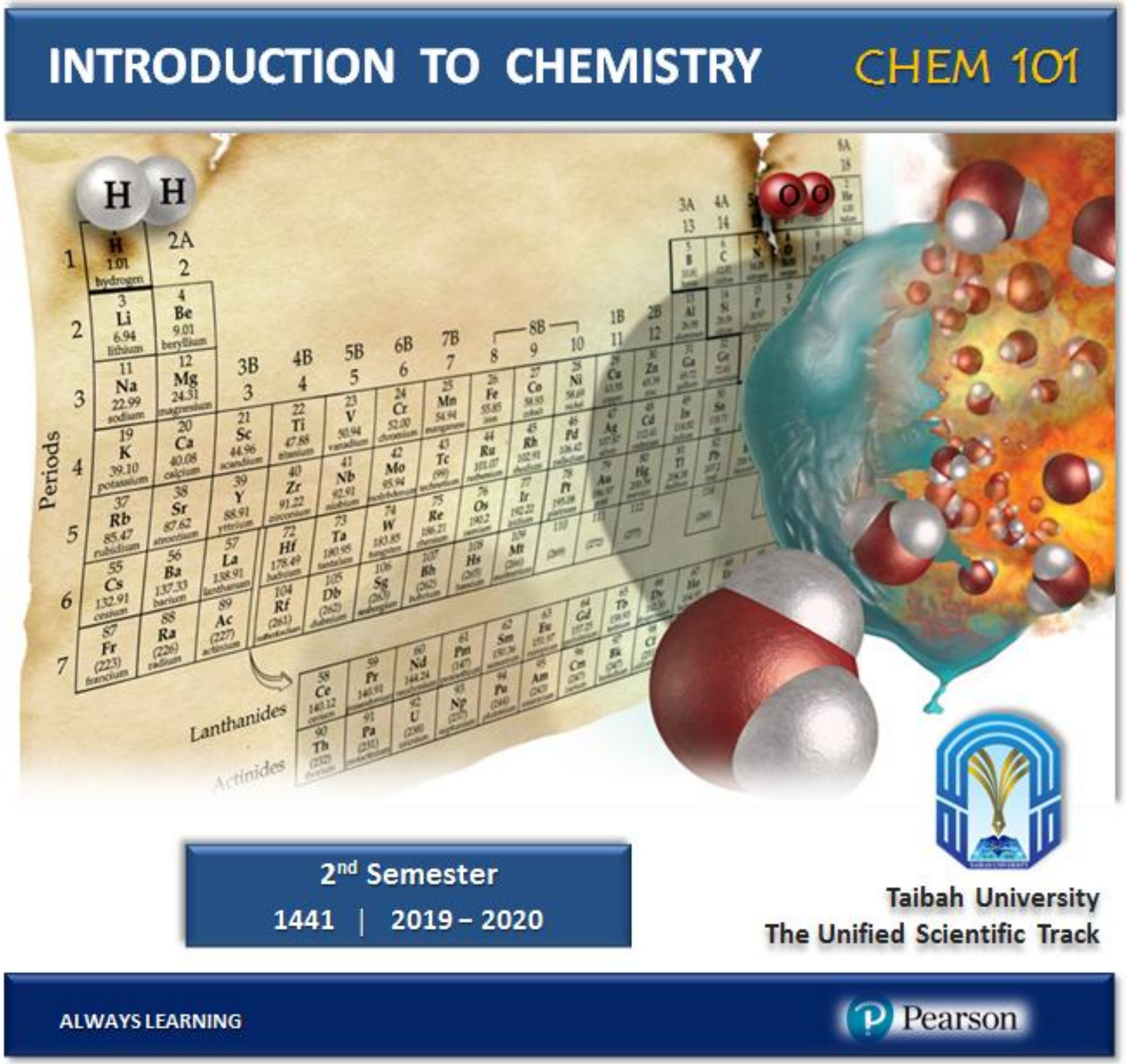


Lecture Presentation

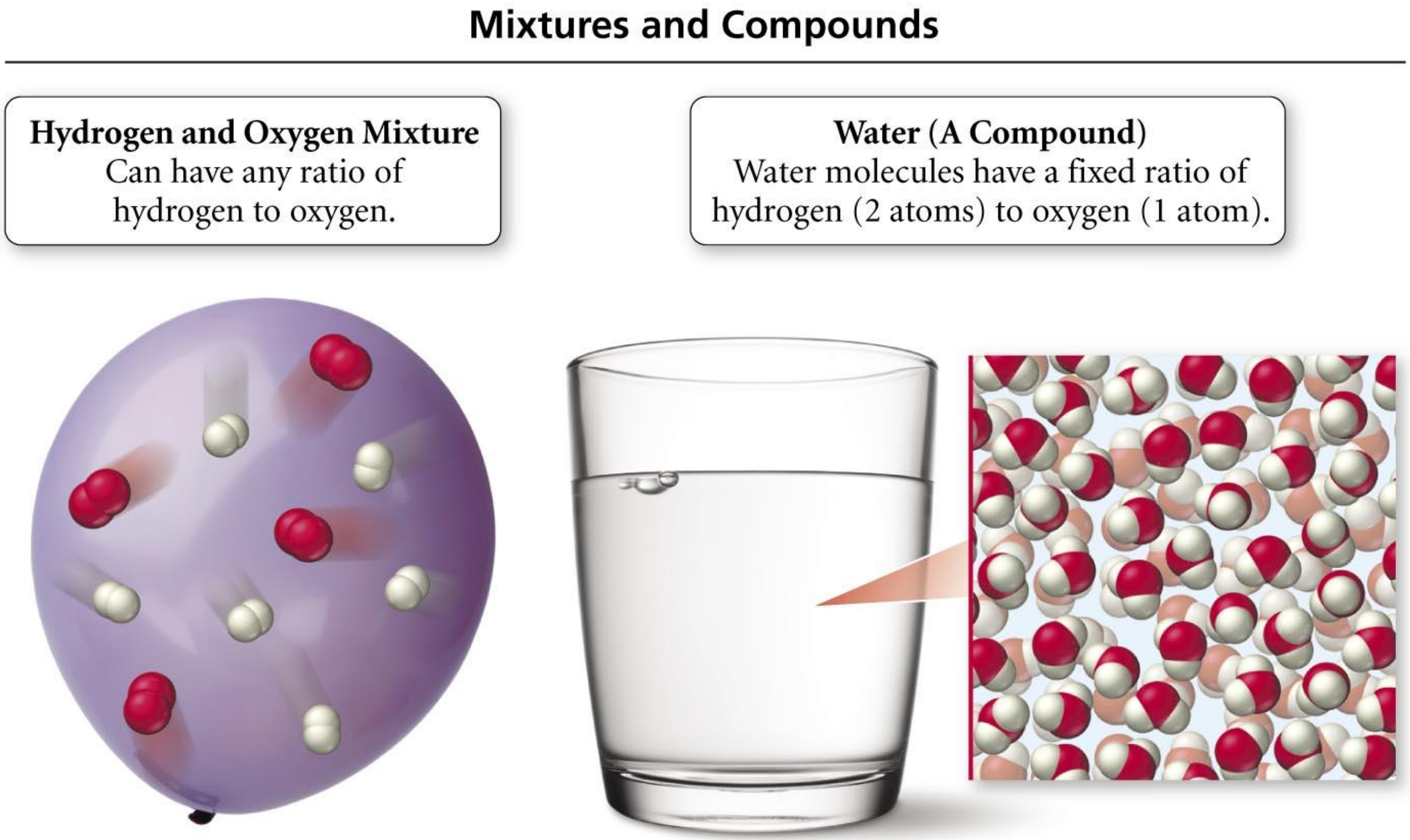
Chapter 3

Molecules, Compounds, and Chemical Equations

Topic 08

Chemical Formulas and Molecular Models

3.1 Elements, Compounds & Mixtures



Elements & Compounds

- **Elements** are the simplest form of matter, they can combine together to make a limitless number of **compounds**.
- The properties of the compounds are totally different from their constituent elements.

✓ **Example**: the properties of water are different from the properties of both H_2 and O_2:

Selected Properties	Hydrogen 	Oxygen 	Water 
Boiling Point	−253 °C	−183 °C	100 °C
State at Room Temperature	Gas	Gas	Liquid
Flammability	Explosive	Necessary for combustion	Used to extinguish flame

- A **Compound** is a distinct substance that is composed of bonded atoms of two or more elements.
- Describe the compound by describing the number and type of each atom in the simplest unit of the compound: molecules
- Each **element** is represented by its letter symbol (from the periodic table)
- The number of atoms of each element is written to the right of the element as a subscript (if there is only one atom of an element, the "1" subscript is not written), e.g. $C_6H_{12}O_6$, CH_3Br
- Polyatomic ions are placed in parentheses (if more than one), e.g. $Mg(NO_3)_2$, $Al(OH)_3$, $NaOH$

3.2 Representing Compounds: Chemical Formulas & Molecular Models

- Compounds are generally represented with their Chemical Formulas or Molecular Models.
- Chemical formula indicates the type and number of each element present in the compound (using the letter symbols of the elements from the periodic table):
  - Water is represented as H_2O
  - Carbon dioxide is represented as CO_2
  - Sodium chloride is represented as NaCl
  - Carbon tetrachloride is represented as CCl_4

Types of Chemical Formulas

- Chemical formulas can generally be categorized into **three different types**:

  1. **Empirical Formula**
  2. **Molecular Formula**
  3. **Structural Formula**

- The amount of information about the structure of the compound varies with the type of formula.
- All formulas and models convey a limited amount of information – none is a perfect representation!

Types of Chemical Formulas: The Empirical Formula

- **Empirical Formula:** gives the **<u>relative number</u>** of atoms of each element in a compound.
- It does not describe the actual number of atoms, the order of attachment, or the shape of molecules.
- It is the <u>simplest whole-number</u> (ratio) representation of the type and number of elements present in a molecule.
- The <u>Ionic compounds</u> (metal + nonmetal) are usually represented using their <u>Empirical Formulas</u> (formula unites): (e.g. MgO **<u>not</u>** Mg_2O_2 & CaS **<u>not</u>** Ca_2S_2 , ….)

Types of Chemical Formulas: The Molecular Formula

- **Molecular Formula:** gives the **actual number** of atoms of each element in a molecule of a compound.
- It does not describe the order of attachment, or the shape of molecules.

✓ **Examples**:

a) H_2O is the molecular formula of water, which means that the water molecule is actually composed of 2 hydrogen atoms + 1 oxygen atom.

b) C_4H_8 is a molecular formula, which means that the molecule is actually composed of 4 carbon atoms + 8 hydrogen atoms.

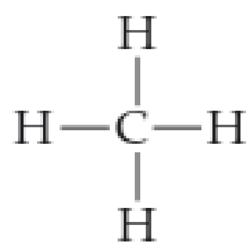
c) B_2H_6 is a molecular formula, which means that the molecule is actually composed of 2 boron atoms + 6 hydrogen atoms.

Types of Chemical Formulas: The Structural Formula

- **Structural Formula:** is a **sketch or diagram** of how the atoms in the molecule are bonded to each other.
- It uses lines to represent covalent bonds and shows how atoms in a molecule are connected or bonded to each other.
- lines describe the number of electrons shared by the bonded atoms:
  - Single line = two shared electrons, a **single** covalent bond
  - Double line = four shared electrons, a **double** covalent bond
  - Triple line = six shared electrons, a **triple** covalent bond
- It's used only with "Molecular Compounds" (Nonmetal + Nonmetal), but NOT with "Ionic Compounds" (Metal + Nonmetal)

✓ Example: The structural formula for **CO_2** is: O=C=O

✓ Example: The structural formula for methane **CH_4** is:



Molecular Models

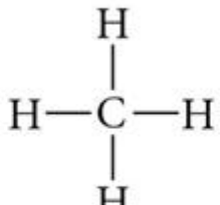
There are two common <u>molecular models</u> used to represent the molecules of compounds:

- Ball-and-Stick Model
- Space-Filling Model

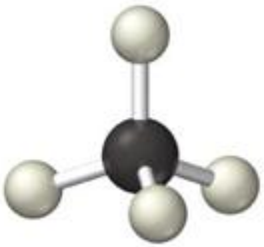
<u>Example</u>: the different ways to represent the methane molecule (CH_4):

CH_4

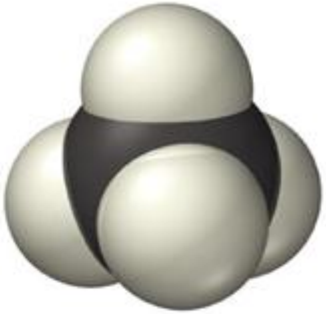
Molecular formula
(also is empirical formula)



Structural formula



Ball-and-stick model



Space-filling model

Exercises: Molecular and Empirical Formulas

Write empirical formulas for the compounds represented by the molecular formulas.

(a) C_4H_8
(b) B_2H_6
(c) CCl_4

SOLUTION

To determine the empirical formula from a molecular formula, divide the subscripts by the greatest common factor (the largest number that divides exactly into all of the subscripts).

(a) For C_4H_8, the greatest common factor is 4. The empirical formula is therefore CH_2.
(b) For B_2H_6, the greatest common factor is 2. The empirical formula is therefore BH_3.
(c) For CCl_4, the only common factor is 1, so the empirical formula and the molecular formula are the same.

FOR PRACTICE 5.1

Write the empirical formula for the compounds represented by the molecular formulas.

(a) C_5H_{12}
(b) Hg_2Cl_2
(c) $C_2H_4O_2$

Exercise: Write the Empirical Formulas for the following compounds:

a. CH_4 **b.** C_2H_6 **c.** C_2H_2 **d.** C_2H_5Cl **e.** $C_6H_6Br_6$

3.3 An Atomic Level View of Elements and Compounds

1. **Atomic Elements**: elements whose particles are single atoms, most of the elements in the periodic table are "atomic elements":

   e.g. Fe , Na , Al , Ne , Hg, …..

2. **Molecular Elements**: elements whose particles are multi-atom molecules, having the same type of atoms (i.e. atoms of the same element), (atoms are bonded by covalent bond):

   e.g. H_2 , O_2 , N_2 , Cl_2 …. (see the next slide)

3. **Molecular Compounds**: compounds whose particles are molecules made of only nonmetals, (bonded by covalent bond):

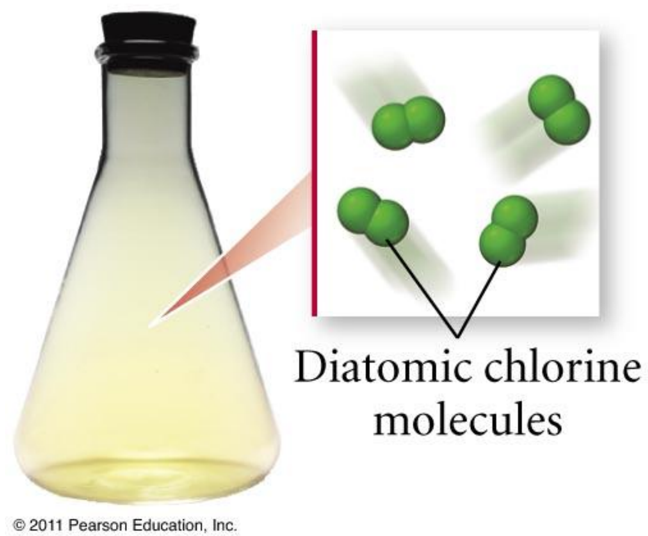
   e.g. H_2O , NH_3 , HCl , CH_4

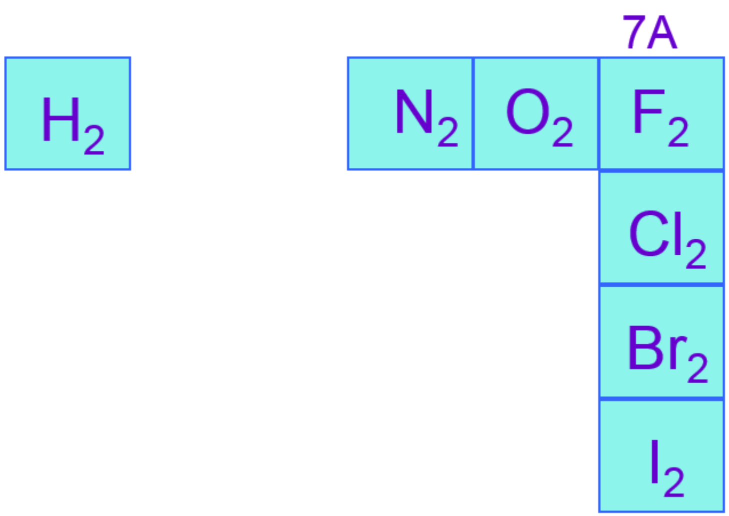
4. **Ionic Compounds**: compounds whose particles are composed of cations (of metals) and anions (of nonmetals), (bonded by ionic bond):

   e.g. $NaCl$, AlF_3 , Fe_2O_3 , Mg_2S

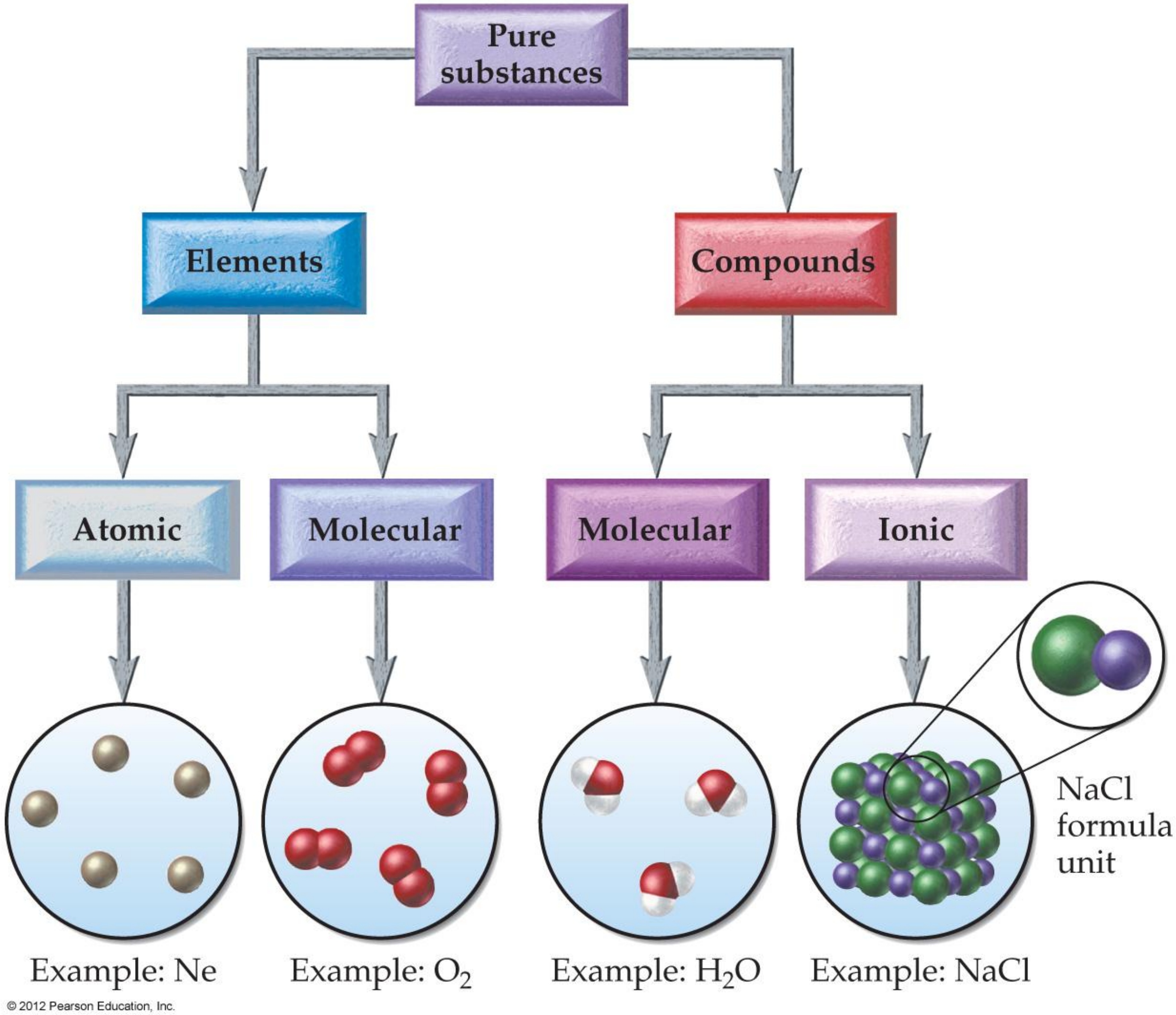
Molecular Elements

- Certain elements occur as diatomic molecules (only 7 out of the 118 elements of the periodic table):
  - ✓ H_2 , N_2 , O_2 , F_2 , Cl_2 , Br_2 and I_2
- Some other elements occur as polyatomic molecules:
  - ✓ P_4 , S_8 , Se_8 and O_3 (Ozone Gas)





Classification of Elements and Compounds: A Summary



Assessment

Classify the following substances as: Atomic Elements, Molecular Elements, Molecular compounds, or Ionic Compounds:

a. Barium, Ba ....

b. Iron (III) chloride, $FeCl_3$ ....

c. Bromine, Br_2 ....

d. Ethanol, C_2H_6O ....

e. Nitrogen monoxide, NO ....

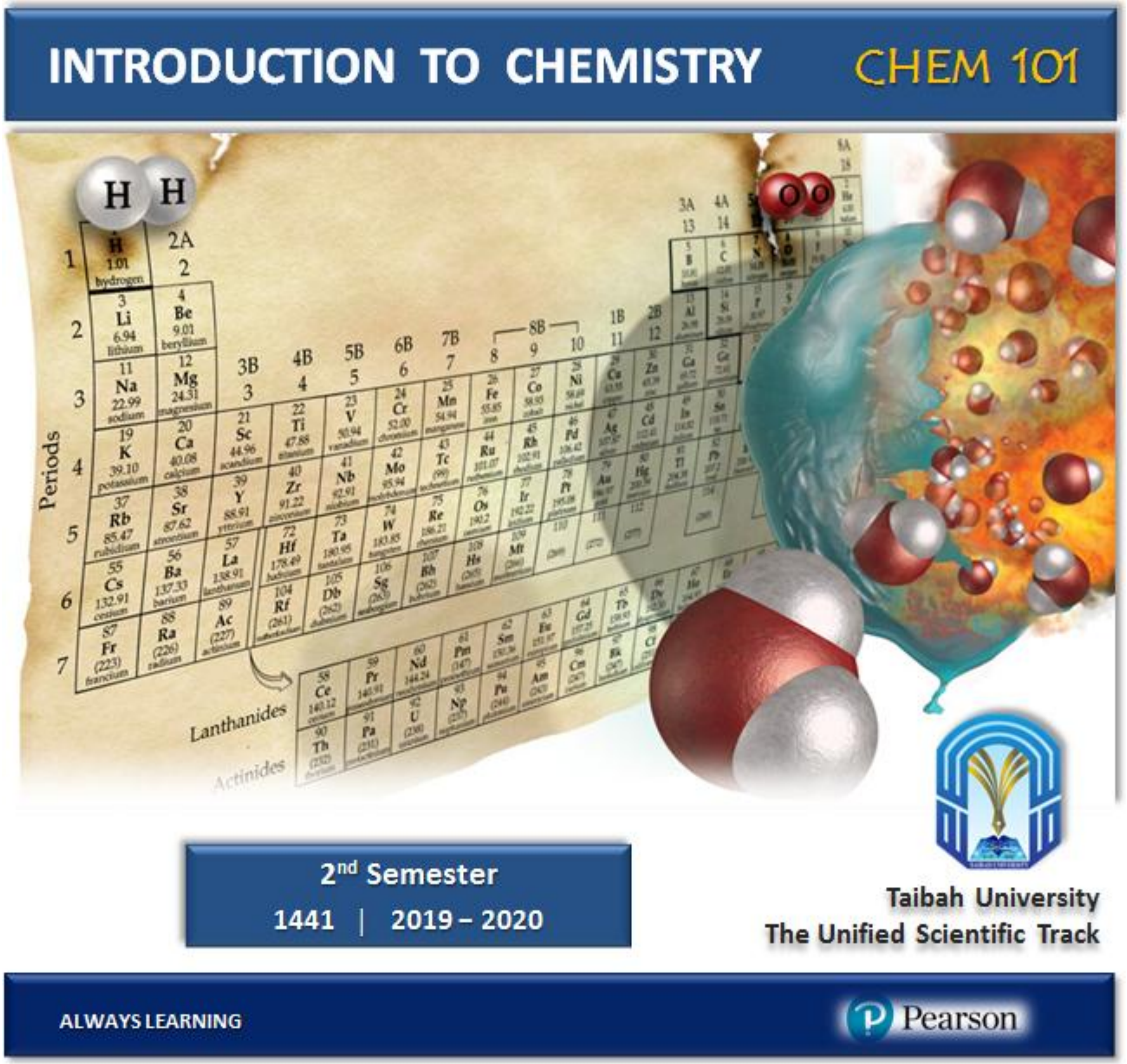
f. Cobalt, Co ....

g. Carbon monoxide, CO ....

h. Nickel(II) chloride, $NiCl_2$, ...

i. Sodium iodide, NaI ....

j. Phosphorus chloride, PCl_3 ....



Lecture Presentation

Chapter 3

Molecules, Compounds, and Chemical Equations

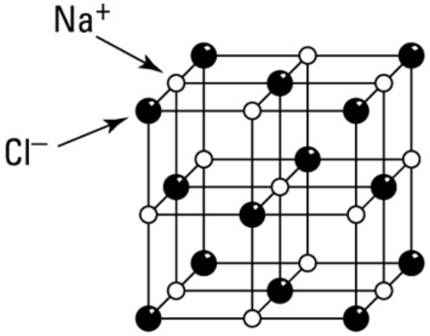
Topic 09

Ionic and Molecular Compounds: Formulas and Names

3.4 Ionic Compounds: Formulas and Names

- **Ionic compounds:** are composed of metal ions combined with nonmetals ions, bonded by "**an ionic bond**":
  - ✓ metal atoms lose electrons to form cations (+ions), while nonmetal atoms gain electrons to form anions (-ions).
- Many ionic compounds contain "**polyatomic ions**"**:** several atoms attached together by covalent bonds into one ion, carrying a specific charge, e.g. CO_3^{2-}, SO_4^{2-}, NH_4^{1+}, OH^{1-} ……

Note: Ionic compounds have no individual molecules, instead, they have a 3-dimensional array of cations and anions made of "**formula units**". e.g. NaCl is the formula unit of sodium chloride (not called a molecule)



Writing Formulas for Ionic Compounds (Name to Formula)

- Compound must have a total charge = 0, therefore we must balance the numbers of cations and anions in a compound to get "0" charge.

✓ Example: If Na^+ is combined with S^{2-}, you will need two Na^+ ions for every one S^{2-} ion to balance the charges, therefore the formula must be Na_2S

Steps to Write the Formula of Ionic Compounds

1. Write the symbol for the metal cation and its charge (use the periodic table to find both info).
2. Write the symbol for the nonmetal anion and its charge (use the periodic table to find both info).
3. Charge (without sign) of each ion becomes subscript for the other ion.
4. Whenever possible, reduce subscripts to smallest whole number ratio (to obtain the empirical formula, since this is an ionic compound).
5. Check that the total charge of the cations cancels the total charge of the anions!

Example: Write the formula of a compound composed of aluminum ions and oxide ions

1. Write the symbol for the metal cation and its charge.

   Al^{3+} (group 3A charge = +3)

2. Write the symbol for the nonmetal anion and its charge.

   O^{2-} (group 6A charge = –2)

3. Charge (without the sign) becomes subscript for other ion.

   $Al^{3+} \quad O^{2-}$

4. If possible, reduce subscripts to smallest whole number ratio (in this example, reduction is not possible).

   $\mathbf{Al_2O_3}$

5. Check that the total charge of cations cancels the total charge of anions.

   $Al = (2)\times(+3) = +6$

   $O = (3)\times(-2) = -6$

Naming the Binary Ionic Compounds (Formula to Name)

- Ionic compounds are made of cation and anion.
- Some ionic compounds have one or more common names (or commercial names) that are only learned by experience:
  - ➢ e.g. NaCl = table salt, $NaHCO_3$ = baking soda, CaO = lime
- Write systematic name by simply naming the ions:

✓ **If the cation is:**

- ➢ metal with invariant charge = metal name (as in the periodic table)
- ➢ metal with variable charge = metal name + (charge as a Roman numeral)
- ➢ polyatomic ion = name of polyatomic ion (see the table of polyatomic ions)

✓ **If the anion is:**

- ➢ nonmetal = stem of nonmetal name + ide
- ➢ polyatomic ion = name of polyatomic ion (see the table of polyatomic ions)

Naming Binary Ionic Compounds of Metals with Invariant Charge

- Contain metal cation + nonmetal anion
- Metal ion is written first in both formula and name.

1. Name metal cation first, name nonmetal anion second.
2. Cation name is the metal name (from the periodic table).
3. Nonmetal anion is named by changing the ending on the nonmetal name to **–ide**.

Metals with Invariant Charge

- **Metals with Invariant Charge (main-group metals):**
  - ✓ Metals whose ions can only have one possible charge
    - Groups $1A^{1+}$ & $2A^{2+}$, Al^{3+}, Sc^{3+}
  - ✓ Cation name = metal name (as written in the periodic table)

Metals Whose Charge Is Invariant from One Compound to Another

Metal	Ion	Name	Group Number
Li	Li^+	Lithium	1A
Na	Na^+	Sodium	1A
K	K^+	Potassium	1A
Rb	Rb^+	Rubidium	1A
Cs	Cs^+	Cesium	1A
Be	Be^{2+}	Beryllium	2A
Mg	Mg^{2+}	Magnesium	2A
Ca	Ca^{2+}	Calcium	2A
Sr	Sr^{2+}	Strontium	2A
Ba	Ba^{2+}	Barium	2A
Al	Al^{3+}	Aluminum	3A
Zn	Zn^{2+}	Zinc	*
Sc	Sc^{3+}	Scandium	*
Ag**	Ag^+	Silver	*

Example: Naming Binary Ionic Compound with Invariant Charge Metal: CsF

1. Identify both cation and anion:

   $Cs = Cs^{+}$, because it is in Group 1A

   $F = F^{-}$, because it is in Group 7A

2. Name the cation:

   Cs^{+} = cesium

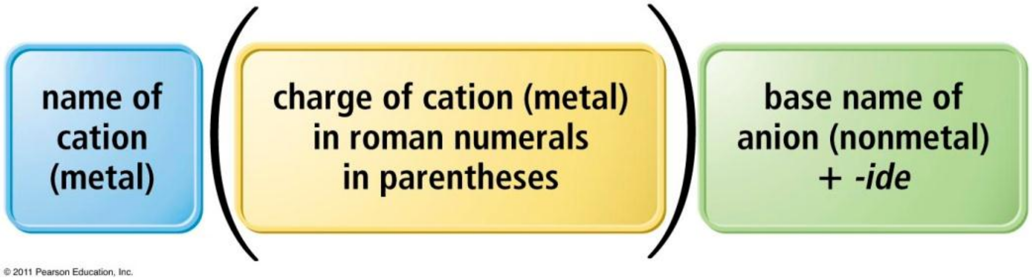
3. Name the anion:

   F^{-} = fluoride

4. Write the cation name first, then the anion name:

   Cesium fluoride

Naming Binary Ionic Compounds for Metals with Variable Charge



- Contain metal cation + nonmetal anion.
- Metal listed first in formula and name.

1. Name metal cation first, name nonmetal anion second.
2. Metal cation name is the metal name followed by a Roman number (in brackets) to indicate its charge e.g. (I) , (II) , (III), (IV) , (V) , (VI) , (VII) …
   - ✓ determine charge usinganion charge
   - ✓ See the table of common ions
3. Nonmetal anion is named by changing the ending on the nonmetal base name to **−ide**

Metals with variable Charges

- **Metals with Variable Charges (transition elements):**
- ✓ Metals whose ions can have more than one possible charge.
- ✓ Determine charge by charge on anion and cation.
- ✓ Name = metal name + (charge as a Roman numeral in parentheses).

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

Example: Naming Binary Ionic Compounds for Metals with Variable Charge: CuF_2

1. Identify cation and anion

   $F = F^-$ because it is in Group 7

   $Cu = Cu^{2+}$ to balance the two (−) charges from $2F^-$

2. Name the cation:

   Cu^{2+} = copper(II)

3. Name the anion:

   F^- = fluoride

4. Write the cation name first, then the anion name:

Copper(II) fluoride

Naming Binary Ionic Compounds Containing Polyatomic Ions

- ✓ Polyatomic ions are single ions that contain more than one atom.
- ✓ Often identified by parentheses around ion in formula.
- ✓ The Name and charge of polyatomic ion do NOT change.
- ✓ Name any ionic compound by naming cation first, and then anion.
- ✓ See the following table for the most common Polyatomic ions (memorize the highlighted ions!).

Common Polyatomic Ions

Name	Formula
acetate	$C_2H_3O_2^-$
carbonate	CO_3^{2-}
bicarbonate	HCO_3^-
hydroxide	OH^-
nitrate	NO_3^-
nitrite	NO_2^-
sulfate	SO_4^{2-}
sulfite	SO_3^{2-}
Phosphate	PO_4^{3-}
ammonium	NH_4^+

Name	Formula
hypochlorite	ClO^-
chlorite	ClO_2^-
chlorate	ClO_3^-
perchlorate	ClO_4^-
chromate	CrO_4^{2-}
dichromate	$Cr_2O_7^{2-}$
hydrogen sulfate (aka bisulfate)	HSO_4^-
hydrogen sulfite (aka bisulfite)	HSO_3^-

- **Important**: The ions indicted with this arrow are to be carefully memorized (names, formulas, and charges!)

Notice that the name of the polyatomic ion is written as is: no "ide" is added to its name!

Examples: Naming Binary Ionic Compounds Including Polyatomic Ions

Examples:

$FeSO_4$	iron(II) sulfate
NH_4NO_3	ammonium nitrate
$NaNO_2$	sodium nitrite

Exercise: Name The Following Compounds:

1. $Ca_3(PO_4)_2$ ....
2. $Mg(NO_2)_2$ ....
3. KNO_3 ....
4. $CuNO_2$ ....
5. $Ba(NO_3)_2$ ....
6. $MgSO_4$ ....

3.5 Molecular Compounds: Formulas and Names

- **Molecular compounds:** are composed of two or more nonmetals, bonded by "**covalent bonds**".

- 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 and name:
  - e.g. 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.

Naming The Binary Molecular Compounds

- Write the name of the element with the smallest group number first (e.g. Nitrogen before Oxygen)
- If the two elements lie in the same group, then write the element with the greatest row number first (e.g. Sulfur before Oxygen).
- The prefixes given to each element indicate the number of its atoms present.

Naming The Binary Molecular Compounds

prefix	name of 1st element	prefix	base name of 2nd element + *-ide*

Note: If there is <u>only one atom</u> of the <u>first element</u> in the formula, the prefix *<u>mono</u>* is normally <u>omitted</u> (not written).

Prefixes Used in Naming Binary Compounds Formed between Nonmetals

Prefix	Meaning
Mono-	*1*
Di-	*2*
Tri-	*3*
Tetra-	*4*
Penta-	*5*
Hexa-	*6*
Hepta-	*7*
Octa-	*8*
Nona-	*9*
Deca-	*10*

Example: Naming The Binary Molecular Compounds

Naming Molecular Compounds

Name each compound.

(a) NI_3 **(b)** PCl_5 **(c)** P_4S_{10}

SOLUTION

(a) The name of the compound is the name of the first element, *nitrogen*, followed by the base name of the second element, *iod*, prefixed by *tri-* to indicate three and given the suffix *-ide*.

NI_3 nitrogen triiodide

(b) The name of the compound is the name of the first element, *phosphorus*, followed by the base name of the second element, *chlor*, prefixed by *penta-* to indicate five and given the suffix *-ide*.

PCl_5 phosphorus pentachloride

(c) The name of the compound is the name of the first element, *phosphorus*, prefixed by *tetra-* to indicate four, followed by the base name of the second element, *sulf*, prefixed by *deca* to indicate ten and given the suffix *-ide*.

P_4S_{10} tetraphosphorus decasulfide

FOR PRACTICE

Name the compound N_2O_5.

FOR MORE PRACTICE

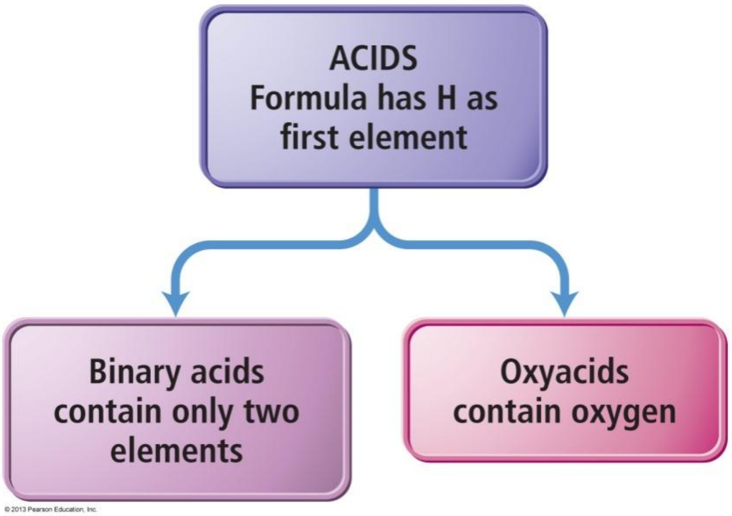
Write the formula for phosphorus tribromide.

Naming of Binary Acids

- Acids are molecular compounds that produce **H^+** when dissolved in water.
  - not named as "acids" if not dissolved in water.
  - to indicate the compound is dissolved in water, (*aq*) is written after the formula.
- Acids have sour taste.
- Acids react with many active metals:
  - such as Zn, Fe, Mg; but not with Au, Ag, Pt
- Acids formulas generally start with: H
  - e.g. HCl, HNO_3, HBr, HI, H_2SO_4

Types of Acids

- **Types of acids:**
  - Binary acids:
    - H^+ and nonmetal anion
  - Oxyacids:
    - H^+ cation and polyatomic anion



Forming an Acid by Adding H^+ to The Anion

- "ide" and "ate" suffixes of the anions change to "ic":

  - $NO_3^- + H^+ \rightarrow HNO_3$

    Nitrate ion — Nitric acid

  - $Cl^- + H^+ \rightarrow HCl$

    Chloride ion — Hydrochloric acid

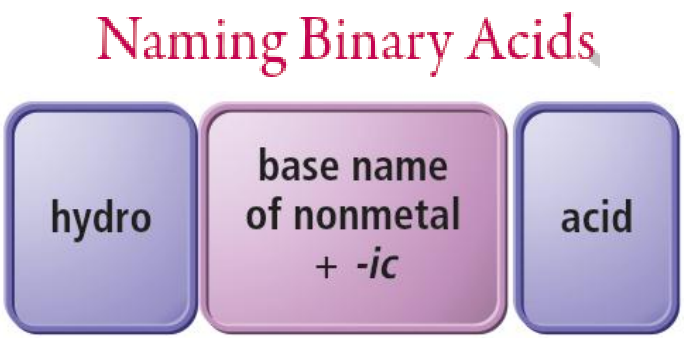
Examples: Name the following binary acids:

HF: Hydrofluoric acid

HCl: Hydrochloric acid

HBr: Hydrobromic acid

HI: Hydroiodic acid



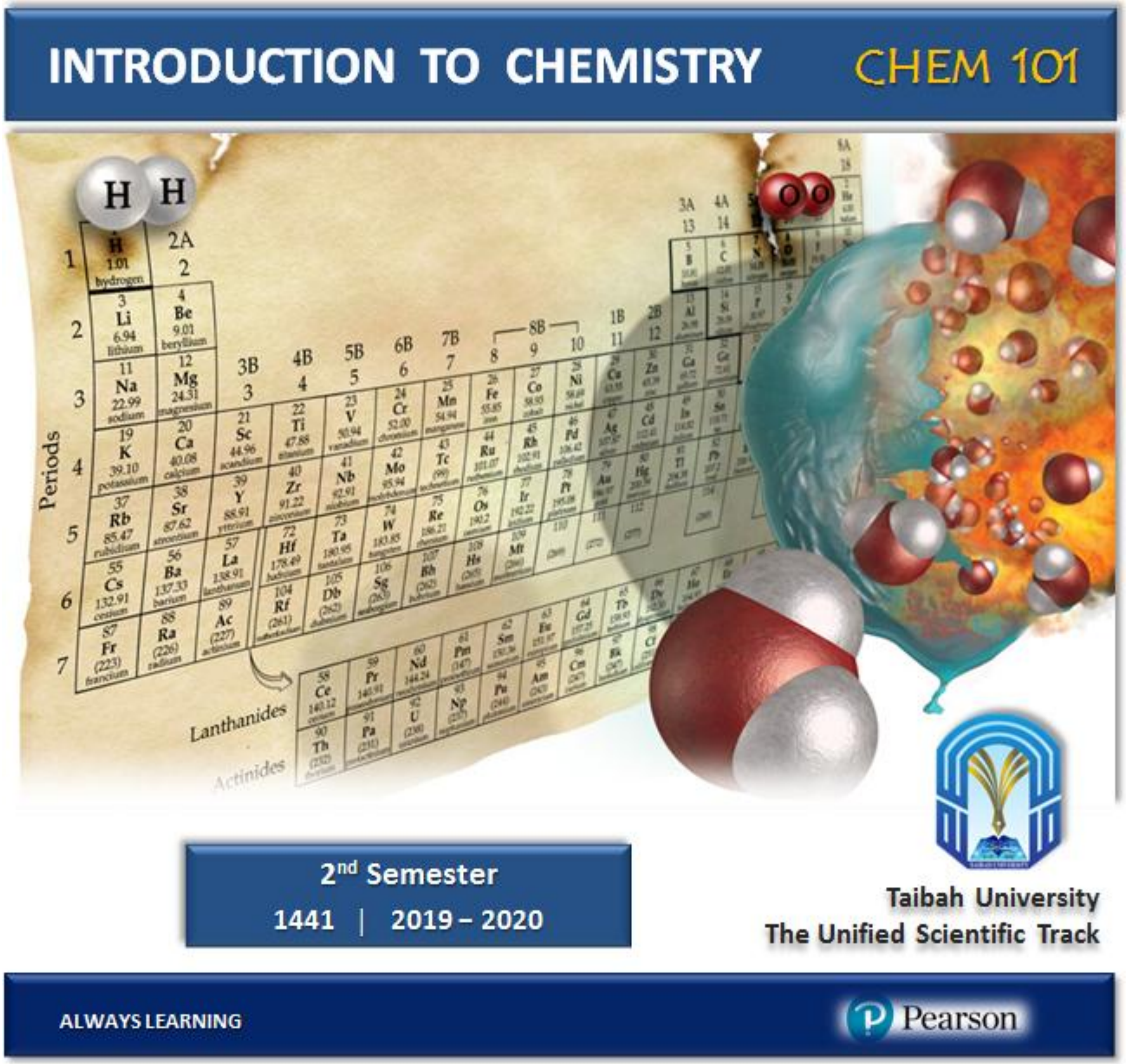
Assessment: **Write the missing names and formulas of the following compounds, and mention the type of each of them (ionic, molecular or acid):**

Section:

ID:

Name:

Formula	Name	Compound Type	Formula	Name	Compound Type
	Calcium chloride			Carbon dioxide	
Fe_2O_3				Hydrochloric acid	
LiF			P_2O_5		
	Copper(I) bromide			Nitrogen trifluoride	
$Mg(NO_3)_2$			N_2O_4		
Fe_2S_3			CO		
CuS			HBr		
	Sodium nitrite		Na_3PO_4		
Al_2S_3				Dihydrogen monoxide	
	Aluminium hydroxide		N_2O		
Cu_2CO_3			H_2S		
$Fe(HCO_3)_3$				Dinitrogen tetrahydride	
	Ammonium sulfite		S_2F_{10}		
NH_4NO_3			SF_6		
$Al_2(SO_4)_3$				Nitrogen monoxide	
	Radium hydroxide		SrF_2		
	Barium bicarbonate		N_2O_3		
FeF_2			$Ca(OH)_2$		
$Cu(NO_3)_2$				Iron(II) oxide	



Lecture Presentation

Chapter 3

Molecules, Compounds, and Chemical Equations

Topic 10

- **Molar Mass:** **Counting Atoms by Weighing Them**
- **The Mole:** **The "Chemist's Dozen"**

3.7 Molar Mass: Counting Atoms by Weighing Them

- Definition of the "Mole"
- The Molar Mass of a Compound
- Mole Conversions

What is the “Mole”?

- Chemistry is quantitative in nature:
  - Its unit is the “**mole**” = the chemist’s “dozen”
- **The Mole** as a unit **vs. “Dozen”** as an unit:
  - The unit “dozen” is associated with 12 units.
  - The unit “mole” is associated with 6.022×10^{23} particles.
- **There is Avogadro’s number** of particles in every mole of substance:
  - 6.022×10^{23} particles is known as Avogadro’s number.
  - 1 mole = 6.022×10^{23} particles.
- **1 mole of Cu atoms has**:
  - an atomic mass of 63.55 g, which contains:
  - 6.022×10^{23} Cu atoms.

Formula Mass & Molar Mass

- **Formula Mass (amu):** The mass of an individual molecule or formula unit, expressed in “amu” (atomic mass unit)
  - ✓ also known as molecular mass or molecular weight.
  - ✓ Sum of the masses of the atoms in a single molecule or formula unit

Formula mass of H_2O = [2 × (1.01 amu H)] **+** [1 × (16.00 amu O)] = 18.02 **amu**

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

- **Molar Mass (g/mol):** The mass of one mole of a substance, expressed in “g/mol”
  - ✓ **Molar mass is numerically equal to formula mass, but expressed in g/mol**

Molar mass of H_2O = 18.02 **g/mol**

How to Calculate the Molar Mass of Any Substance?

- To calculate the molar mass of any substance, you must first know its exact chemical formula!
- The molar mass can be calculated for any substance by summation of the atomic masses (*from the periodic table*) of all the atoms of elements present this substance's formula:

> An element's molar mass in grams per mole (g/mol) is numerically equal to the element's atomic mass in atomic mass units (amu).

✓ **Examples: Calculate the molar mass (g/mol) for:**

- Molar mass of water (**H_2O**) = $(2\times1) + (1\times16) = 18$ g/mol

- Molar mass of oxygen (**O_2**) = $2\times16 = 32$ g/mol

- Molar mass of **NaCl** = $(1\times23) + (1\times35) = 58$ g/mol

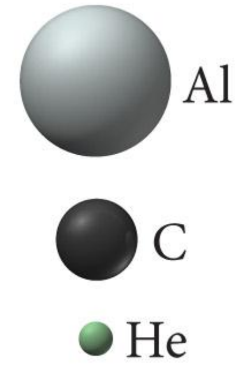
- Molar mass of glucose (**$C_6H_{12}O_6$**) = $(12\times6) + (12\times1) + (16\times6) = 180$ g/mol

Mole, Mass, and Number of Particles

26.98 g aluminum = 1 mol aluminum = 6.022×10^{23} Al atoms

12.01 g carbon = 1 mol carbon = 6.022×10^{23} C atoms

4.003 g helium = 1 mol helium = 6.022×10^{23} He atoms



✓ **Exercises: Calculate the molar mass (g/mol) for each substance:**

- $MgBr_2$
- CuF_2
- $Ca_3(PO_4)_2$
- Ozone gas (O_3)
- Nitrogen gas
- $(NH_4)_2SO_4$
- C_8H_{18}
- $Al_2(CO_3)_3$

Mole Conversions

➢ Converting Between: Mole, Mass and Number of Particles:

Mole Conversions: Grams of Substance $\xrightarrow{\div \text{ Molar Mass}}$ Moles of Substance $\xrightarrow{\times \text{ Avo. Number}}$ Number of Atoms or Molecules

Number of Atoms or Molecules $\xrightarrow{\div \text{ Avo. Number}}$ Moles of Substance $\xrightarrow{\times \text{ Molar Mass}}$ Grams of Substance

Important Note: to calculate the number of **particles** in a given **mass (g)** of a substance (or vice versa); you need first to calculate the number of **moles** of that substance, by using one of the mole conversion relationships!

✓ **In other words**: to go from **mass (g)** to **number of particles** (atoms or molecules); you must go through the "**mole**".

Notes:

- ✓ **Avogadro's Number** = 6.022 x 10^{23} atoms/mol (or molecule/mol)
- ✓ The mass of substance must be in (**grams**), if not, convert it first.
- ✓ **Molar Masses** can be calculated using the periodic table.

Mole Conversions: Number of Particles – Moles

Converting between Number of Moles and Number of Atoms

Calculate the number of copper atoms in 2.45 mol of copper.

SORT You are given the amount of copper in moles and asked to find the number of copper atoms.

GIVEN: 2.45 mol Cu
FIND: Cu atoms

STRATEGIZE Convert between number of moles and number of atoms using Avogadro's number as a conversion factor.

CONCEPTUAL PLAN

$$\text{mol Cu} \xrightarrow{\frac{6.022 \times 10^{23}\ \text{Cu atoms}}{1\ \text{mol Cu}}} \text{number of Cu atoms}$$

RELATIONSHIPS USED
$6.022 \times 10^{23} = 1$ mol (Avogadro's number)

SOLVE Follow the conceptual plan to solve the problem. Begin with 2.45 mol Cu and multiply by Avogadro's number to get to the number of Cu atoms.

SOLUTION

$$2.45\ \cancel{\text{mol Cu}} \times \frac{6.022 \times 10^{23}\ \text{Cu atoms}}{1\ \cancel{\text{mol Cu}}}$$

$$= 1.48 \times 10^{24}\ \text{Cu atoms}$$

CHECK Since atoms are small, it makes sense that the answer is large. The given number of moles of copper is almost 2.5, so the number of atoms is almost 2.5 times Avogadro's number.

FOR PRACTICE 2.9
A pure silver ring contains 2.80×10^{22} silver atoms. How many moles of silver atoms does it contain?

Mole Conversions: Mass – Moles

Problem: How many Mg atoms are in 0.20 g of Mg?

- ✓ Mg has a molar mass of 24.3 g/mol (from the periodic table)
- ✓ You need to convert 0.20 g Mg to moles of Mg:

$$0.20 \text{ g Mg} \times (1 \text{ mol}/24.3 \text{ g}) = 8.23 \times 10^{-3} \text{ mol of Mg}$$

- ✓ Now that you know the moles of Mg, you can determine the number of Mg atoms:

$$8.23 \times 10^{-3} \text{ mol Mg} \times (6.022 \times 10^{23} \text{ atoms/mol of Mg})$$

$$= 4.95 \times 10^{21} \text{ atoms of Mg}$$

Mole Conversions: Mass – Number of Particles

The Mole Concept—Converting between Mass and Number of Atoms

How many copper atoms are in a copper penny with a mass of 3.10 g? (Assume that the penny is composed of pure copper.)

SORT You are given the mass of copper and asked to find the number of copper atoms.

GIVEN: 3.10 g Cu
FIND: Cu atoms

STRATEGIZE Convert between the mass of an element in grams and the number of atoms of the element by first converting to moles (using the molar mass of the element) and then to number of atoms (using Avogadro's number).

CONCEPTUAL PLAN

$$\boxed{\text{g Cu}} \xrightarrow{\frac{1\ \text{mol Cu}}{63.55\ \text{g Cu}}} \boxed{\text{mol Cu}} \xrightarrow{\frac{6.022 \times 10^{23}\ \text{Cu atoms}}{1\ \text{mol Cu}}} \boxed{\text{number of Cu atoms}}$$

RELATIONSHIPS USED

$63.55\ \text{g Cu} = 1\ \text{mol Cu}$ (molar mass of copper)
$6.022 \times 10^{23} = 1\ \text{mol}$ (Avogadro's number)

SOLVE Follow the conceptual plan to solve the problem. Begin with 3.10 g Cu and multiply by the appropriate conversion factors to arrive at the number of Cu atoms.

SOLUTION

$$3.10\ \cancel{\text{g Cu}} \times \frac{1\ \cancel{\text{mol Cu}}}{63.55\ \cancel{\text{g Cu}}} \times \frac{6.022 \times 10^{23}\ \text{Cu atoms}}{1\ \cancel{\text{mol Cu}}} = 2.94 \times 10^{22}\ \text{Cu atoms}$$

CHECK The answer (the number of copper atoms) is less than 6.022×10^{23} (one mole). This is consistent with the given mass of copper atoms, which is less than the molar mass of copper.

FOR PRACTICE 2.11
How many carbon atoms are there in a 1.3-carat diamond? Diamonds are a form of pure carbon. (1 carat = 0.20 g)

FOR MORE PRACTICE 2.11
Calculate the mass of 2.25×10^{22} tungsten atoms.

Mole Conversions: Mass – Number of Particles

Problem: How many grams of CO_2 are there in 6.75×10^{22} molecules of CO_2?

Strategy:

1. You need to know the molar mass of CO_2 (calculate it using the atomic masses from the periodic table)
2. Convert number of molecules of CO_2 to number of moles.
3. Convert number of moles of CO_2 to grams (mass) of CO_2.

The answer ….

The Answer:

Step 1: CO_2 has a molar mass of 44.0 g/mol:

- There is 1 mole carbon in every mole of CO_2

1 mole carbon = 12.0 g (from periodic table)

- There are 2 moles oxygen in every mole of CO_2

2 mole oxygen = 2 × 16.0 = 32.0 g (from periodic table)

- Therefore, the molar mass of CO_2 = 12 + 32 = **44.0 g/mol**

Step 2: Determine the number of moles of CO_2 :

6.75×10^{22} molecules CO_2 x (1 mole CO_2/6.022×10^{23} molecules)

= **1.12×10^{-1} mol of CO_2**

Step 3: Determine the grams of CO_2 :

1.12×10^{-1} moles CO_2 x (44.0 g/1 mol CO_2) = **4.93 g of CO_2**

Assessment

1- How many **moles** of H_2O are there in **100 g** H_2O?

2- Calculate the number of iron **atoms** present in a **4 g** piece of iron.

3- How many CO **molecules** are there in **2.67 moles** of CO?

4- How many **moles** of NH_3 are there in **0.2 Kg** of NH_3?

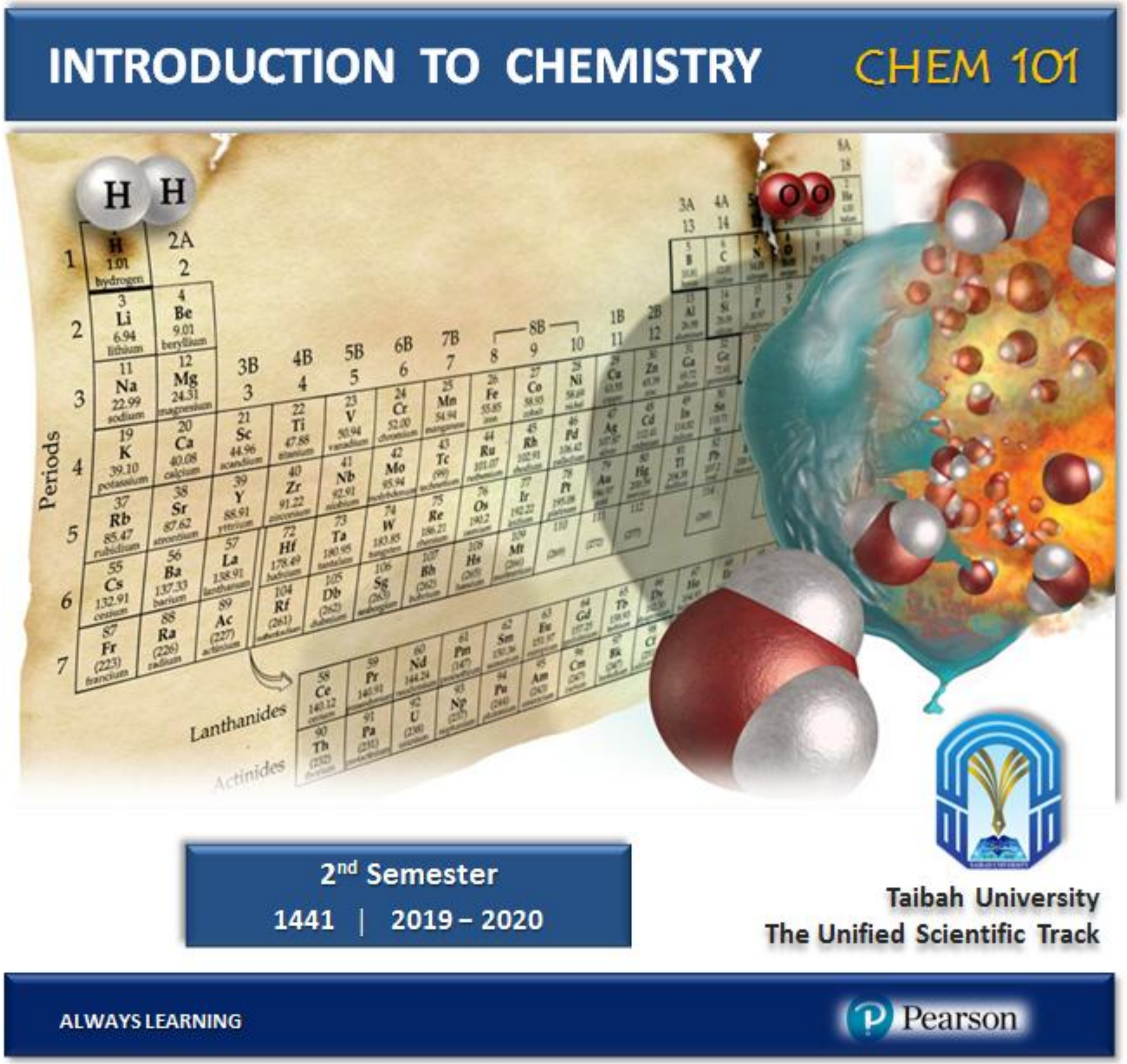
5- What is the **mass (g)** of 4.3×10^{24} **atoms** of silver?

6- Calculate the number of oxygen **molecules** in **250 g** oxygen.

7- What is the **mass (g)** of 9.2×10^{23} **particles** of $Al_2(CO_3)_3$?

Mole Conversions:

$$\boxed{\text{Grams of Substance}} \underset{\times \text{ Molar Mass}}{\overset{\div \text{ Molar Mass}}{\rightleftarrows}} \boxed{\text{Moles of Substance}} \underset{\div \text{ Avo. Number}}{\overset{\times \text{ Avo. Number}}{\rightleftarrows}} \boxed{\text{Number of Atoms or Molecules}}$$



Lecture Presentation

Chapter 3

Molecules, Compounds, and Chemical Equations

Topic 11

- **Calculating Mass Percent**
- **Calculating Empirical & Molecular Formulas**
- **Balancing Chemical Equations**

3.6 Composition of Compounds: Mass Percent

- A chemical formula, in combination with the molar masses of its constituent elements, indicates the relative quantities of each element in a compound.
- The mass percentage (%) 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.

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

- The percentages may not always total to exactly 100% due to rounding!

3.6 Composition of Compounds: Practice

Example: Calculate the mass percent composition of "Cl" in the chlorofluorocarbon **CCl_2F_2**

$$\text{Mass percent Cl} = \frac{2 \times \text{Molar mass Cl}}{\text{Molar mass } CCl_2F_2} \times 100\%$$

Answer:

$$\begin{aligned}\text{Molar mass} &= 12.01\text{ g/mol} + 2(35.45\text{ g/mol}) + 2(19.00\text{ g/mol}) \\ &= 120.91\text{ g/mol}\end{aligned}$$

So the mass percent of Cl in CCl_2F_2 is

$$\begin{aligned}\text{Mass percent Cl} &= \frac{2 \times \text{molar mass Cl}}{\text{molar mass } CCl_2F_2} \times 100\% \\ &= \frac{2 \times 35.45\ \cancel{\text{g/mol}}}{120.91\ \cancel{\text{g/mol}}} \times 100\% \\ &= 58.64\%\end{aligned}$$

Exercise: Calculate the mass% of C in **$C_{12}H_{24}O_{12}$**

EXAMPLE **Mass Percent Composition**

Calculate the mass percent of Cl in Freon-112 ($C_2Cl_4F_2$), a CFC refrigerant.

SORT You are given the molecular formula of Freon-112 and asked to find the mass percent of Cl.	**GIVEN** $C_2Cl_4F_2$ **FIND** mass percent Cl
STRATEGIZE The molecular formula indicates that there are 4 mol of Cl in each mole of Freon-112. Find the mass percent composition from the chemical formula by using the equation that defines mass percent. The conceptual plan shows how you can use the mass of Cl in 1 mol of $C_2Cl_4F_2$ and the molar mass of $C_2Cl_4F_2$ to determine the mass percent of Cl.	**CONCEPTUAL PLAN** $\text{Mass \% Cl} = \dfrac{4 \times \text{molar mass Cl}}{\text{molar mass } C_2Cl_4F_2} \times 100\%$ **RELATIONSHIPS USED** $\text{Mass percent of element X} = \dfrac{\text{mass of element X in 1 mol of compound}}{\text{mass of 1 mol of compound}} \times 100\%$
SOLVE Calculate the necessary parts of the equation and substitute the values into the equation to find mass percent Cl.	**SOLUTION** $4 \times \text{molar mass Cl} = 4(35.45\text{ g/mol}) = 141.8\text{ g/mol}$ $\text{Molar mass } C_2Cl_4F_2 = 2(12.01\text{ g/mol}) + 4(35.45\text{ g/mol}) + 2(19.00\text{ g/mol})$ $= 24.02\text{ g/mol} + 141.8\text{ g/mol} + 38.00\text{ g/mol} = 203.8\text{ g/mol}$ $\text{Mass \% Cl} = \dfrac{4 \times \text{molar mass Cl}}{\text{molar mass } C_2Cl_4F_2} \times 100\%$ $= \dfrac{141.8\ \cancel{\text{g/mol}}}{203.8\ \cancel{\text{g/mol}}} \times 100\%$ $= 69.58\%$

➢ **Empirical Formula:**

- Simplest whole-number ratio of the atoms of elements in a compound.
- The empirical formula of a compounds can be determined from the results of elemental analysis (in laboratory).
  - Masses of elements formed when a compound is decomposed, or that react together to form a compound.
  - Percent composition.

Note: An empirical formula represents a ratio of atoms or a ratio of moles of atoms, NOT a ratio of masses (g)!

Obtaining Empirical Formula from Experimental Data

1. **Convert the percentages to <u>grams</u>:**
   a) If not given, assume you start with 100 g of the compound.
   b) Example: 24.5% C means 24.5 g C.
2. **Convert mass (in grams) to <u>moles</u> (using mole conversions):**
   a) Use molar mass of each element.
   b) Example: 24.5 g C × (1 mol C/12.01 grams) = 2.0 mol C.
3. **Divide all by the smallest number of moles to obtain the <u>atom-to-atom ratio</u> for each of the elements in the compound:**
   a) If the result is within 0.1 of a whole number, round to the whole number.
4. **Multiply all mole ratios by a number to make all <u>whole numbers:</u>**
   a) If the ratio is 0.5, multiply all by 2; If the ratio is 0.25, multiply all by 4; if the ratio is 0.33 or 0.66, multiply all by 3; and so on …..
   b) Skip step 4 if already whole numbers.

Obtaining Empirical Formula from Experimental Data: Practice

Problem: A compound containing nitrogen and oxygen is decomposed in the laboratory and produced 24.5 g nitrogen and 70.0 g oxygen. **Calculate its empirical formula**.

Given: 24.5 g N, 70.0 g O , **Find**: empirical formula

Answer:

1- Convert each of the masses in step 1 to moles by using the appropriate molar mass for each element as a conversion factor:

$$24.5\ \cancel{\text{g N}} \times \frac{1\ \text{mol N}}{14.01\ \cancel{\text{g N}}} = 1.75\ \text{mol N}$$

$$70.0\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 4.38\ \text{mol O}$$

2- Write down a pseudo formula for the compound using the number of moles of each element (from step 1) as subscripts: $\mathbf{N_{1.75}O_{4.38}}$

3- Divide all the subscripts in the formula by the smallest subscript: $N_{\frac{1.75}{1.75}}O_{\frac{4.38}{1.75}} \rightarrow N_1O_{2.5}$

4- If the subscripts are not whole numbers, multiply all the subscripts by a small whole number to get whole-number subscripts:

$$N_1O_{2.5} \times 2 \rightarrow N_2O_5$$

The correct empirical formula is N_2O_5.

Obtaining Empirical Formula from Experimental Data: Exercise

- A laboratory analysis of aspirin determined the following mass percent composition:

  C = 60%

  H = 4.48%

  O = 35.52%

Find the empirical formula of aspirin.

Hint: Since you are NOT given the real mass of each element, but given their % instead, then assume you started the analysis with **a 100 g** sample!

This means: In a 100 g sample of aspirin:

60% C = 60 g (C) , 4.48% H = 4.48 g (H) , 35.52% O = 35.52 g (O)

Calculating Molecular Formula from Empirical Formula

- The molecular formula is a multiple of the empirical formula.
  - It is the actual formula of the compound.
- To determine the molecular formula, you need to know the empirical formula and the molar mass of the compound (i.e. the molar mass of the molecular formula):

$$\boxed{\text{Molecular formula} = \text{empirical formula} \times n, \text{ where } n = 1, 2, 3, \ldots}$$

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

Note: Usually, you are given the molar mass of molecular formula but need to calculate the molar mass of the empirical formula (from the given empirical Formula itself!)

Calculating Molecular Formula from Empirical Formula: Example

Calculating a Molecular Formula from an Empirical Formula and Molar Mass

Butanedione—a main component responsible for the smell and taste of butter and cheese—contains the elements carbon, hydrogen, and oxygen. The empirical formula of butanedione is C_2H_3O and its molar mass is 86.09 g/mol. Find its molecular formula.

SORT You are given the empirical formula and molar mass of butanedione and asked to find the molecular formula.	**GIVEN:** Empirical formula $= C_2H_3O$ molar mass $= 86.09$ g/mol **FIND:** molecular formula
STRATEGIZE A molecular formula is always a whole-number multiple of the empirical formula. Divide the molar mass by the empirical formula mass to get the whole number.	Molecular formula $=$ empirical formula $\times\ n$ $n = \dfrac{\text{molar mass}}{\text{empirical formula molar mass}}$
SOLVE Calculate the empirical formula mass. Divide the molar mass by the empirical formula mass to find *n*. Multiply the empirical formula by *n* to obtain the molecular formula.	Empirical formula molar mass $= 2(12.01\text{ g/mol}) + 3(1.008\text{ g/mol}) + 16.00\text{ g/mol} = 43.04\text{ g/mol}$ $n = \dfrac{\text{molar mass}}{\text{empirical formula mass}} = \dfrac{86.09\ \cancel{\text{g/mol}}}{43.04\ \cancel{\text{g/mol}}} = 2$ Molecular formula $= C_2H_3O \times 2 = C_4H_6O_2$

Chemical Reactions

- **Chemical Reactions** involve chemical changes in matter resulting in new substances.

- Reactions involve rearrangement and exchange of atoms to produce new molecules.
  - elements cannot be changed into other elements in a chemical reaction.

$$\boxed{\text{Reactants} \quad \rightarrow \quad \text{Products}}$$

Chemical Equations

- **Chemical Equation:** is a shorthand way of describing a chemical reaction.

- Provides some basic 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
  - Example: $CH_4(g) + \mathbf{2}\ O_2(g) \rightarrow CO_2(g) + \mathbf{2}\ H_2O(g)$

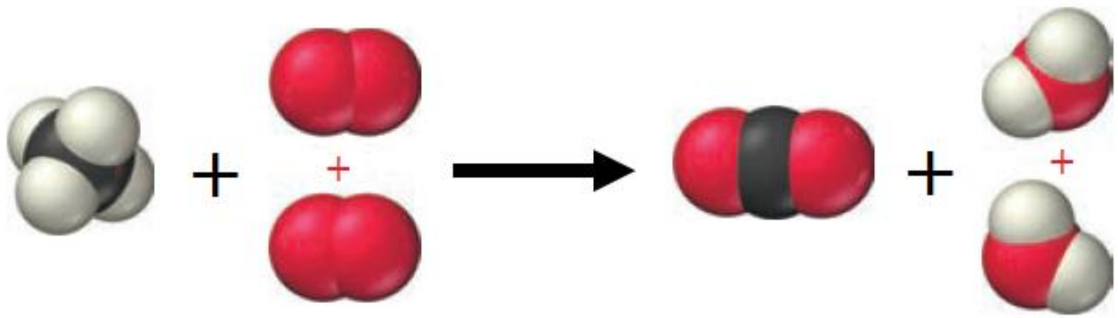
Symbols Used in Chemical Equations

- **State Symbols** (written after the substance's formula):
  - (g) = gas
  - (l) = liquid
  - (s) = solid
  - (aq) = aqueous = dissolved in water

- **Energy Symbols** (Written above the arrow):
  - Δ = heat
  - $h\nu$ = light
  - shock = mechanical
  - elec = electrical

- To show 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:

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



$$1\,C + 4\,H + 4\,O \longrightarrow 1\,C + 4\,H + 4\,O$$

- When aluminum metal reacts with oxygen, it produces a white powdery compound, aluminum oxide

aluminum(*s*) + oxygen(*g*) → aluminum oxide(*s*)

$$\ldots\ldots\ Al(s) + \ldots\ldots\ O_2(g) \rightarrow \ldots\ldots\ Al_2O_3(s)$$

$$\underline{4}\ Al(s) + \underline{3}\ O_2(g) \rightarrow \underline{2}\ Al_2O_3(s)$$

✓ The **coefficients** required to balance this equation are: **4, 3, 2**, respectively.

Assessment

1- **Give the coefficients that are necessary to balance each of the following equations:**

$$H_2(g) + Cl_2(g) \rightarrow HCl(g)$$

$$Cu_2O(s) + C(s) \rightarrow Cu(s) + CO(g)$$

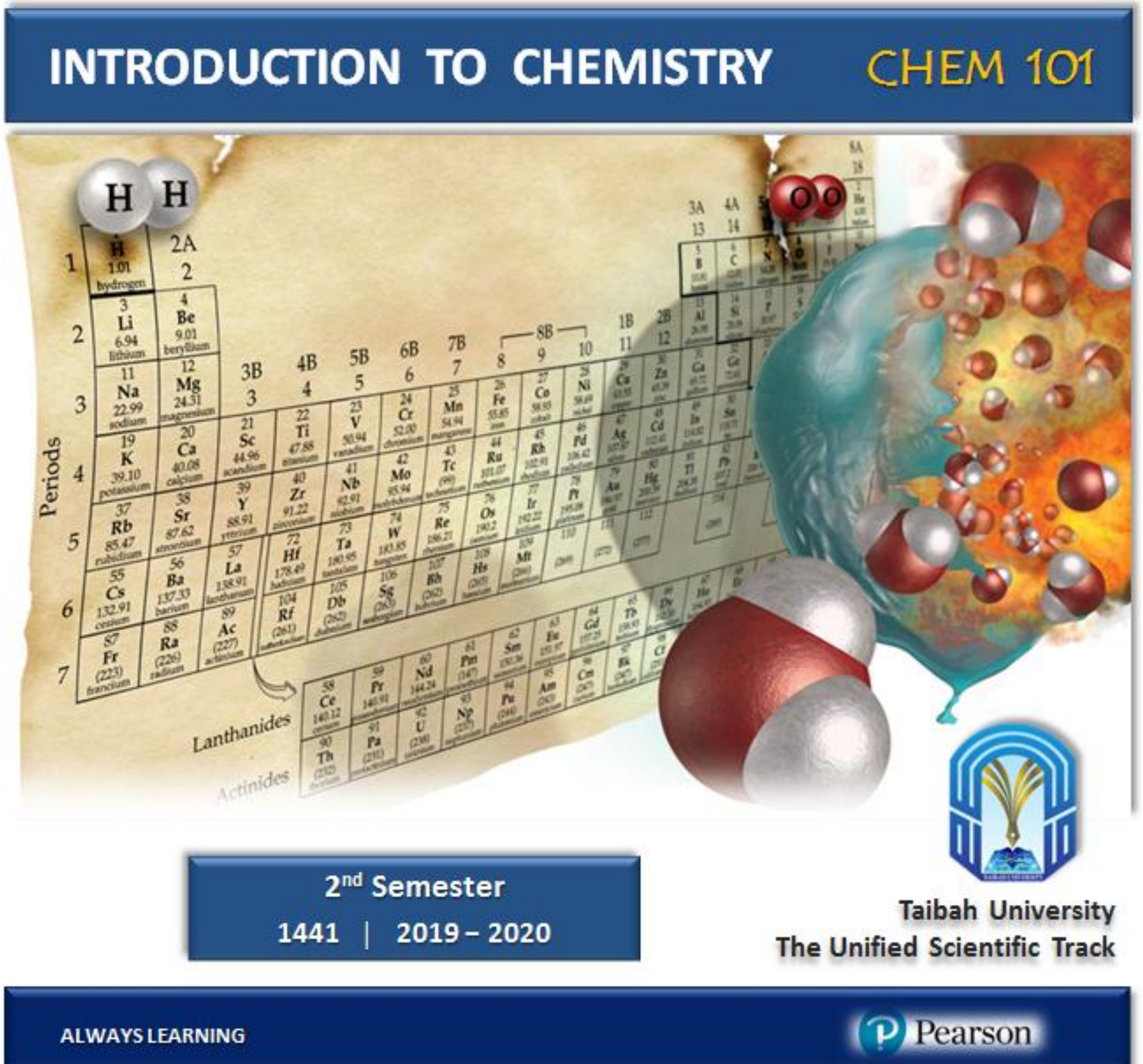
$$Co(NO_3)_3(aq) + (NH_4)_2S(aq) \rightarrow Co_2S_3(s) + NH_4NO_3(aq)$$

$$Co_2O_3(s) + C(s) \rightarrow Co(s) + CO_2(g)$$

2- **What is the coefficient of H_2O when each of the following equations are balanced?**

$$CO_2(g) + CaSiO_3(s) + H_2O(l) \rightarrow SiO_2(s) + Ca(HCO_3)_2(aq)$$

$$C_4H_{10}(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$



Lecture Presentation

Chapter 3

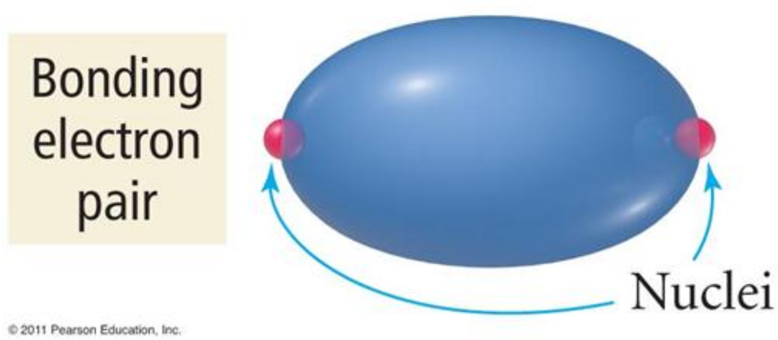
Molecules, Compounds, and Chemical Equations

Topic 12

- Chemical Bonding
- Lewis Structures
- Lattice Energy
- Bond Polarity
- Bond Energy & Length

3.10 Chemical Bonds

- Compounds are made of atoms held together by bonds.
- **Chemical bonds** are forces of attraction between atoms.
- The bonding attraction comes from attractions between protons and electrons of bonded atoms.
- Chemical bonds form because they lower the potential energy between the charged particles that compose atoms.



Types of Chemical Bonds

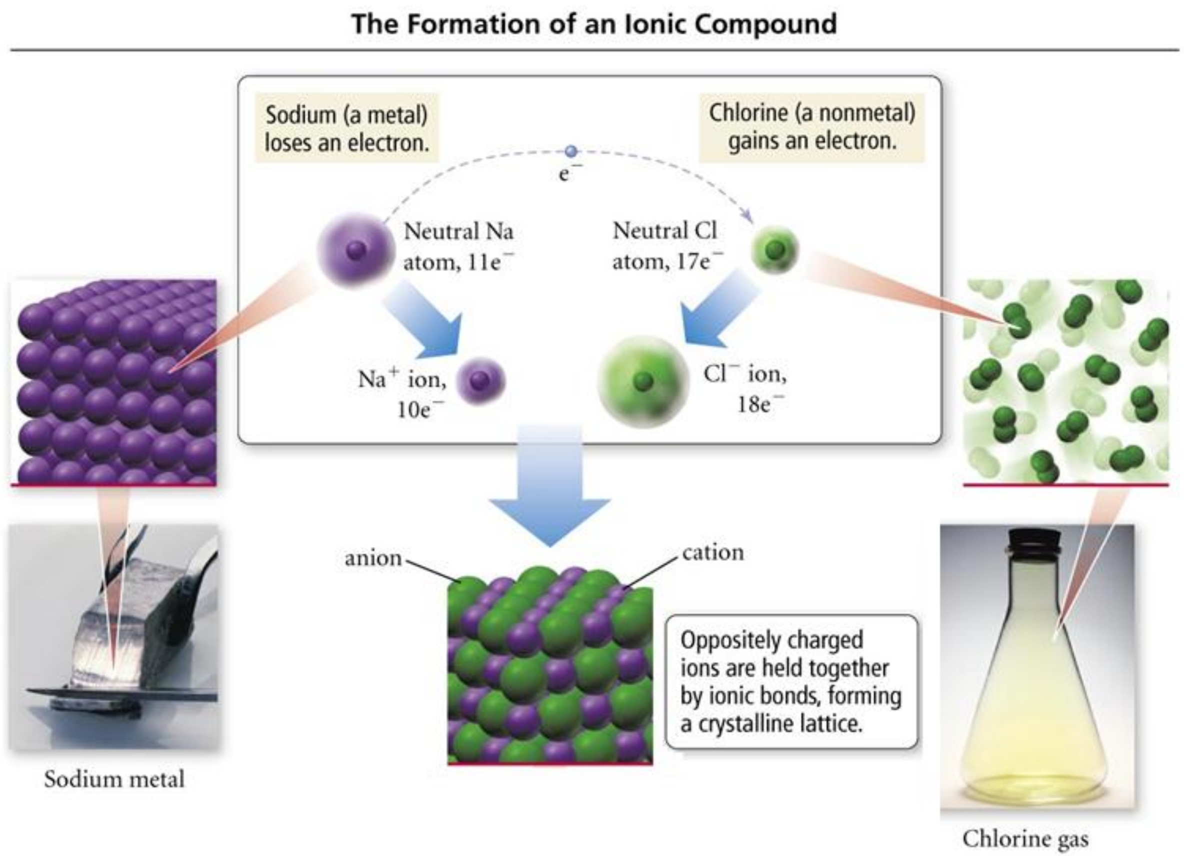
- **Chemical bonds** can be classified into <u>three types</u>, depending on the <u>types of atoms</u> involved in the bonding:
  - **Ionic bond**
  - **Covalent bond**
  - **Metallic bond**

Types of Atoms	Type of Bond	Characteristic of Bond
Metal and nonmetal	Ionic	Electrons transferred
Nonmetal and nonmetal	Covalent	Electrons shared
Metal and metal	Metallic	Electrons pooled

Types of Chemical Bonds: The Ionic Bond

Ionic bond: results when electrons have been transferred between atoms, resulting in oppositely charged ions that attract each other:

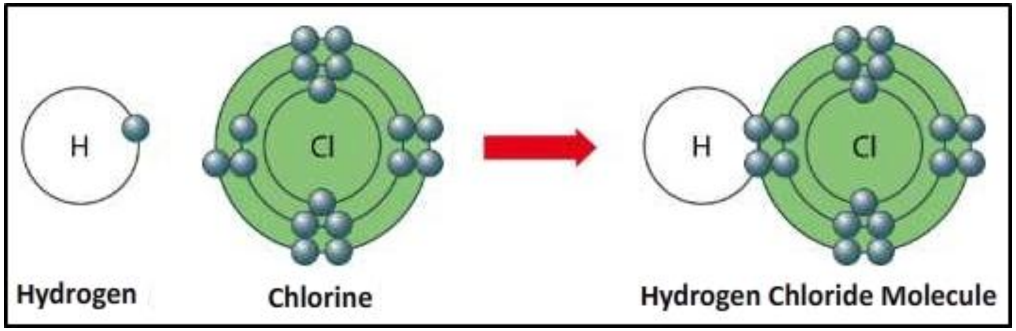
- Generally formed when metal atoms (cations) bond to nonmetal atoms (anions)
- **Method**: electron transfer



Types of Chemical Bonds: The Covalent Bond

Covalent bond: results when two atoms share some of their electrons:

- Generally formed when nonmetal atoms bond together
- Shared electrons hold the atoms together by attracting nuclei of both atoms.
- **Method**: electron sharing



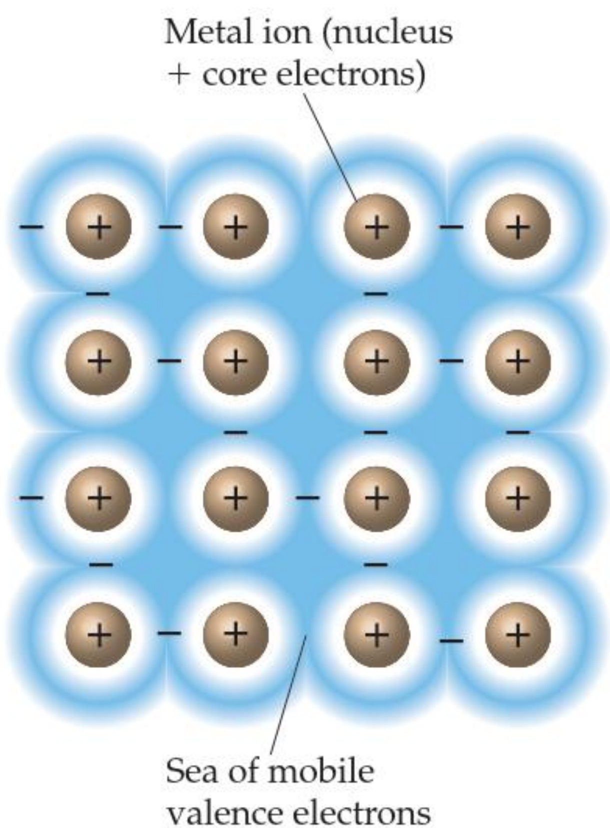
- **Multiple Covalent Bonds:**
  - ✓ **Single covalent bond**: A covalent bond formed by sharing one electron pair.
    Represented by a single line: **H–H**
  - ✓ **Double covalent bond**: formed by sharing two electron pairs.
    Represented by a double line: **O = O**
  - ✓ **Triple covalent bond**: formed by sharing three electron pairs.
    Represented by a triple line: **N ≡ N**

Types of Chemical Bonds: The Metallic Bond

Metallic Bond: occurs in metals. Since metals have low ionization energies, they tend to lose electrons easily, forming an electron sea, in which, all of the atoms in a metal lattice pool (release) their valence electrons (delocalize).

- Metallic bonding results from attraction of cation to the delocalized electrons.

- **Method**: electron pooling (electron sea of delocalized electrons)



- **Lewis Structures**: simple diagrams to visualize the number of valence electrons in atoms of main-group elements by dots. The dots are placed around the element's symbol with a maximum of two dots per side. Each dot represents one valence electron.
  - ✓ **Remember**: the number of valence electrons for main group element is equal to the group number of the element (except for helium, which is in group 8A but has only two valence electrons).
  - ✓ **Note**: While the exact location of dots is not critical, here we first place dots singly before pairing (except for helium which always has two paired dots)

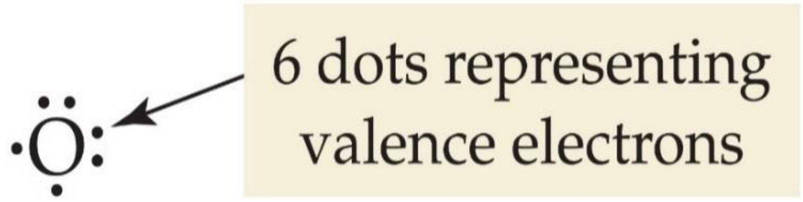
3.11 Representing Valence Electrons with Dots (Lewis Structures)

- The electron configuration of Oxygen is as follows:

$$1s^2\,\boxed{2s^2 2p^4}$$

6 valence electrons

- Its **Lewis structure** is as follows:

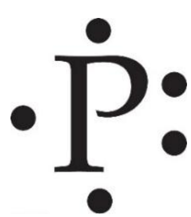


3.11 Representing Valence Electrons with Dots (Lewis Structures)

- **Lewis structure** for all period 2 elements:

- **Practice**: Draw the Lewis dot structure of a phosphorus atom.

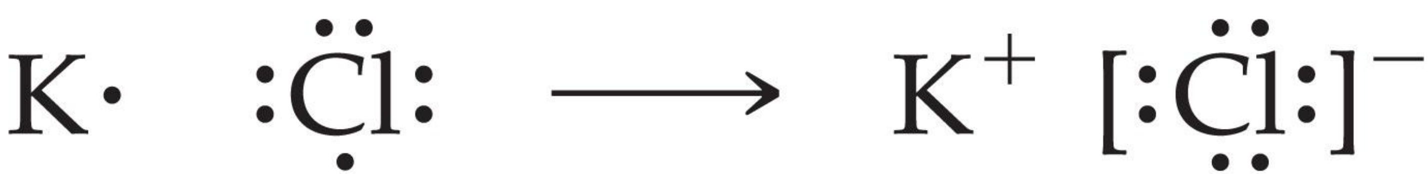
- **Solution**: Since phosphorus is in Group 5A in the periodic table, it has 5 valence electrons. Represent these as five dots surrounding the symbol for phosphorus:



3.12 Lewis Structures: For Ionic Bonding

- For example, **potassium** and **chlorine** have the Lewis structures:

- When potassium and chlorine bond, potassium **transfers** its valence electron to chlorine (forming octet Cl):



3.12 Lewis Structures: For Ionic Bonding

- Consider the ionic compound formed between sodium and sulfur. The Lewis structures for **sodium** and **sulfur** are as follows:

- Sodium must **lose** one valence electron to obtain an octet (in the previous principal shell), while sulfur must **gain** two electrons to obtain an octet.
- The compound that forms between sodium and sulfur requires two sodium atoms to every one sulfur atom. The Lewis structure is as follows:

- The correct chemical formula is Na_2S.

3.12 Lewis Structures: For Ionic Bonding

EXAMPLE 3.19 Using the Lewis Model to Predict the Chemical Formula of an Ionic Compound

Use the Lewis model to predict the formula for the compound that forms between calcium and chlorine.

SOLUTION

Draw Lewis structures for calcium and chlorine based on their respective number of valence electrons, determined from their group number in the periodic table.	$:\ddot{\underset{\cdot}{Cl}}:$ $Ca:$ $:\ddot{\underset{\cdot}{Cl}}:$
Calcium must lose its two valence electrons (to be left with an octet in its previous principal level), while chlorine only needs to gain one electron to get an octet. Draw two chloride anions, each with an octet and a 1− charge, and one calcium cation with a 2+ charge. Draw the chloride anions in brackets and indicate the charges on each ion.	$[:\ddot{\underset{\cdot\cdot}{Cl}}:]^-$ Ca^{2+} $[:\ddot{\underset{\cdot\cdot}{Cl}}:]^-$
Finally, write the formula with subscripts that indicate the number of atoms.	$CaCl_2$

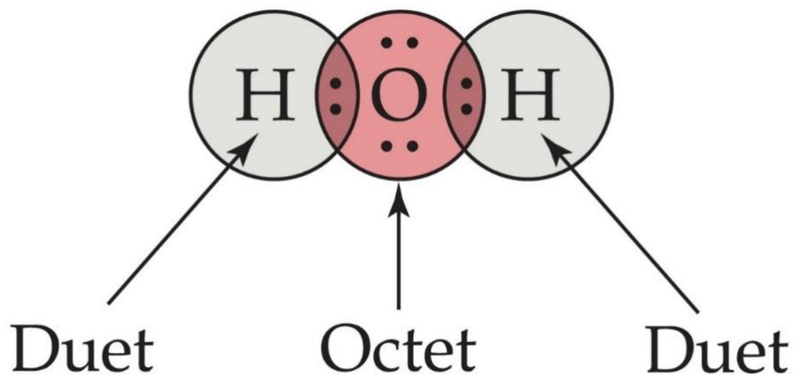
FOR PRACTICE 3.11

Use the Lewis model to predict the formula for the compound that forms between magnesium and nitrogen.

3.13 Lewis Structures: For Covalent Bonding

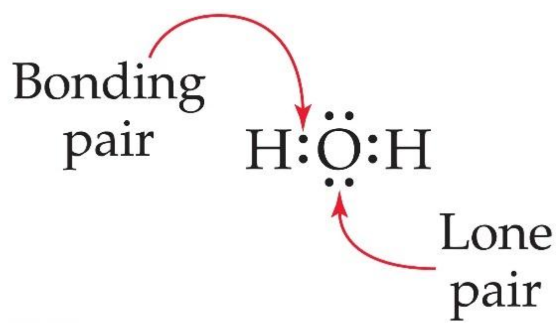
- **Hydrogen and oxygen** have the following Lewis structures:

- In **water**, hydrogen and oxygen <u>share</u> their valence electrons so that each hydrogen atom gets a <u>duet</u> and the oxygen atom gets an <u>octet</u>.

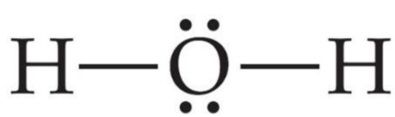


3.13 Lewis Structures: Electron Pairs

- **Bonding Pairs of Electrons** (shared pairs): electrons that are shared between two atoms.
  - ✓ example: 2 bonding pairs in water molecule
- **Lone pairs of Electrons** (unshared pairs): electrons that were not shared (i.e. present only on one atom).
  - ✓ example: 2 lone pairs in water molecule.

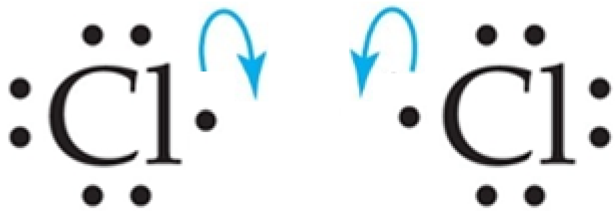


Remember that each dash represents a pair of shared electrons:



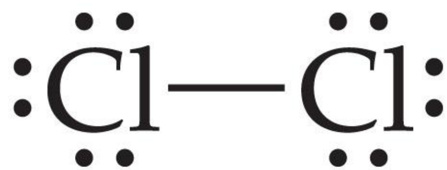
Lewis Theory Explains Why **Halogens form Diatomic Molecules**

- Consider the **Lewis structure of chlorine atoms**:



- If two Cl atoms pair together, they can each attain an **octet**:

or



- Lewis theory explains why chlorine exists as diatomic molecule (Cl_2), as most atoms tend to follow the **octet rule**.
- The same is true for the other halogens (F_2, Br_2, I_2).

Lewis Theory Predicts That Hydrogen Should Exist as H_2

- The individual **Hydrogen atoms** has the following Lewis structure:

$$H\cdot \quad \cdot H$$

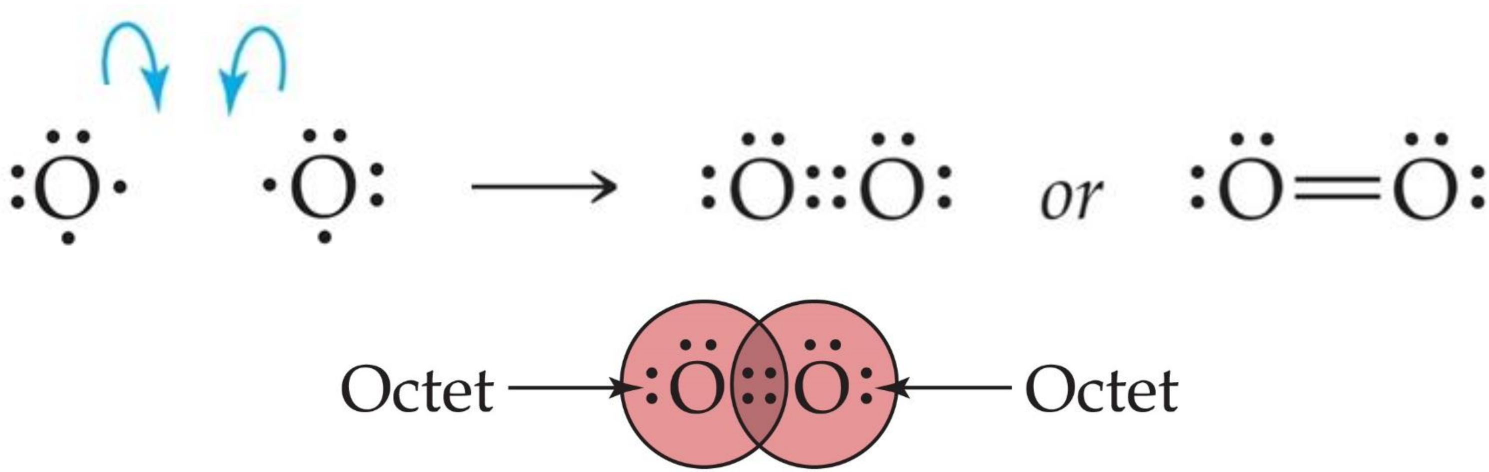
- When two hydrogen atoms share their valence electrons, they each get a **duet**, a stable configuration for hydrogen.

$$H{:}H \quad \textit{or} \quad H{-}H$$

- Lewis theory predicts that elemental hydrogen exists as a diatomic molecule (H_2).

3.13 Lewis Structures: Double and Triple Covalent Bonds

- **Oxygen** exists as a diatomic molecule (**O_2**):



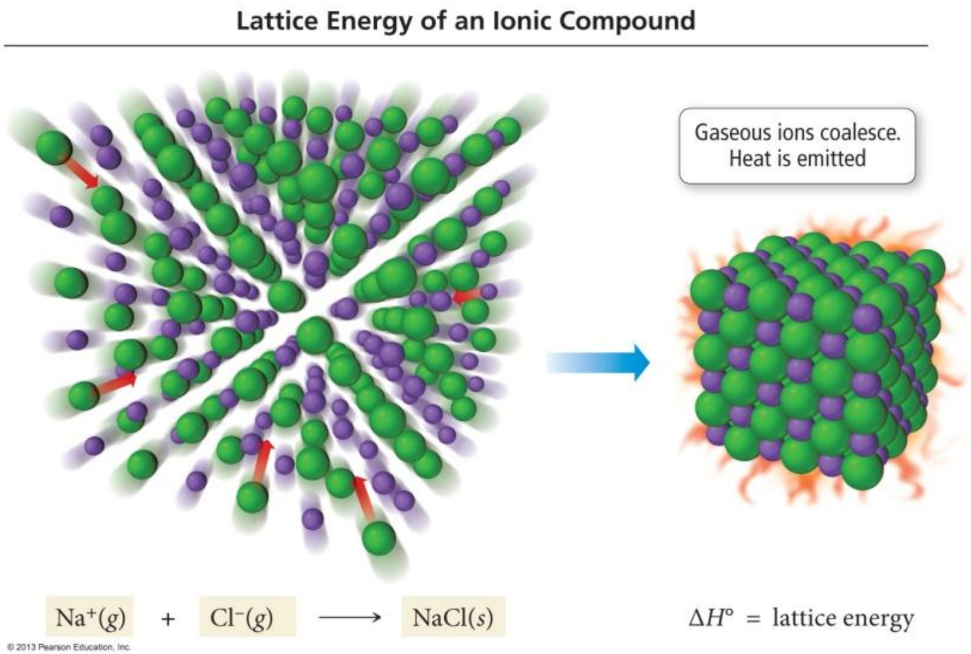
- **Nitrogen** exists as a diatomic molecule (**N_2**):

Lattice Energy (for Crystalline Ionic Compounds)

Lattice Energy: The energy required to completely separate a mole of a crystalline ionic compound into its gaseous ions.

✓ The ions are arranged in a pattern called a crystal lattice.

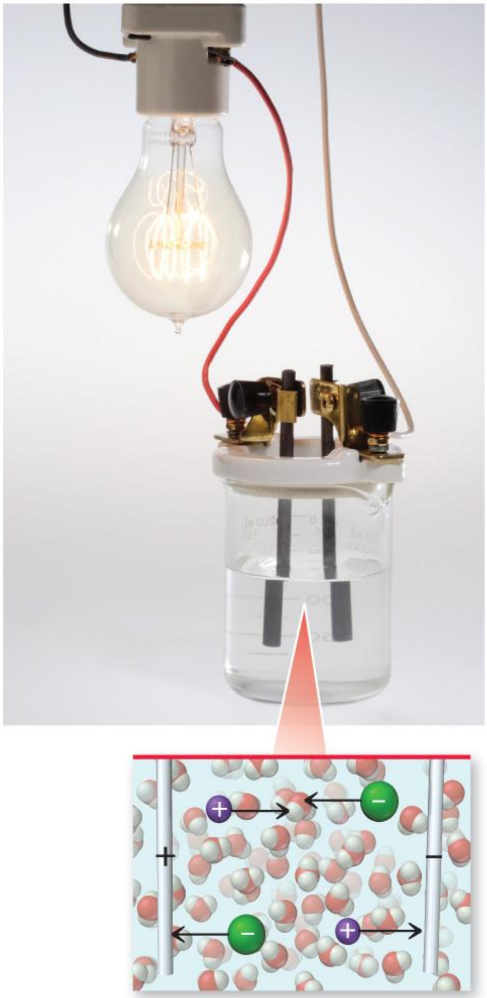
✓ Lattice energy is the most important factor in determining the stability of an ionic compound.



- **Lattice energy increases with:**
  - increasing charge on the ions
  - decreasing size of ions

Physical Properties of Ionic Compounds

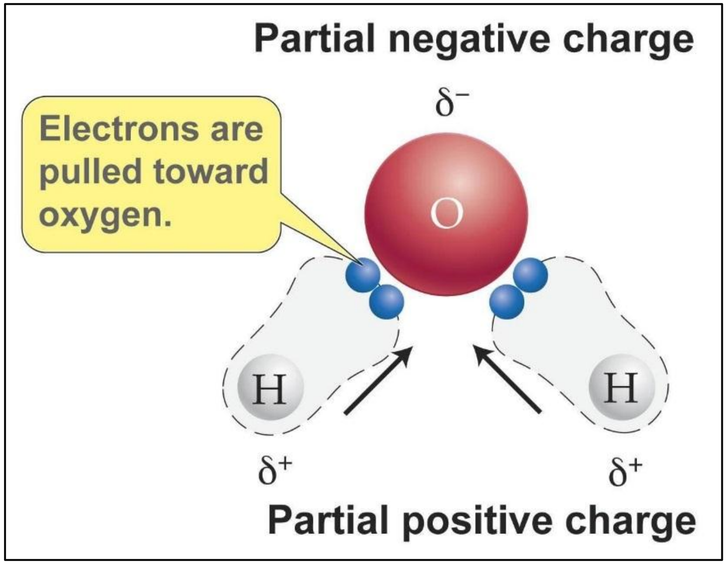
- ✓ High melting points (>300 °C)
- ✓ High boiling points
- ✓ Hard and brittle solids
- ✓ All are crystalline solids at room temperature.
- ✓ Their solids do not conduct electricity, but when in their liquid state they do, i.e. act as strong electrolytes (known as: *molten salts*).
- ✓ Solid and liquid states are thermal insulators.
- ✓ Many ionic compounds are soluble in water.



NaCl(*aq*)

3.14 Electronegativity and Bond Polarity

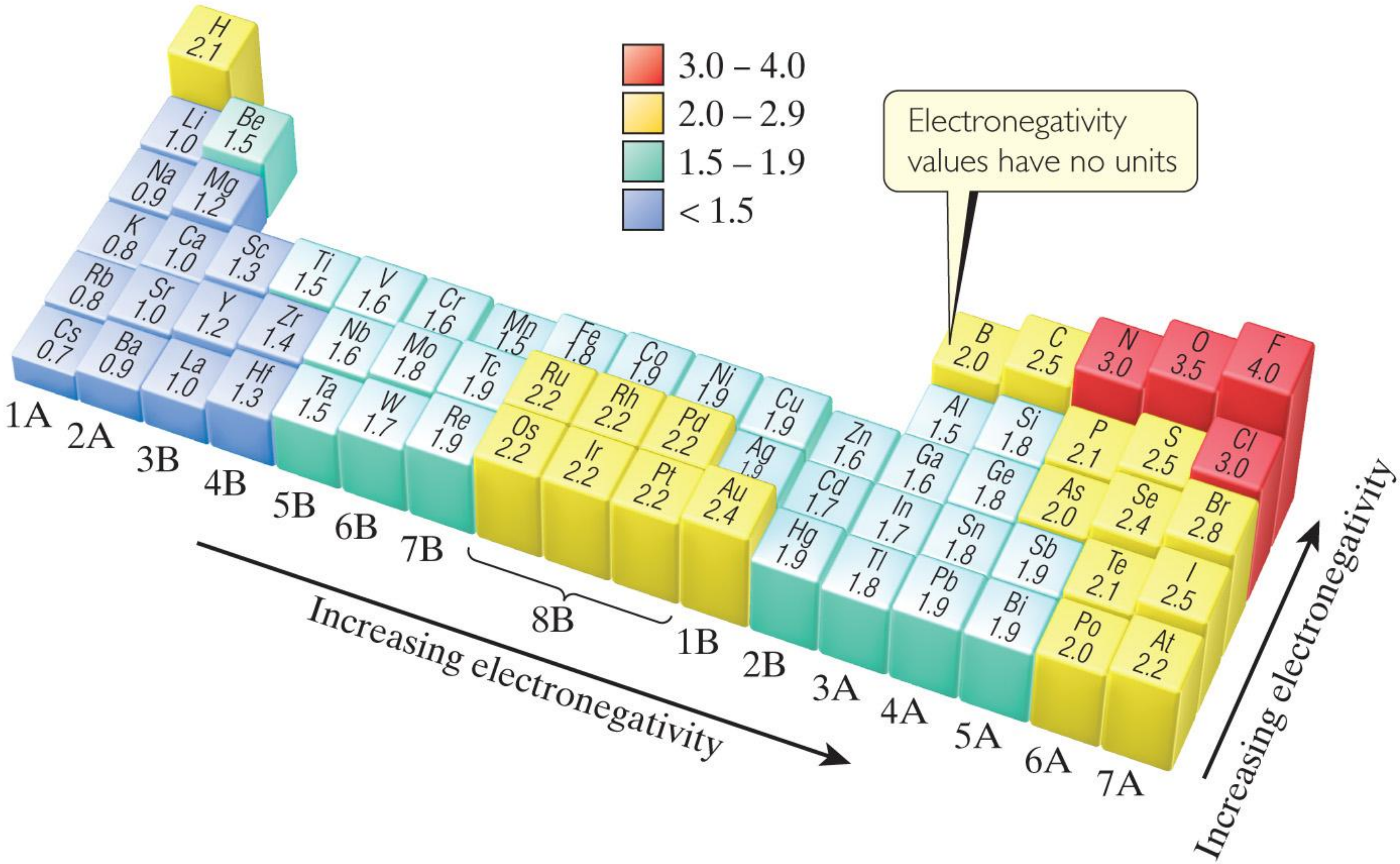
- **Electronegativity (EN):** is the ability of an atom (*in a molecule*) to attract the bond electrons to itself.
  - ✓ is higher for nonmetals; and lower for metals
  - ✓ **The greater the difference in electronegativity (ΔEN), the more polar the bond.**



$\mathrm{H{-}F}$ *or* $\overset{\delta^+}{\mathrm{H}}{-}\overset{\delta^-}{\mathrm{F}}$

$\boldsymbol{\delta^+}$ and $\boldsymbol{\delta^-}$ in polar molecular compounds represent the partial positive and negative charges, to differentiate them from the full charge (+ or -) on ions in ionic compounds.

3.14 Electronegativity Values for Elements (Unitless)



Electronegativity and Bond Types

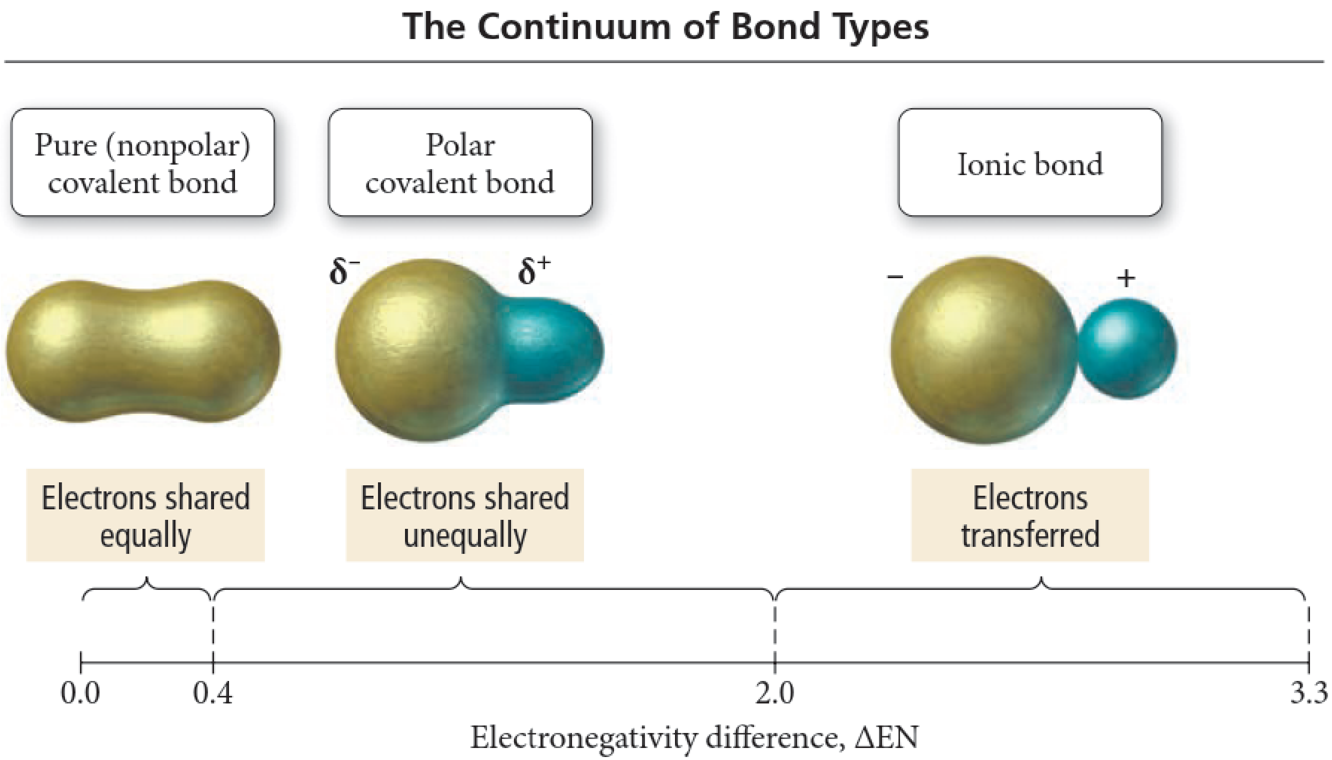


TABLE 3.6 The Effect of Electronegativity Difference on Bond Type

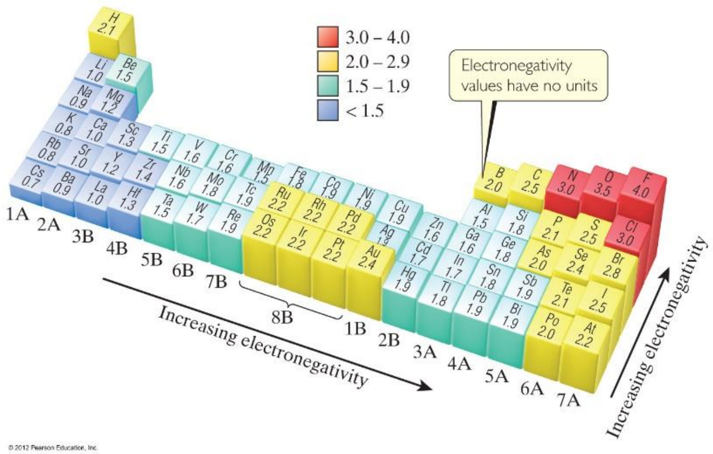
Electronegativity Difference (ΔEN)	Bond Type	Example
Small (0–0.4)	Covalent	Cl_2
Intermediate (0.4–2.0)	Polar covalent	HCl
Large (2.0+)	Ionic	NaCl

- **Example:** Based on the of values of electronegativity (EN) of elements, which bond is more polar: (B-Cl) or (C-Cl)?

Answer:

- The ΔEN of **Cl** and **B** = 3.0 - 2.0 = **1.0**
- The ΔEN of **Cl** and **C** = 3.0 - 2.5 = **0.5**

✓ Hence, the B-Cl bond is more polar.



- **Exercise 1:** Which of the following bonds is the most polar?

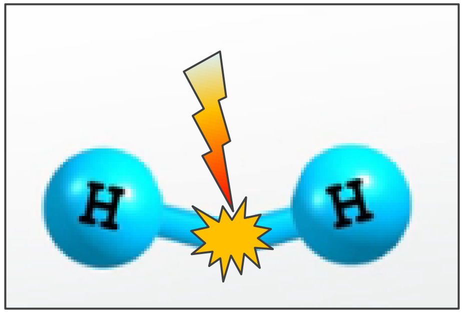
(a) H-F (b) Se-F (c) N-P (d) Ga-Cl

- **Exercise 2:** Predict the type of each bond (use the table of EN values):

(a) H-Br (b) O-O (c) H-O (d) S-O

4.14 Bond Energy and Bond Length

- **Bond Energy:** the amount of energy needed to break one mole of a bond in a compound (in the gaseous state).



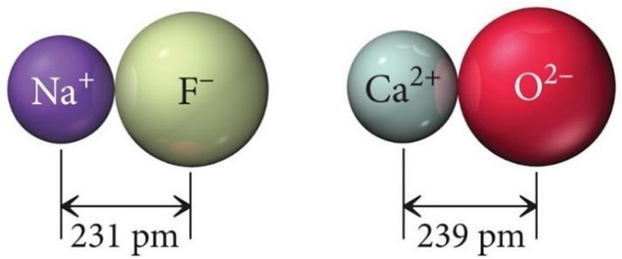
- **Trends in Bond Energies:**
- In general, the more electrons two atoms share, the stronger the covalent bond.

C≡C (837 kJ) > C═C (611 kJ) > C—C (347 kJ)
C≡N (891 kJ) > C═N (615 kJ) > C—N (305 kJ)

  – Bonds get weaker down the column.

  – Bonds get stronger across the period.

- **Bond Length:** the distance between the nuclei of bonded atoms.



- **Trends in Bond Lengths:**
- In general, the more electrons two atoms share, the shorter the covalent bond.

C≡C (120 pm) < C═C (134 pm) < C—C (154 pm)
C≡N (116 pm) < C═N (128 pm) < C—N (147 pm)

- In general, as bonds get longer, they also get weaker.

Conclusion: triple bond (≡) is short and strong, while single bond (—) is long and weak.

1- Write the Lewis structure for each atom or ion:

a. Al b. sodium ion c. magnesium ion d. chloride ion

2- Use Lewis structures to explain why each element occurs as diatomic molecules:

a. hydrogen b. bromine c. oxygen d. nitrogen

3- Write the Lewis structure for each compound:

a. PH_3 b. SCl_2 c. HI d. CH_4

e. NaF f. CaO g. $SrBr_2$ h. K_2O

4- Determine whether a bond between each pair of atoms would be nonpolar covalent, polar covalent, or ionic.

a. Br & Br b. C & Cl c. Mg & I d. Sr & O

5- Order these compounds in order of increasing carbon–carbon **bond strength** and in order of decreasing carbon–carbon **bond length**:

$$HC{\equiv}CH \;,\; H_2C{=}CH_2 \;,\; H_3C{-}CH_3$$

▲ Periodic Table of the Elements

Key

Atomic number → 6; Symbol → C; Name → Carbon; atomic mass → 12.01

An element

Main group | Transition metals | Main group

Period number	1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
Group number	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H Hydrogen 1.008																	2 He Helium 4.003
2	3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
3	11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
4	19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.41	31 Ga Gallium 69.72	32 Ge Germanium 72.64	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
5	37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9	46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3
6	55 Cs Cesium 132.9	56 Ba Barium 137.3	57 La Lanthanum 138.9	72 Hf Hafnium 178.5	73 Ta Tantalum 180.9	74 W Tungsten 183.8	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.1	79 Au Gold 197.0	80 Hg Mercury 200.6	81 Tl Thallium 204.4	82 Pb Lead 207.2	83 Bi Bismuth 209.0	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
7	87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (267)	105 Db Dubnium (268)	106 Sg Seaborgium (271)	107 Bh Bohrium (272)	108 Hs Hassium (270)	109 Mt Meitnerium (276)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (280)	112 Cn Copernicium (285)	113 Nh Nihonium (284)	114 Fl Flerovium (289)	115 Mc Moscovium (288)	116 Lv Livermorium (293)	117 Ts Tennessine (293)	118 Og Oganesson (294)

Lanthanides 6	58 Ce Cerium 140.1	59 Pr Praseodymium 140.9	60 Nd Neodymium 144.2	61 Pm Promethium (145)	62 Sm Samarium 150.4	63 Eu Europium 152.0	64 Gd Gadolinium 157.3	65 Tb Terbium 158.9	66 Dy Dysprosium 162.5	67 Ho Holmium 164.9	68 Er Erbium 167.3	69 Tm Thulium 168.9	70 Yb Ytterbium 173.0	71 Lu Lutetium 175.0
Actinides 7	90 Th Thorium 232.0	91 Pa Protactinium 231.0	92 U Uranium 238.0	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (262)