composed of cations (usually a metal) and anions (usually one or more nonmetals) bound together by ionic bonds.
- The basic unit of an ionic compound is the **formula unit**, the smallest, electrically neutral collection of ions.
- Table salt is an ionic compound with the formula unit NaCl, which is composed of Na^+ and Cl^- ions in a one-to-one ratio.

Molecular and Ionic Compounds



(a) (b)

Polyatomic Ions

- Many common ionic compounds contain ions that are themselves composed of a group of covalently bonded atoms with an overall charge.
- This group of charged species is called polyatomic ions.
  - $NaNO_3$ contains Na^+ and NO_3^-.
  - $CaCO_3$ contains Ca^{2+} and CO_3^{2-}.
  - $Mg(ClO_3)_2$ contains Mg^{2+} and ClO_3^-.

Ionic Compounds: Formulas and Names

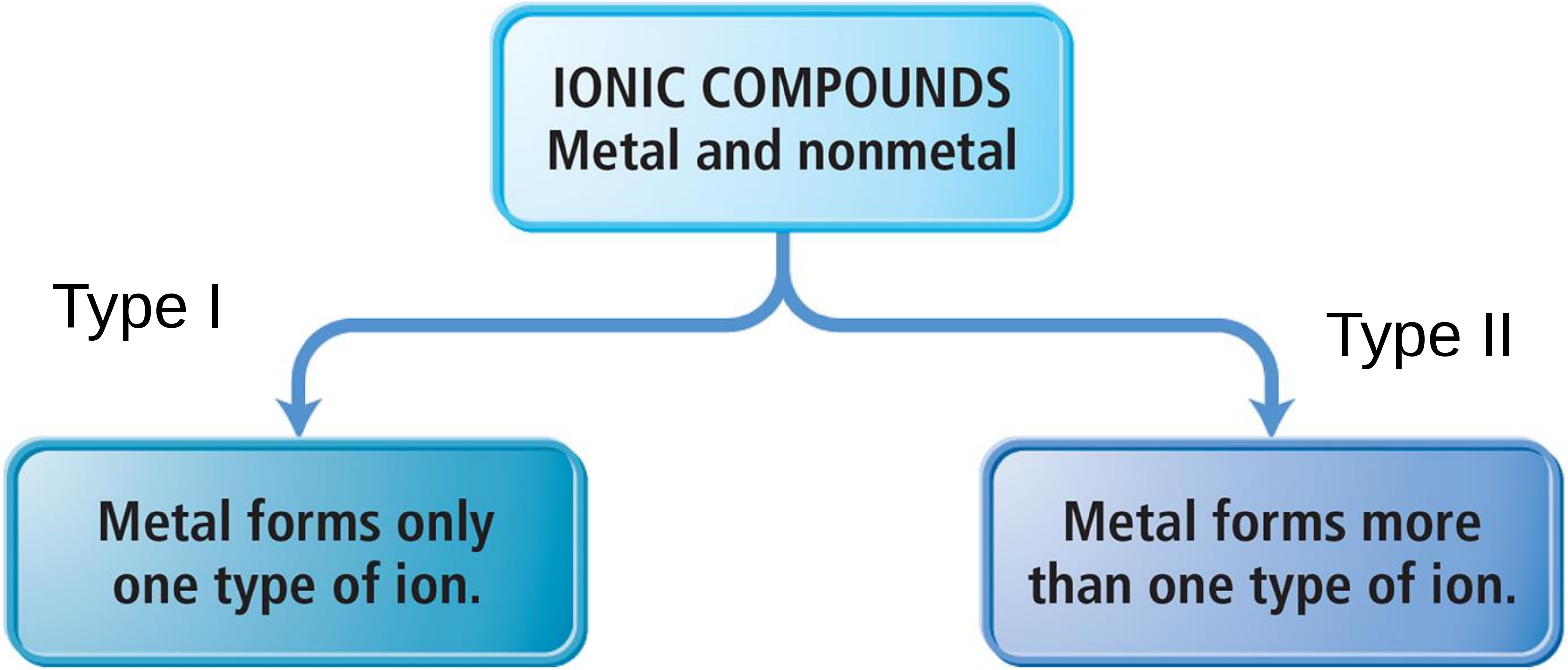
- ***Summarizing Ionic Compound Formulas***
  - Ionic compounds always contain positive and negative ions.
  - In a chemical formula, the sum of the charges of the cations must equal the sum of the charges of the anions.
  - The formula of an ionic compound reflects the smallest whole-number ratio of ions.

Ionic Compounds: Formulas and Names

- The charges of the representative elements can be predicted from their group numbers.
- The representative elements form only one type of charge.
- Transition metals tend to form multiple types of charges.
- Hence, their charges cannot be predicted as in the case of most representative elements.

Naming Ionic Compounds

- Ionic compounds can be categorized into two types, depending on the metal in the compound.



Naming Type I Ionic Compounds

- Type I ionic compounds contain a metal whose charge is invariant from one compound to another when bonded with a nonmetal anion.
- The metal ion always has the same charge.

Metals Whose Charge Are Invariant from One Compound to Another

Metals Whose Charge Is Invariant from One Compound to Another

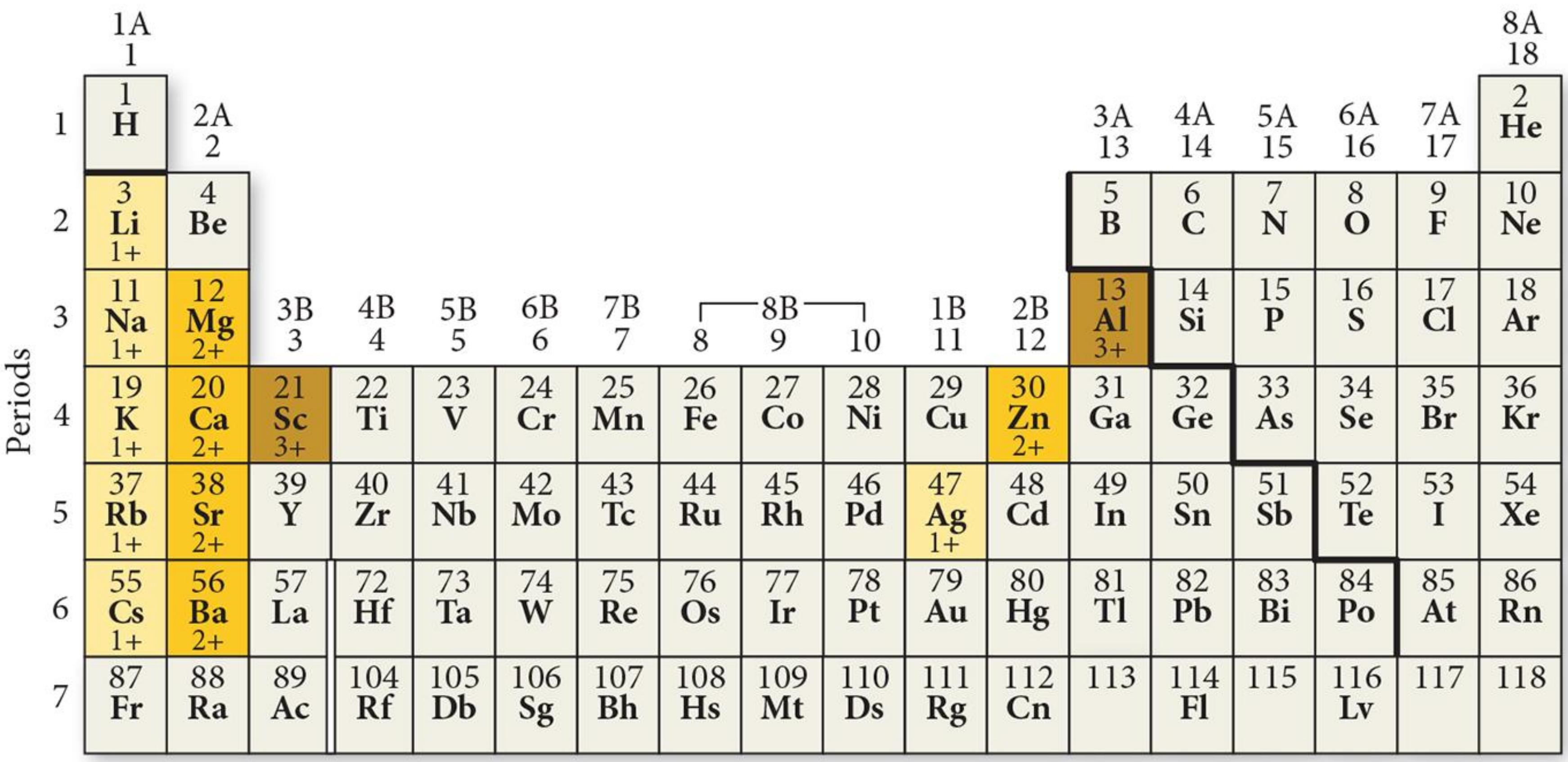


TABLE 3.2 Some Common Monoatomic Anions

Nonmetal	Symbol for Ion	Base Name	Anion Name
Fluorine	F^{-}	fluor	Fluoride
Chlorine	Cl^{-}	chlor	Chloride
Bromine	Br^{-}	brom	Bromide
Iodine	I^{-}	iod	Iodide
Oxygen	O^{2-}	ox	Oxide
Sulfur	S^{2-}	sulf	Sulfide
Nitrogen	N^{3-}	nitr	Nitride
Phosphorus	P^{3-}	phosph	Phosphide

Naming Binary Ionic Compounds of Type I Cations

- **Binary compounds** contain only two different elements. The names of binary ionic compounds take the following form:

name of cation (metal)

base name of anion (nonmetal) + *-ide*

Examples: Type I Binary Ionic Compounds

- The name for KCl consists of the name of the cation, *potassium*, followed by the base name of the anion, *chlor*, with the ending *-ide*.
  - KCl is *potassium chloride*.
- The name for CaO consists of the name of the cation, *calcium*, followed by the base name of the anion, *ox*, with the ending *-ide.*
  - CaO is *calcium oxide*.

Naming Type II Ionic Compounds

- The second type of ionic compound contains a metal that can form more than one kind of cation, depending on the compound, bonded to a nonmetal anion.
- The metal's charge must be specified for a given compound.
- The proportion of metal cation to nonmetal anion helps us determine the charge on the metal ion.

Type II Ionic Compounds

- Iron, for instance, forms a 2+ cation in some of its compounds and a 3+ cation in others.
- Metals of this type are often *transition metals*.
  - FeS: Here, iron is +2 cation (Fe^{2+}).
  - Fe_2S_3: Here, iron is +3 cation (Fe^{3+}).
  - Cu_2O: Here, copper is +1 cation (Cu^{+}).
  - CuO: Here, copper is +2 cation (Cu^{2+}).
- Some main group metals, such as Pb, Ti, and Sn, form more than one type of cation.

Naming Type II Binary Ionic Compounds

- The full name of compounds containing metals that form more than one kind of cation have the following form:

name of cation (metal)	(charge of cation (metal) in Roman numerals in parentheses)	base name of anion (nonmetal) + *-ide*

- The charge of the metal cation can be determined by inference from the sum of the charges of the nonmetal.

Naming Type II Binary Ionic Compounds

- For these types of metals, the name of the cation is followed by a roman numeral (in parentheses) that indicates the charge of the metal in that particular compound.
  - For example, we distinguish between Fe^{2+} and Fe^{3+} as follows:
    - Fe^{2+} Iron(II)
    - Fe^{3+} Iron(III)

TABLE 3.3 Some Metals That Form Cations with Different Charges

Metal	Ion	Name	Older Name*
Chromium	Cr^{2+}	Chromium(II)	Chromous
	Cr^{3+}	Chromium(III)	Chromic
Iron	Fe^{2+}	Iron(II)	Ferrous
	Fe^{3+}	Iron(III)	Ferric
Cobalt	Co^{2+}	Cobalt(II)	Cobaltous
	Co^{3+}	Cobalt(III)	Cobaltic
Copper	Cu^{+}	Copper(I)	Cuprous
	Cu^{2+}	Copper(II)	Cupric
Tin	Sn^{2+}	Tin(II)	Stannous
	Sn^{4+}	Tin(IV)	Stannic
Mercury	Hg_2^{2+}	Mercury(I)	Mercurous
	Hg^{2+}	Mercury(II)	Mercuric
Lead	Pb^{2+}	Lead(II)	Plumbous
	Pb^{4+}	Lead(IV)	Plumbic

*An older naming system substitutes the names found in this column for the name of the metal and its charge. Under this system, chromium(II) oxide is named chromous oxide. Additionally, the suffix *-ous* indicates the ion with the lesser charge, and *-ic* indicates the ion with the greater charge. We will *not* use the older system in this text.

Example: Type II Binary Ionic Compounds

- To name $CrBr_3$,determine the charge on the chromium.
  - Total charge on cation + total anion charge = 0.
  - Cr charge + 3($Br^–$ charge) = 0.
  - Since each Br has a –1 charge, then:
    - Cr charge + 3(–1) = 0.
    - Cr charge –3 = 0.
    - Cr = +3.
  - Hence, the cation Cr^{3+} is called chromium(III), while $Br^–$ is called bromide.
- Therefore, $CrBr_3$ is **chromium(III) bromide.**

Naming Ionic Compounds Containing Polyatomic Ions

- We name ionic compounds that contain a polyatomic ion in the same way as other ionic compounds, except that we use the name of the polyatomic ion whenever it occurs.
- For example, $NaNO_2$ is named according to its cation, Na^+, *sodium*, and its polyatomic anion, NO_2^-, *nitrite*.
- Hence, $NaNO_2$ is *sodium nitrite*.

TABLE 3.4 Some Common Polyatomic Ions

Name	Formula	Name	Formula
Acetate	$C_2H_3O_2^-$	Hypochlorite	ClO^-
Carbonate	CO_3^{2-}	Chlorite	ClO_2^-
Hydrogen carbonate (or bicarbonate)	HCO_3^-	Chlorate	ClO_3^-
Hydroxide	OH^-	Perchlorate	ClO_4^-
Nitrite	NO_2^-	Permanganate	MnO_4^-
Nitrate	NO_3^-	Sulfite	SO_3^{2-}
Chromate	CrO_4^{2-}	Hydrogen sulfite (or bisulfite)	HSO_3^-
Dichromate	$Cr_2O_7^{2-}$	Sulfate	SO_4^{2-}
Phosphate	PO_4^{3-}	Hydrogen sulfate (or bisulfate)	HSO_4^-
Hydrogen phosphate	HPO_4^{2-}	Cyanide	CN^-
Dihydrogen phosphate	$H_2PO_4^-$	Peroxide	O_2^{2-}
Ammonium	NH_4^+		

Oxyanions

- Most polyatomic ions are **oxyanions**, anions containing oxygen and another element.
- Notice that when a series of oxyanions contains different numbers of oxygen atoms, they are named according to the number of oxygen atoms in the ion.
- If there are two ions in the series,
  - the one with more oxygen atoms has the ending *-ate*, and
  - the one with fewer has the ending *-ite*.
- For example,
  - NO_3^- is ***nitrate***.
  - SO_4^{2-} is ***sulfate***.
  - NO_2^- is ***nitrite***.
  - SO_3^{2-} is ***sulfite***.

Oxyanions

- If there are more than two ions in the series, then the prefixes *hypo-*, meaning *less than*, and *per-*, meaning *more than*, are used.

ClO^-	*hypo*chlor*ite*	BrO^-	*hypo*brom*ite*
ClO_2^-	chlor*ite*	BrO_2^-	brom*ite*
ClO_3^-	chlor*ate*	BrO_3^-	brom*ate*
ClO_4^-	*per*chlor*ate*	BrO_4^-	*per*brom*ate*

Hydrated Ionic Compounds

- **Hydrates** are ionic compounds containing a specific number of water molecules associated with each formula unit.
  - For example, the formula for epsom salts is $MgSO_4 \cdot 7H_2O$.
  - Its systematic name is magnesium sulfate heptahydrate.
  - $CoCl_2 \cdot 6H_2O$ is cobalt(II) chloride hexahydrate.

Common Hydrate Prefixes

- Common hydrate prefixes

hemi = ½	tri = 3	hexa = 6
mono = 1	tetra = 4	hepta = 7
di = 2	penta = 5	octa = 8

- Other common hydrated ionic compounds and their names are as follows:
  - $CaSO_4 \cdot \frac{1}{2} H_2O$ is called calcium sulfate hemihydrate.
  - $BaCl_2 \cdot 6H_2O$ is called barium chloride hexahydrate.
  - $CuSO_4 \cdot 6H_2O$ is called copper sulfate hexahydrate.

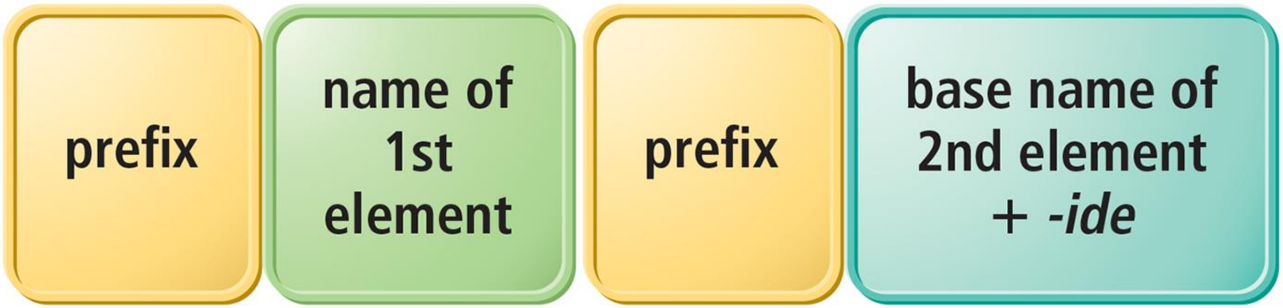
Molecular Compounds: Formulas and Names

- The formula for a molecular compound *cannot* readily be determined from its constituent elements because the same combination of elements may form many different molecular compounds, each with a different formula.
  - Nitrogen and oxygen form all of the following unique molecular compounds: NO, NO_2, N_2O, N_2O_3, N_2O_4, and N_2O_5.

Molecular Compounds: Formulas and Names

- *Molecular compounds are composed of two or more nonmetals*.
- Generally, write the name of the element with the smallest group number first.
- If the two elements lie in the same group, then write the element with the greatest row number first.
  - The prefixes given to each element indicate the number of atoms present.

Binary Molecular Compounds



- These prefixes are the same as those used in naming hydrates:

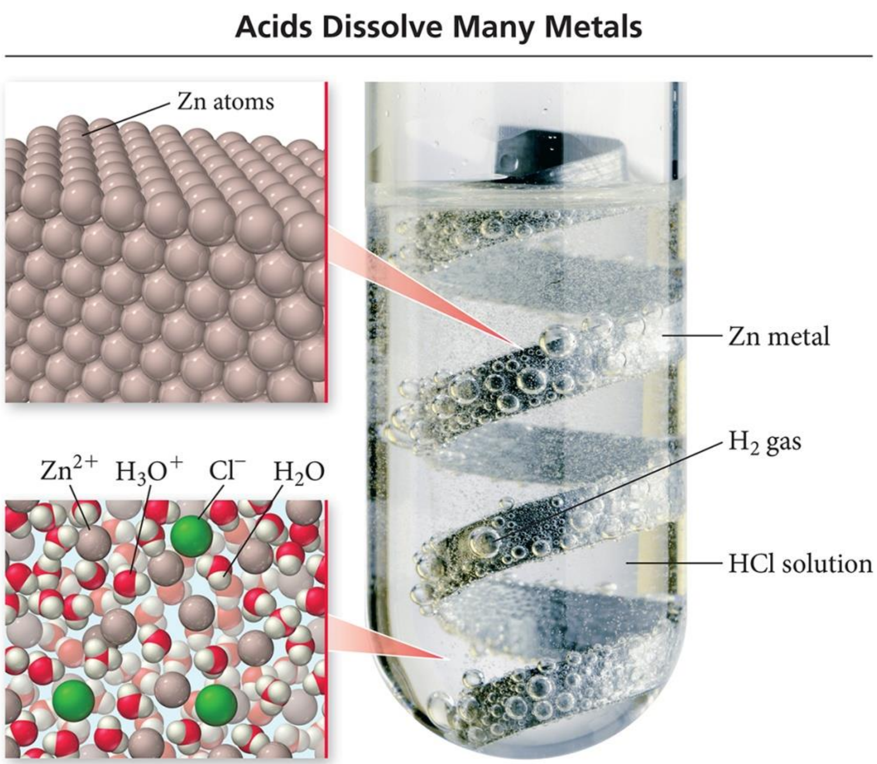
mono = 1	hexa = 6
di = 2	hepta = 7
tri = 3	octa = 8
tetra = 4	nona = 9
penta = 5	deca = 10

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

Acids

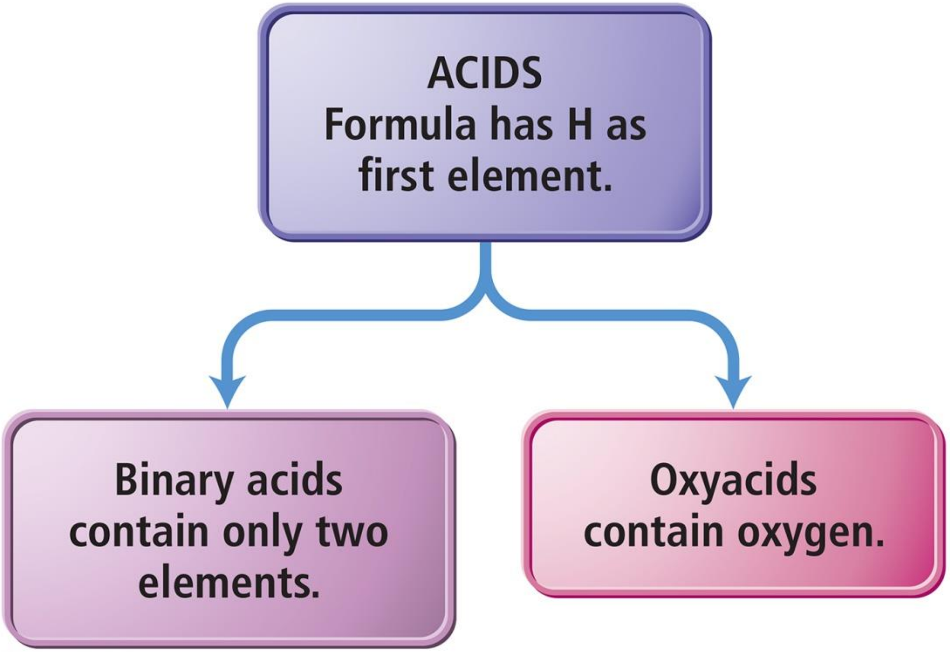
- Acids are molecular compounds that release hydrogen ions (H^+) when dissolved in water.
- Acids are composed of hydrogen, usually written first in their formulas, and one or more nonmetals, written second.

Acids



- Sour taste
- Dissolve many metals
  - such as Zn, Fe, and Mg; but not Au, Ag, or Pt
- Formulas generally start with H,
  - e.g., HCl, H_2SO_4
- HCl is a molecular compound that, when dissolved in water, forms $H^+_{(aq)}$ and $Cl^-_{(aq)}$ ions, where *aqueous* (*aq*) means *dissolved in water*.

Acids



- Binary acids have H^+ cation and nonmetal anion.
- Oxyacids have H^+ cation and polyatomic anion.

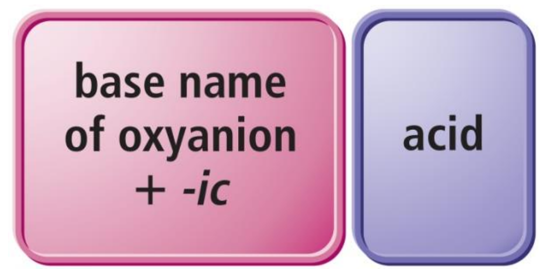
Naming Binary Acids

- Write a ***hydro-*** prefix.
- Follow with the nonmetal base name.
- Add *–ic*.
- Write the word **acid** at the end of the name.

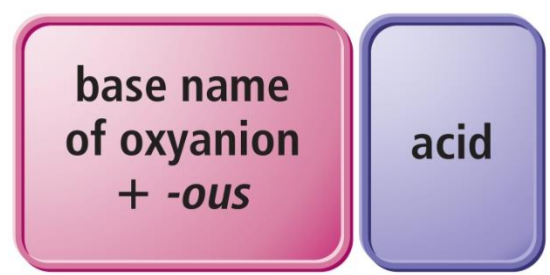
Naming Oxyacids

- If the polyatomic ion name ends in *–ate,* change ending to *–ic*.
- If the polyatomic ion name ends in *–ite*, change ending to *–ous*.
- Write word **acid** at the end of all names.

oxyanions ending with -*ate*



oxyanions ending with -*ite*



Practice: Name the Acid

1. H_2S

2. $HClO_3$

3. $HC_2H_3O_2$

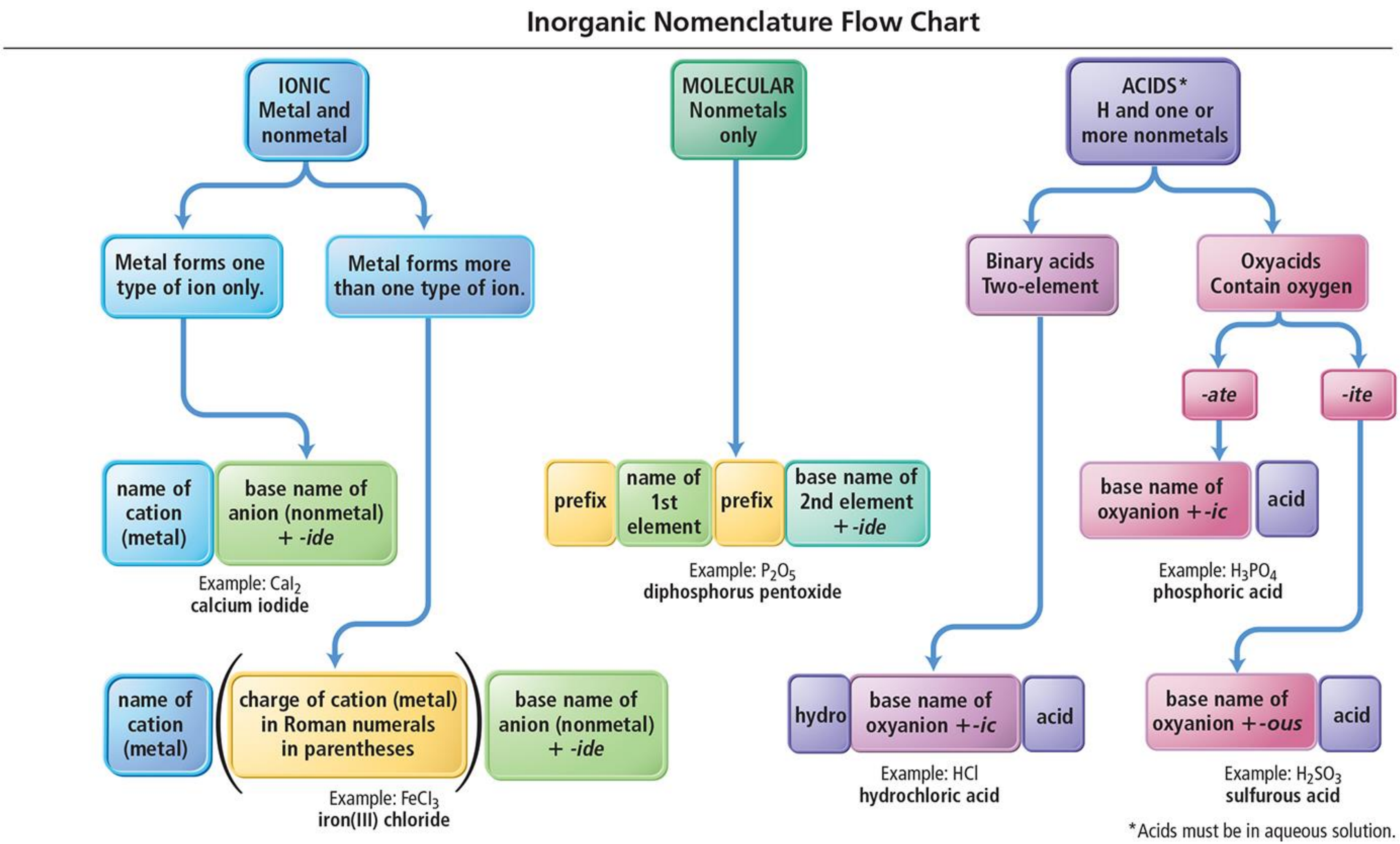
Writing Formulas for Acids

- When the name ends in **acid**, the formula starts with **H** followed by an anion.
- Write the formula as if it is ionic, even though it is molecular.
- *Hydro-* prefix means it is binary acid; no prefix means it is an oxyacid.
- For an oxyacid,
  - if the ending is ***–ic***, the polyatomic ion ends in *–ate*.
  - if the ending is ***–ous***, the polyatomic ion ends in *–ous*.

Acid Rain

- Certain pollutants, such as NO, NO_2, SO_2, and SO_3, form acids when mixed with water, resulting in acidic rainwater.
- Acid rain can fall or flow into lakes and streams, making these bodies of water more acidic.


Inorganic Nomenclature Flow Chart



Formula Mass

- The mass of an individual molecule or formula unit
  - also known as molecular mass or molecular weight
- Sum of the masses of the atoms in a single molecule or formula unit
  - whole = sum of the parts!

Mass of 1 molecule of H_2O =

(2 atoms H)(1.01 amu/H atom) + (1 atom O)(16.00 amu/atom O)

= 18.02 amu

$$\text{Formula mass} = \left(\boxed{\begin{array}{c}\text{Number of atoms}\\ \text{of 1st element in}\\ \text{chemical formula}\end{array}} \times \boxed{\begin{array}{c}\text{Atomic mass}\\ \text{of}\\ \text{1st element}\end{array}}\right) + \left(\boxed{\begin{array}{c}\text{Number of atoms}\\ \text{of 2nd element in}\\ \text{chemical formula}\end{array}} \times \boxed{\begin{array}{c}\text{Atomic mass}\\ \text{of}\\ \text{2nd element}\end{array}}\right) + \ldots$$

Molar Mass of Compounds

- ***The molar mass of a compound*—the mass, in grams, of 1 mol of its molecules or formula units—is numerically equivalent to its formula mass with units of g/mol.**

Molar Mass of Compounds

- The relative masses of molecules can be calculated from atomic masses:

  formula mass = 1 molecule of H_2O = 2(1.01 amu H) + 16.00 amu O = 18.02 amu

- 1 mole of H_2O contains 2 moles of H and 1 mole of O:

  molar mass = 1 mole H_2O

  = 2mol(1.01 g/1 mol H) + 1mol(16.00 g/1 mol O) = 18.02 g/1 mol H_2O

  So the molar mass of H_2O is 18.02 g/mole.

- Molar mass = formula mass (in g/mole)

Using Molar Mass to Count Molecules by Weighing

- Molar mass in combination with Avogadro's number can be used to determine the number of atoms in a given mass of the element.
  - Use molar mass to convert to the amount in moles. Then use Avogadro's number to convert to number of molecules.

$$\#\ molecule\ A = \frac{gram\ A}{1} * \frac{1\ mole\ A}{molar\ mass\ A} * \frac{6.02\times10^{23}\ molecules\ A}{1\ mole\ A}$$

Composition of Compounds

- **A chemical formula, in combination with the molar masses of its constituent elements, indicates the relative quantities of each element in a compound.**

Composition of Compounds

- Percentage by mass of each element in a compound. Can be determined from
  1. the formula of the compound and
  2. the experimental mass analysis of the compound.
- The percentages may not always total to 100% due to rounding.

$$\text{Mass percent of element X} = \frac{\text{mass of element X in 1 mol of compound}}{\text{mass of 1 mol of compound}} \times 100\%$$

Conversion Factors from Chemical Formulas

- Chemical formulas show the relationship between numbers of atoms and molecules.
  - Or moles of atoms and molecules

$$58.64 \text{ g Cl} : 100 \text{ g } CCl_2F_2$$

$$1 \text{ mol } CCl_2F_2 : 2 \text{ mol Cl}$$

- These relationships can be used to determine the amounts of constituent elements and molecules.
  - Like percent composition

Determining a Chemical Formula from Experimental Data

Empirical Formula

- Simplest, whole-number ratio of the atoms or moles of elements in a compound, *not a ratio of masses*
- Can be determined from elemental analysis
  - Percent composition
  - Masses of elements formed when a compound is decomposed, or that react together to form a compound

Finding an Empirical Formula

1. Convert the percentages to grams.
   (a) Assume you start with 100 g of the compound.
   (b) Skip if it is already in grams.
2. Convert grams to moles.
   (a) Use the molar mass of each element.
3. Write a pseudoformula using moles as subscripts.

Finding an Empirical Formula (continued)

4. Divide all by the smallest number of moles.
   (a) If the result is within 0.1 of a whole number, round to the whole number.
5. Multiply all mole ratios by a number to make all whole numbers.
   (a) If ratio is .5, multiply all by 2.
   (b) If ratio is .33 or .67, multiply all by 3.
   (c) If ratio is 0.25 or 0.75, multiply all by 4, etc.
   (d) Skip if ratios are already whole numbers.

Molecular Formulas for Compounds

- The molecular formula is a whole-number multiple of the empirical formula.
- To determine the molecular formula, you need to know the empirical formula and the molar mass of the compound.

Molecular formula = (empirical formula)***n***,

*where **n** is a positive integer.*

$$n = \frac{\text{molar mass}}{\text{empirical formula molar mass}}$$

Chemical Equations

- Shorthand way of describing a reaction
- Provide information about the reaction
  - Formulas of reactants and products
  - States of reactants and products
  - Relative numbers of reactant and product molecules that are required can be used to determine weights of reactants used and products that can be made

TABLE 3.5 States of Reactants and Products in Chemical Equations

Abbreviation	State
(*g*)	Gas
(*l*)	Liquid
(*s*)	Solid
(*aq*)	Aqueous (water solution)

Combustion of Methane

- Methane gas burns to produce carbon dioxide gas and gaseous water.
  - Whenever something burns it combines with $O_2(g)$.

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

- If you look closely, you should immediately spot a problem.

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

2 O atoms

2 O atoms + 1 O atom =
3 O atoms

Combustion of Methane

- Notice also that the left side has four hydrogen atoms while the right side has only two.

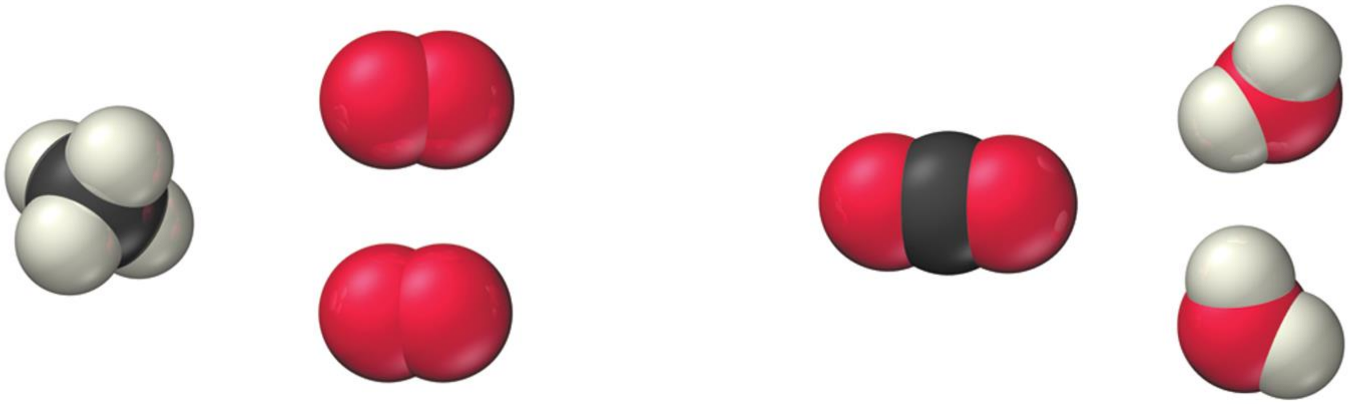
$$\underset{\text{4 H atoms}}{CH_4(g)} + O_2(g) \longrightarrow CO_2(g) + \underset{\text{2 H atoms}}{H_2O(g)}$$

- To correct these problems, we must **balance** the equation by changing the coefficients, not the subscripts.

Combustion of Methane: Balanced

- To show that the reaction obeys the Law of Conservation of Mass, the equation must be **balanced**.
  - We adjust the numbers of molecules so there are equal numbers of atoms of each element on both sides of the arrow.

$$CH_4(g) \quad + \quad 2\,O_2(g) \longrightarrow CO_2(g) \quad + \quad 2\,H_2O(g)$$



Quantities in Chemical Reactions

- The amount of every substance used and made in a chemical reaction is related to the amounts of all the other substances in the reaction.
  - Law of conservation of mass
  - Balancing equations by balancing atoms
- The study of the numerical relationship between chemical quantities in a chemical reaction is called **stoichiometry**.

Reaction Stoichiometry

- **The coefficients in a chemical reaction specify the relative amounts in moles of each of the substances involved in the reaction.**

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

  - 2 molecules of C_8H_{18} react with 25 molecules of O_2 to form 16 molecules of CO_2 and 18 molecules of H_2O.
  - 2 moles of C_8H_{18} react with 25 moles of O_2 to form 16 moles of CO_2 and 18 moles of H_2O.

2 mol C_8H_{18} : 25 mol O_2 : 16 mol CO_2 : 18 mol H_2O

Making Pizza

- The number of pizzas you can make depends on the amount of the ingredients you use.

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

This relationship can be expressed mathematically.

1 crust : 5 oz sauce : 2 cups cheese : 1 pizza

- We can compare the amount of pizza that can be made from 10 cups of cheese:

2 cups cheese : 1 pizza, then,

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

Making Molecules: Mole-to-Mole Conversions

- We use the ratio from the balanced chemical equation in the same way that we used the ratio from the pizza recipe.

The ratio of the coefficients acts as a conversion factor between the amount in moles of the reactants and products.

$$\mathbf{2\ C_8H_{18}(\mathit{l}) + 25\ O_2(\mathit{g}) \rightarrow 16\ CO_2(\mathit{g}) + 18\ H_2O(\mathit{g})}$$

stoichiometric ratio: 2 moles C_8H_{18} : 16 moles CO_2

The ratio acts as a conversion factor between the amount in moles of the reactant, C_8H_{18}, and the amount in moles of the product, CO_2.

Mole-to-Mole Conversions

- Suppose we burn 22.0 moles of C_8H_{18}; how many moles of CO_2 form?

$$\mathbf{2\ C_8H_{18}(\mathit{l}) + 25\ O_2(\mathit{g}) \rightarrow 16\ CO_2(\mathit{g}) + 18\ H_2O(\mathit{g})}$$

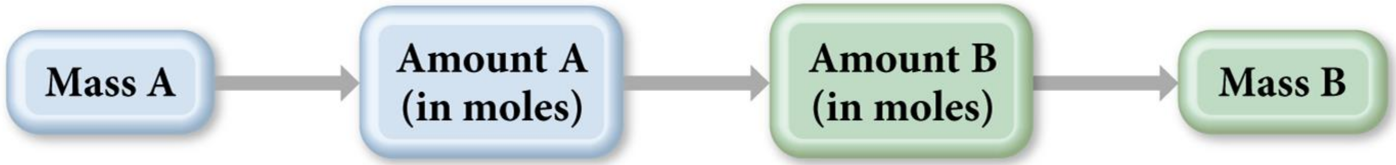
stoichiometric ratio: 2 moles C_8H_{18} : 16 moles CO_2

$$22.0\ \cancel{\text{mol}\ C_8H_{18}} \times \frac{16\ \text{mol}\ CO_2}{2\ \cancel{\text{mol}\ C_8H_{18}}} = 176\ \text{mol}\ CO_2$$

- The combustion of 22.0 moles of C_8H_{18} adds 176 moles of CO_2 to the atmosphere.

Making Molecules: Mass-to-Mass Conversions

- The world burned the equivalent of 3.7×10^{15} g of gasoline (octane) in 2013. We can estimate the mass of CO_2 produced based on the flow chart below.



- We use molar mass as a conversion factor between the mass given and amount in moles.
- We use coefficients as the conversion factor between the reactant, C_8H_{18}, and the amount in moles of the product, CO_2, and then molar mass as the conversion factor to get the mass of CO_2 produced.

Mass-to-Mass Conversions

- If we burn 3.7×10^{15} g C_8H_{18}, how many grams of CO_2 form?

$2\ C_8H_{18}(\mathit{l}) + 25\ O_2(\mathit{g}) \rightarrow 16\ CO_2(\mathit{g}) + 18\ H_2O(\mathit{g})$

molar mass: C_8H_{18} 114.22 g/mol, CO_2 44.01 g/mol
stoichiometric ratio: 2 moles C_8H_{18} : 16 moles CO_2

$$22.0\ \cancel{\text{mol } C_8H_{18}} \times \frac{16\ \text{mol } CO_2}{2\ \cancel{\text{mol } C_8H_{18}}} = 176\ \text{mol } CO_2$$

- The combustion 3.7×10^{15} g C_8H_{18} adds 1.1×10^{16} g CO_2 to the atmosphere.

Limiting Reactant, Theoretical Yield, Percent Yield

- Recall our 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 tomato sauce, how many pizzas can we make?

We have enough crusts to make $4\ \cancel{\text{crusts}} \times \dfrac{1\ \text{pizza}}{1\ \cancel{\text{crust}}} = 4\ \text{pizzas}$

We have enough cheese to make $10\ \cancel{\text{cups cheese}} \times \dfrac{1\ \text{pizza}}{2\ \cancel{\text{cups cheese}}} = 5\ \text{pizzas}$

We have enough tomato sauce to make

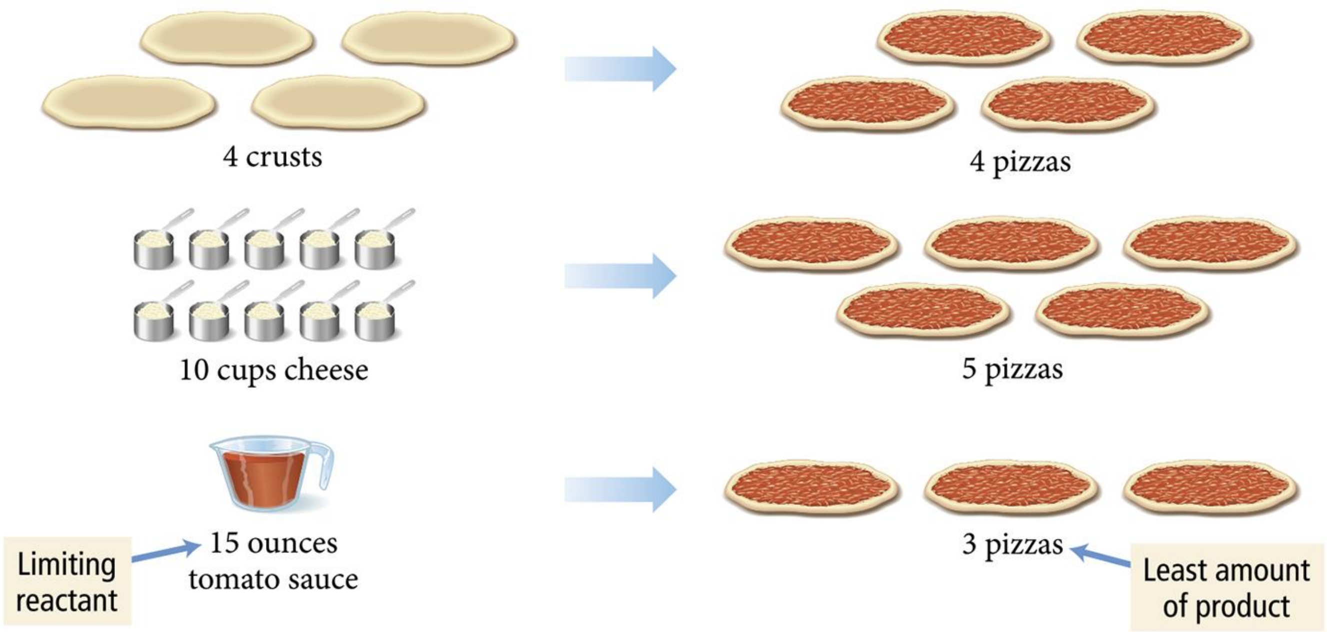
$$15\ \cancel{\text{ounces tomato sauce}} \times \frac{1\ \text{pizza}}{5\ \cancel{\text{ounces tomato sauce}}} = 3\ \text{pizzas}$$

Limiting reactant

Smallest number of pizzas

Limiting Reactant

- We have enough crusts for four pizzas, enough cheese for five pizzas, but enough tomato sauce for only three pizzas.
  - We can make only three pizzas. The tomato sauce *limits* how many pizzas we can make.



Theoretical Yield

- Tomato sauce is the **limiting reactant**, the reactant that makes *the least amount of product.*
  - The limiting reactant is also known as the *limiting reagent.*
- The maximum number of pizzas we can make depends on this ingredient. In chemical reactions, we call this the **theoretical yield**.
  - This is the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.
  - The ingredient that makes the least amount of pizza determines how many pizzas you can make (**theoretical yield**).

Percent Yield

Assume that while making pizzas, we burn a pizza, drop one on the floor, or other uncontrollable events happen so that we make only two pizzas. The actual amount of product made in a chemical reaction is called the **actual yield.**

We can determine the efficiency of making pizzas by calculating the percentage of the maximum number of pizzas we actually make. In chemical reactions, we call this the **percent yield**.

$$\%\ \text{yield} = \frac{\overbrace{2\ \cancel{\text{pizzas}}}^{\text{Actual yield}}}{\underbrace{3\ \cancel{\text{pizzas}}}_{\text{Theoretical yield}}} \times 100\% = 67\%$$

In a Chemical Reaction

- For reactions with multiple reactants, it is likely that one of the reactants will be completely used before the others.
- When this reactant is used up, the reaction stops and no more product is made.
- The reactant that limits the amount of product is called the **limiting reactant**.
  - It is sometimes called the limiting reagent.
  - The limiting reactant gets completely consumed.
- Reactants not completely consumed are called **excess reactants**.
- The amount of product that can be made from the limiting reactant is called the **theoretical yield**.

Summarizing Limiting Reactant and Yield

- The **limiting reactant** (or **limiting reagent**) is the reactant that is completely consumed in a chemical reaction and limits the amount of product.

- The **reactant in excess** is any reactant that occurs in a quantity greater than is required to completely react with the limiting reactant.

Summarizing Limiting Reactant and Yield

- The **theoretical yield** is the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.
- The **actual yield** is the amount of product actually produced by a chemical reaction.
- The **percent yield** is calculated as follows:

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

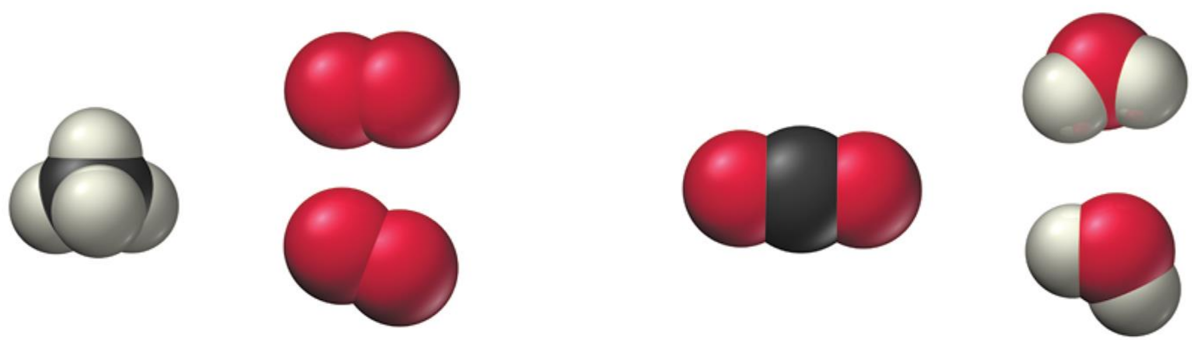
Calculating Limiting Reactant, Theoretical Yield, and Percent Yield

- Recall our balanced equation for the combustion of methane:

$$CH_4(g) + \mathbf{2}\ O_2(g) \rightarrow CO_2(g) + \mathbf{2}\ H_2O(g)$$

  - Our balanced equation for the combustion of methane implies that every one molecule of CH_4 reacts with two molecules of O_2.

$$CH_4(g) + 2\ O_2(g) \longrightarrow CO_2(g) + 2\ H_2O(l)$$



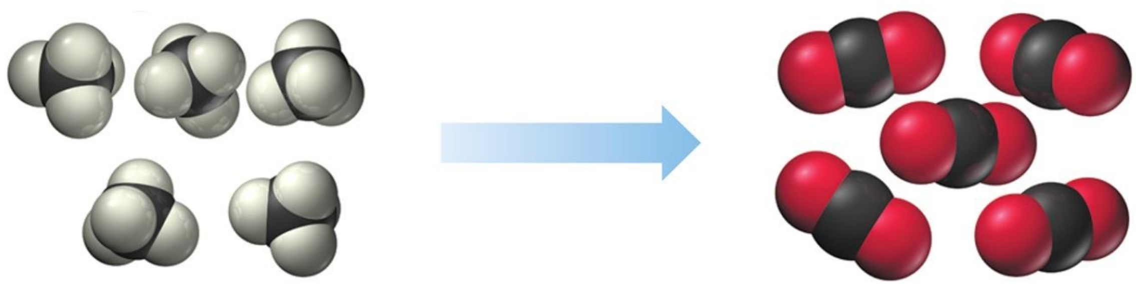
Combustion of Methane

- If we have five molecules of CH_4 and eight molecules of O_2, which is the limiting reactant?

$CH_4(g)$ + **2** $O_2(g) \rightarrow CO_2(g)$ + **2** $H_2O(g)$

– First we calculate the number of CO_2 molecules that can be made from five CH_4 molecules.

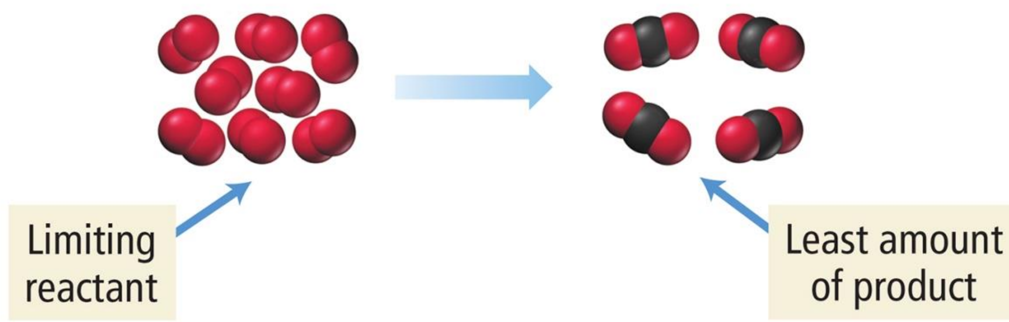
$$5\ \cancel{CH_4} \times \frac{1\ CO_2}{1\ \cancel{CH_4}} = 5\ CO_2$$



Combustion of Methane

- Then we calculate the number of CO_2 molecules that can be made from eight O_2 molecules.

$$8\,\cancel{O_2} \times \frac{1\,CO_2}{2\,\cancel{O_2}} = 4\,CO_2$$



  - We have enough CH_4 to make five CO_2 molecules and four CO_2 molecules.
  - Therefore, O_2 is the limiting reactant, and four CO_2 molecules is the theoretical yield.
  - CH_4 is in excess.

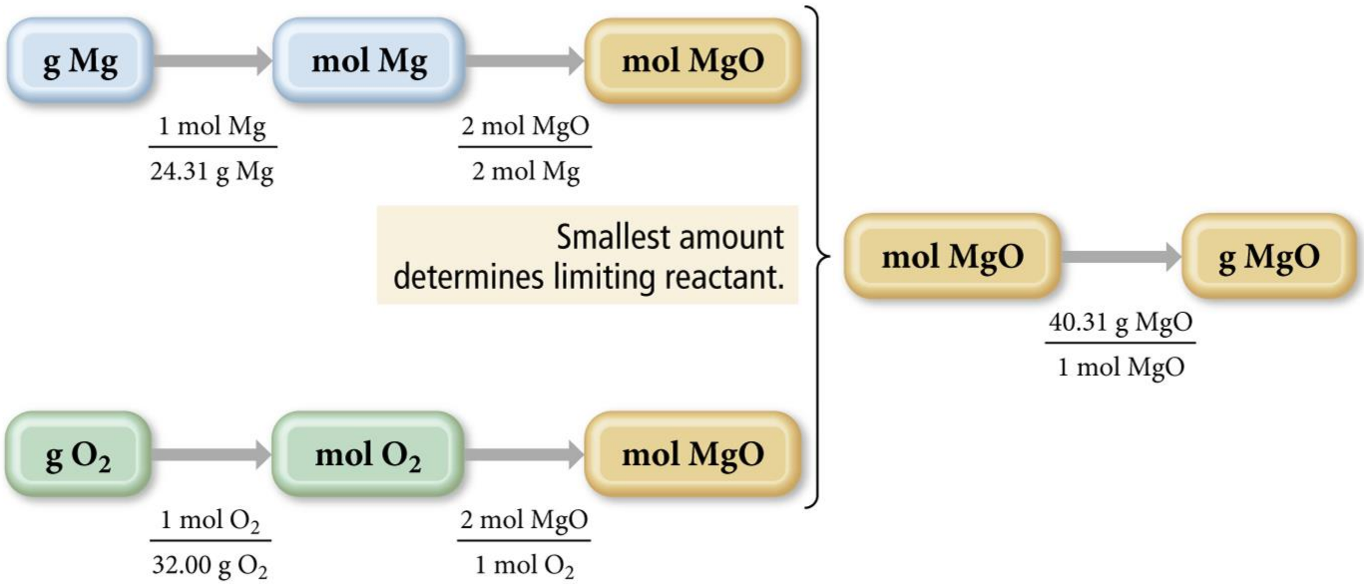
Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses

- When working in the lab, we normally measure reactant quantities in grams.
- To find the limiting reactant and theoretical yield, we must first convert grams to moles.

Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses

- A reactant mixture contains 42.5 g Mg and 33.8 g O_2. What is the limiting reactant and theoretical yield?

$$2\ \mathrm{Mg}(s) + \mathrm{O_2}(g) \rightarrow \mathbf{2}\ \mathrm{MgO}(s)$$



Calculating Limiting Reactant, Theoretical Yield, and Percent Yield from Reactant Masses

$$\mathbf{42.5\ \cancel{g\ Mg}} \times \frac{1\ \cancel{\text{mol Mg}}}{24.31\ \cancel{\text{g Mg}}} \times \frac{2\ \text{mol MgO}}{2\ \cancel{\text{mol Mg}}} = \mathbf{1.7\underline{4}83\ mol\ MgO}$$

Limiting reactant

Least amount of product

$$\mathbf{33.8\ \cancel{g\ O_2}} \times \frac{1\ \cancel{\text{mol } O_2}}{32.00\ \cancel{\text{g } O_2}} \times \frac{2\ \text{mol MgO}}{1\ \cancel{\text{mol } O_2}} = \mathbf{2.1\underline{1}25\ mol\ MgO}$$

$$\mathbf{1.7\underline{4}83\ \cancel{mol\ MgO}} \times \frac{40.31\ \text{g MgO}}{1\ \cancel{\text{mol MgO}}} = \mathbf{70.5\ g\ MgO}$$

Solution Concentration and Solution Stoichiometry

- When table salt is mixed with water, it seems to disappear or become a liquid; the mixture is homogeneous.
  - The salt is still there, as you can tell from the taste or simply boiling away the water.
- Homogeneous mixtures are called **solutions**.
- The majority component is the **solvent**.
- The minority component is the **solute**.
- A solution in which water is the solvent is an **aqueous solution**.

Solution Concentration

- Because solutions are mixtures, the composition can vary from one sample to another.
  - Pure substances have constant composition.
  - Saltwater samples from different seas or lakes have different amounts of salt.

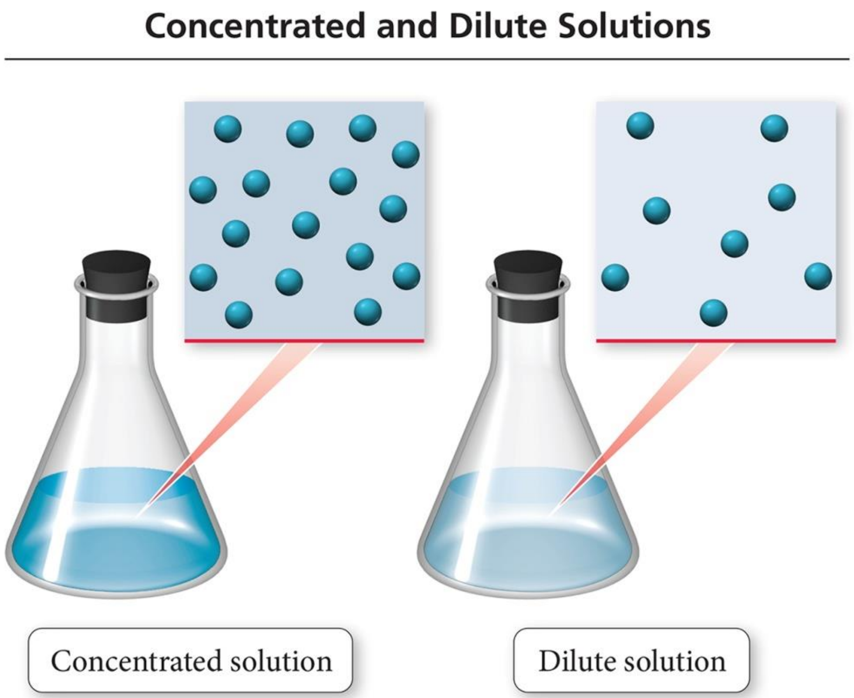
- So, to describe solutions accurately, we quantify the amount of solute relative to solvent, or **concentration of solution**.

Solution Concentration

- Solutions are often described quantitatively, as dilute or concentrated.

- **Dilute solutions** have a small amount of solute compared to solvents.

- **Concentrated solutions** have a large amount of solute compared to solvents.

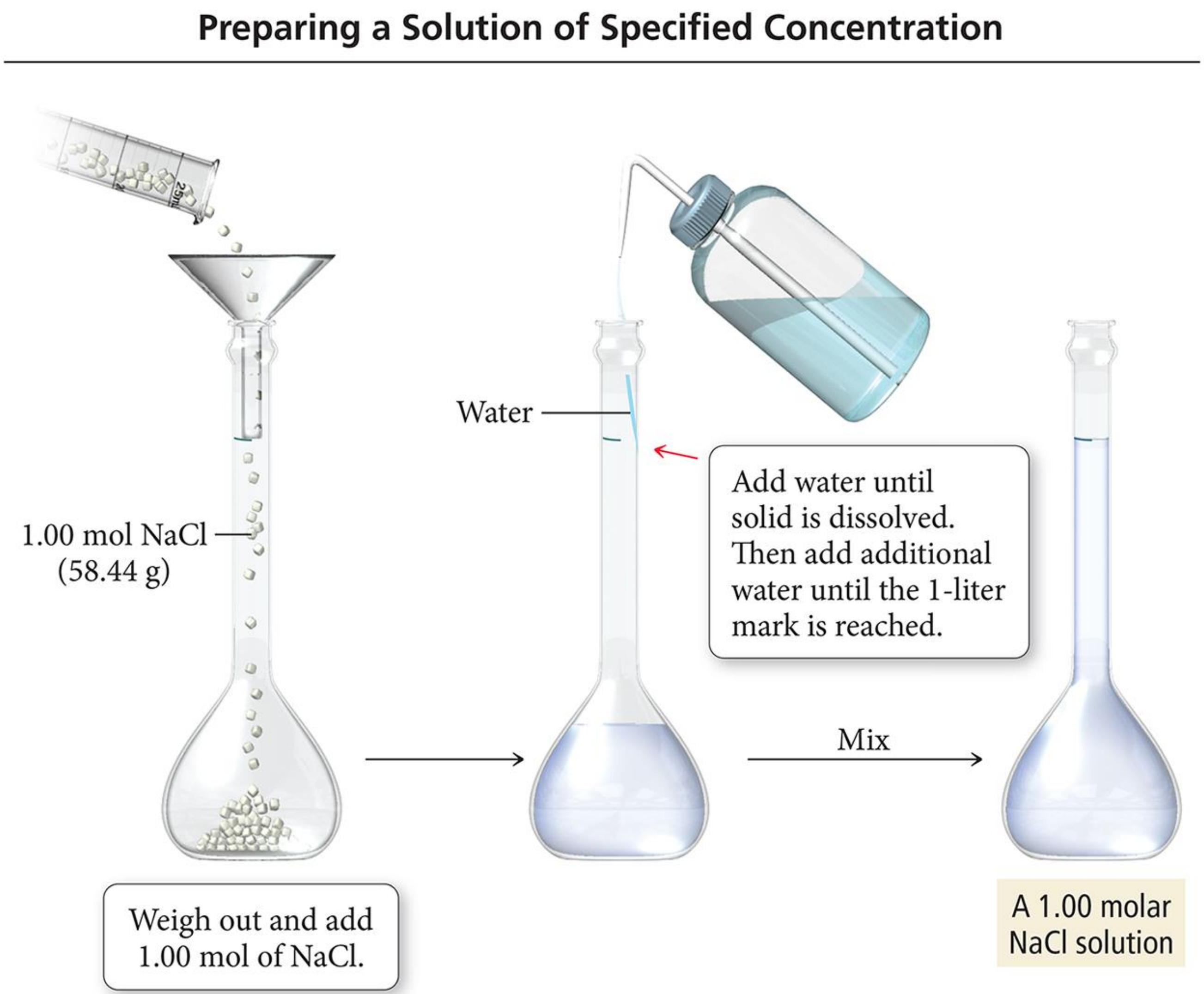


Solution Concentration: Molarity

- A common way to express solution concentration is **molarity (*M*)**.
  - Molarity is the amount of solute (in moles) divided by the volume of solution (in liters).

$$\text{Molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$

Preparing 1 L of a 1.00 M NaCl Solution



Using Molarity in Calculations

- We can use the molarity of a solution as a conversion factor between moles of the solute and liters of the solution.
  - For example, a 0.500 M NaCl solution contains 0.500 mol NaCl for every liter of solution.

$$\frac{0.500 \text{ mol NaCl}}{\text{L solution}} \quad \textit{converts} \quad \textbf{L solution} \rightarrow \textbf{mol NaCl}$$

$$\frac{\text{L solution}}{0.500 \text{ mol NaCl}} \quad \textit{converts} \quad \textbf{mol NaCl} \rightarrow \textbf{L solution}$$

Solution Dilution

- Often, solutions are stored as concentrated **stock solutions**.
- To make solutions of lower concentrations from these stock solutions, more solvent is added.
  - **The amount of solute doesn't change, just the volume of solution:**

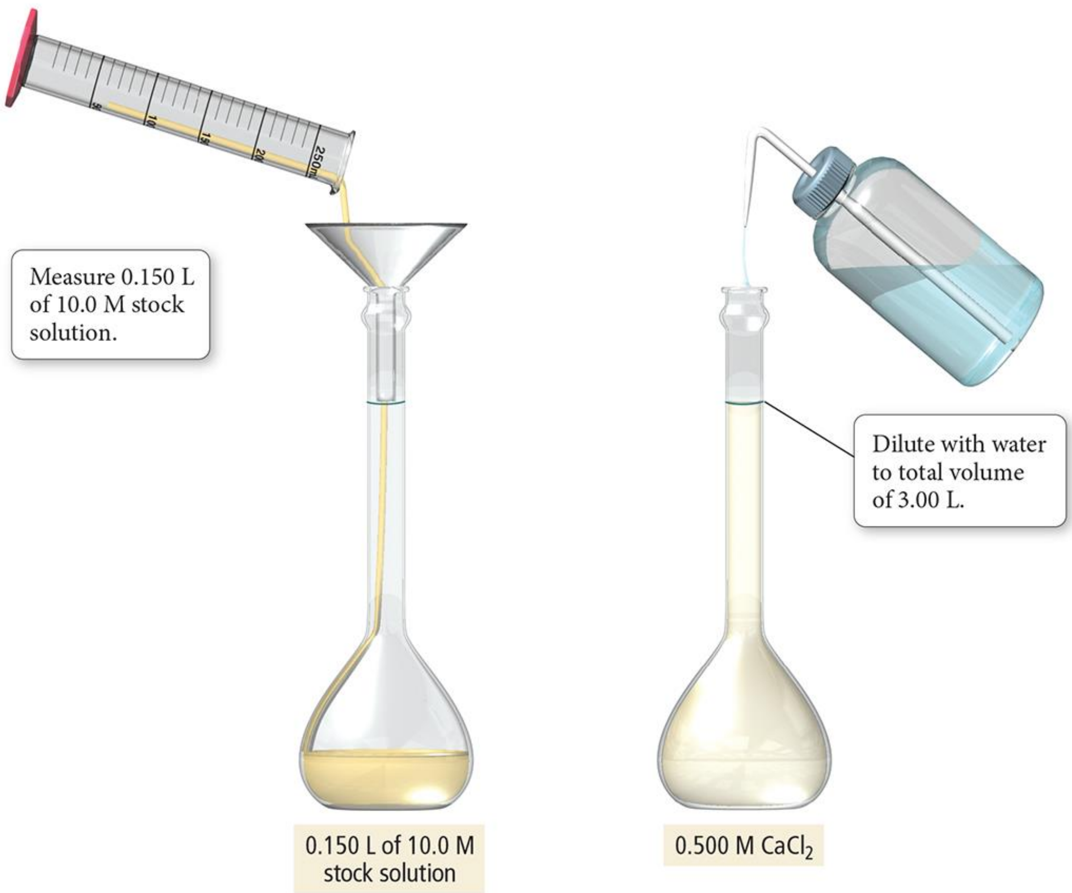
moles solute in solution 1 = moles solute in solution 2

- The concentrations and volumes of the stock and new solutions are inversely proportional:

$$\mathbf{M_1 \cdot V_1 = M_2 \cdot V_2}$$

Preparing 3.00 L of 0.500 M $CaCl_2$ from a 10.0 M Stock Solution

Diluting a Solution



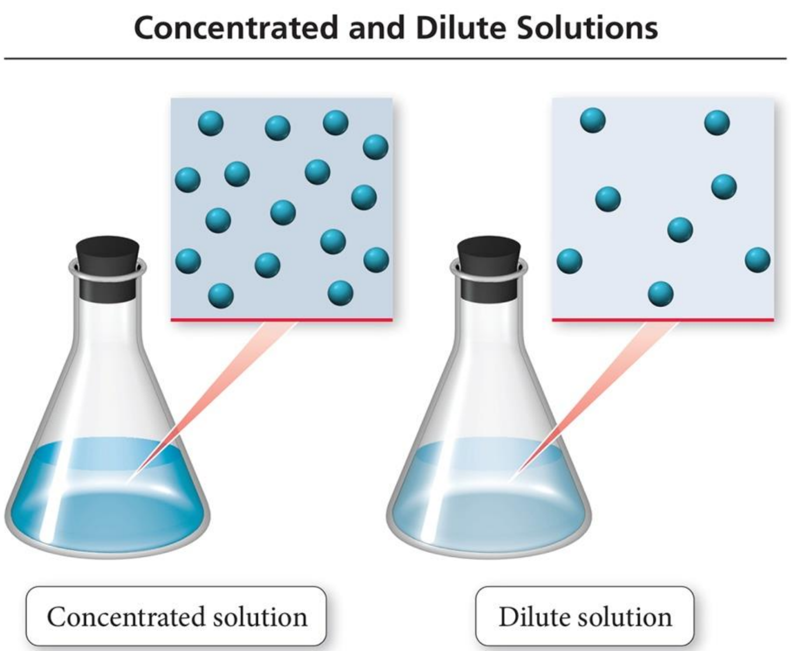
$$M_1V_1 = M_2V_2$$

$$\frac{10.0\ \text{mol}}{\cancel{\text{L}}} \times 0.150\ \cancel{\text{L}} = \frac{0.500\ \text{mol}}{\cancel{\text{L}}} \times 3.00\ \cancel{\text{L}}$$

$$1.50\ \text{mol} = 1.50\ \text{mol}$$

Solution Stoichiometry

- Because molarity relates the moles of solute to the liters of solution, it can be used to convert between amount of reactants and/or products in a chemical reaction.
  - The general conceptual plan for these kinds of calculations begins with the volume of a reactant or product.



Types of Aqueous Solutions and Solubility

- Consider two familiar aqueous solutions: saltwater and sugar water.
  - Saltwater is a homogeneous mixture of NaCl and H_2O.
  - Sugar water is a homogeneous mixture of $C_{12}H_{22}O_{11}$ and H_2O.
- As you stir either of these two substances into the water, it seems to disappear.
  - How do solids such as salt and sugar dissolve in water?