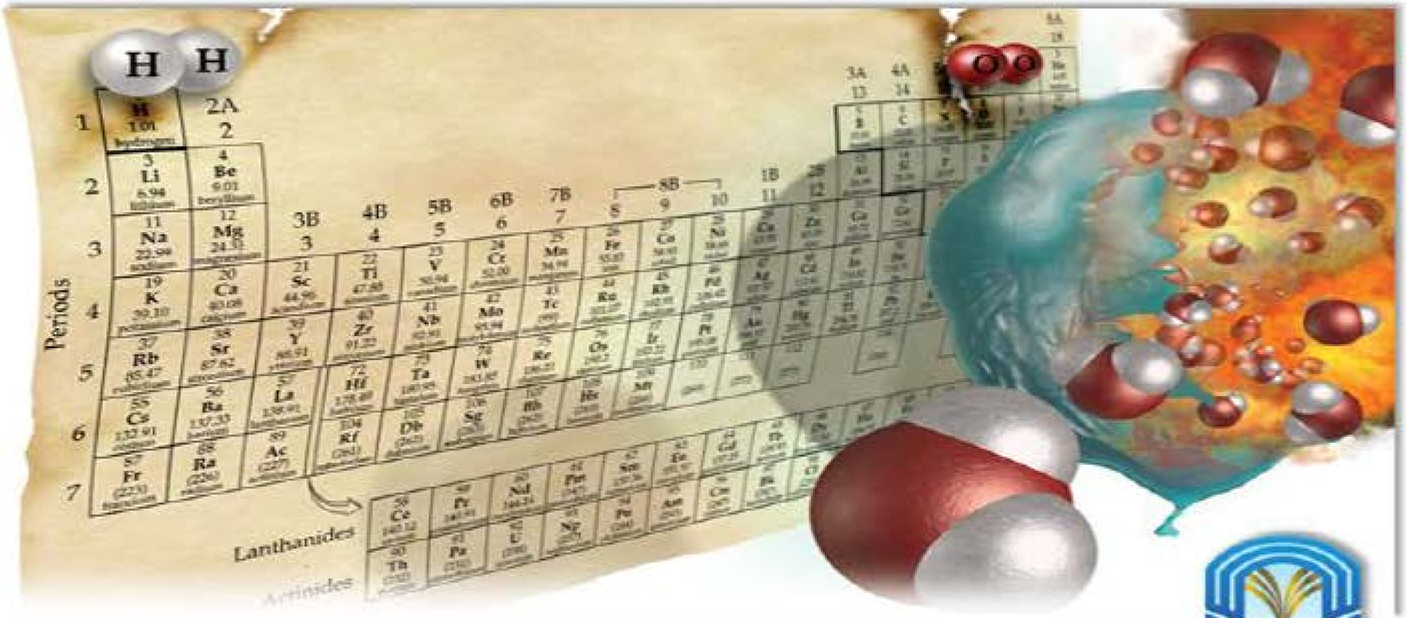
INTRODUCTION TO CHEMISTRY CHEM 101



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Lecture Presentation

Chapter 3

Molecules, Compounds, and Chemical Equations

Topic 08

Chemical Formulas and Molecular Models

Chapter 3

Stoichiometry: calculations with chemical formulas and equations

- Representing compounds.
- An atomic level view of elements and compounds
- Ionic compounds
- Molecular compounds
- Writing and balancing chemical equation.

3.2. Representing compounds:

A compound is represented with its **chemical formula.**

Chemical formula indicates the elements present in the compound and the relative number of atoms or ions of each.

Example:

- 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.

3.2. Representing compounds:

Types of Chemical Formulas

1- Molecular formula (الصيغة الجزيئية):

give the **actual** number of atoms of each element in compound.

e.g. hydrogen peroxide
H_2O_2

2- Empirical formula (الصيغة الأولية):

give the **relative** number of atoms of each element in compound.

e.g. hydrogen peroxide

HO

3- Structural formal (الصيغة البنائية):

use **lines** to represent covalent bonds.

e.g. hydrogen peroxide

H – O – O – H

3.2. Representing compounds:

Example:

1- What the Empirical Formula of H_2O, $C_6H_{12}O_6$, O_3 and N_2H_4?

- The empirical formula of $\mathbf{H_2O}$ is $\mathbf{H_2O}$
- The empirical formula of $\mathbf{C_6H_{12}O_6}$ is $\mathbf{CH_2O}$ (divide the numbers 6, 12 and 6 by the smallest common devisor 6)
- The empirical formula of $\mathbf{N_2H_4}$ is $\mathbf{NH_2}$

2- Write the empirical formulas for the following molecules:

1. Acetylene (C_2H_2)
   CH
2. Nitrous oxide (N_2O)
   $\mathbf{N_2O}$
3. Caffeine $C_8H_{10}N_4O_2$
   $\mathbf{C_{٤}H_{٥}N_{٢}O}$

Divide all the numbers by the smallest common devisor.

3.2. Representing compounds:

Example :

Write the empirical formulas for the following molecules:

(a) C_4H_8,

the greatest common factor is 4 → The empirical formula is therefore CH_2.

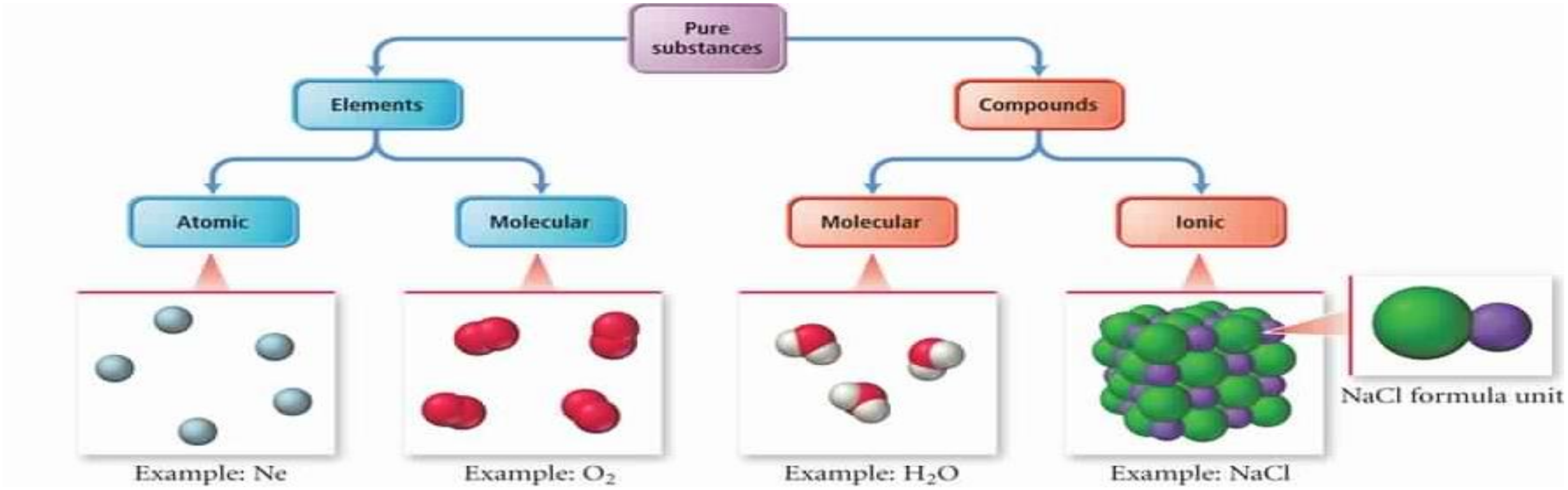
(b) B_2H_6

the greatest common factor is 2 → The empirical formula is therefore BH_3.

(c) CCl_4

the only common factor is 1, → The empirical formula and the molecular formula are **identical**.

3.3. An atomic level view of elements and compounds



3.3. An atomic level view of elements and compounds

- **Atomic element:** Are elements that exist in nature with single atoms as their unite.

- **Molecular element:** Are elements that exist as molecule two or more atoms bonded together.

  1- diatomic molecule: H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2.

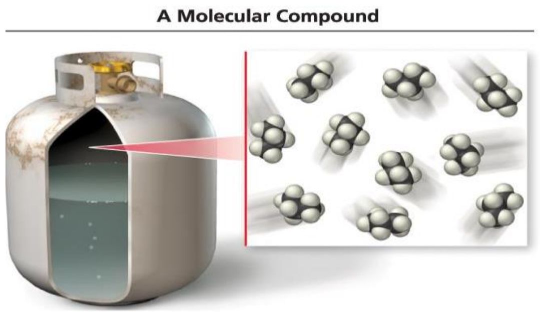
  2- polyatomic molecule ***e.g.*** P_4 and S_8

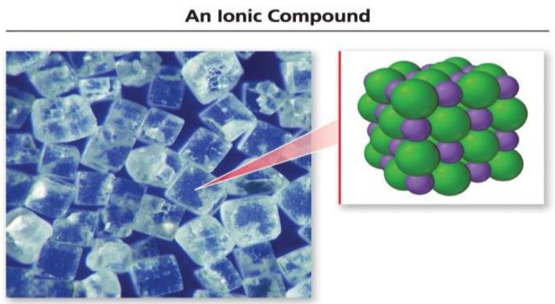
- **Ionic compound:** Are composeds of **cation** (metal) and **anion** (nonmetal) bound together by ionic bond.

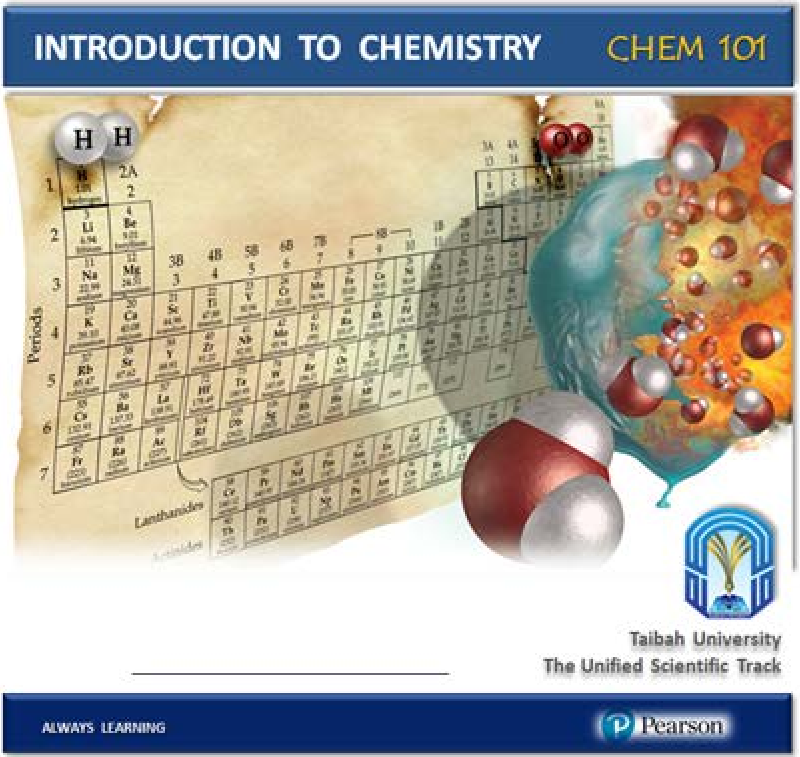
  1- Binary ionic compounds ***e.g.*** NaCl

  2- polyatomic ions compound ***e.g.*** $NaNO_3$

- **Molecular compound:** Are composed of two or more atoms covalently bonded **nonmetal**.







Lecture Presentation

Chapter 3

Molecules, Compounds, and Chemical Equations

Topic 09

Ionic and Molecular Compounds: Formulas and Names

3.4. Ionic Compounds: formal and Naming

Ionic Compounds

Cation Anion

- consist of a combination of cations and an anions, e.g. : Na^+ and Cl^- → NaCl
- The sum of the charges on the cation/s and anion/s in each formula unit **must equal zero**.
- The formula of the ionic compound **must be neutral**.

1- Binary compounds are compounds formed from just two elements.

2- polyatomic ions: are ionic compounds contain ions that are themselves composed of a group of covalently bonded atoms with an overall charge.

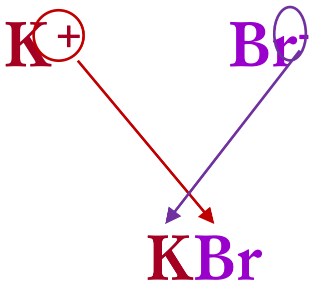
Example:

- $NaNO_3$ contains Na^+ and NO^{3-}.
- $CaCO_3$ contains Ca^{2+} and $CO3^{2-}$.
- KClO Contains K^+ and ClO^-

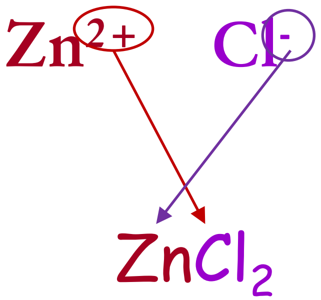
3.4. Ionic Compounds: formal and Naming

How to make the neutral formula?

Potassium Bromide



Zinc Chloride



قاعدة تبادل التكافؤات

The subscript of the cation is numerically **equal** to the charge on the anion, and the subscript on the anion is numerically **equal** to the charge on the cation.

3.4. Ionic Compounds: formal and Naming

How to make the neutral formula?

$2 \times (+3) = +6$ $\quad$ $3 \times (-2) = -6$

Al_2O_3

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

$1 \times (+2) = +2$ $\quad$ $2 \times (-1) = -2$

$CaBr_2$

Ca^{2+} $\quad$ Br^{-}

$1 \times (+2) = +2$ $\quad$ $1 \times (-2) = -2$

Na_2CO_3

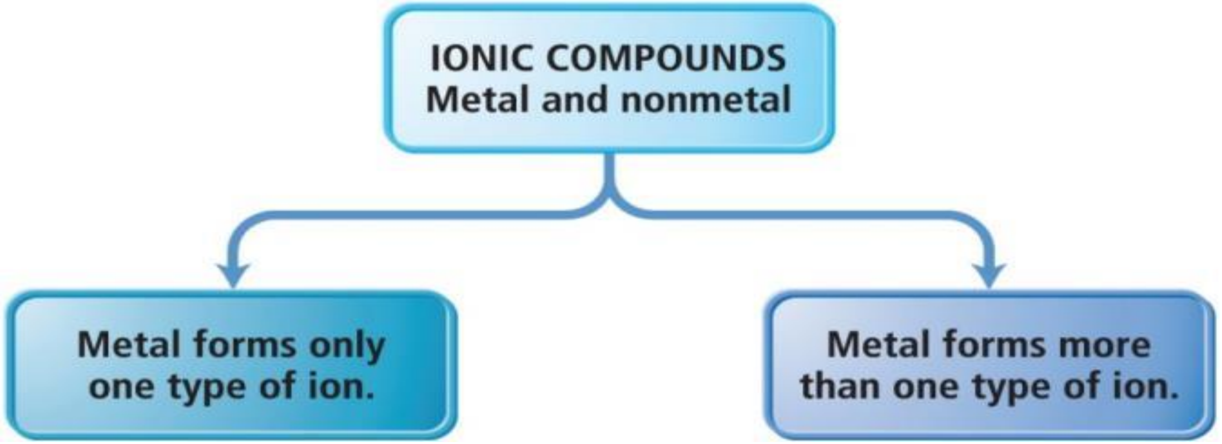
Na^{+} $\quad$ CO_3^{2-}

قاعدة تبادل التكافؤات

The subscript of the cation is numerically **equal** to the charge on the anion, and the subscript on the anion is numerically **equal** to the charge on the cation.

3.4. Ionic Compounds: formal and Naming

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



The ion always has the same charge.

The ion has the different charge.

3.4. Ionic Compounds: formal and Naming

1.1 Naming Binary Ionic Compounds of Type I Cations

The names of binary ionic compounds take the following form:

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

Example:

1- 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 → potassium chlor**ide**.

2- 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 → calcium ox**ide**.

3.4. Ionic Compounds: formal and Naming

1.1 Naming Binary Ionic Compounds of Type I Cations

Table 3.2.Base Names of some common 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

3.4. Ionic Compounds: formal and Naming

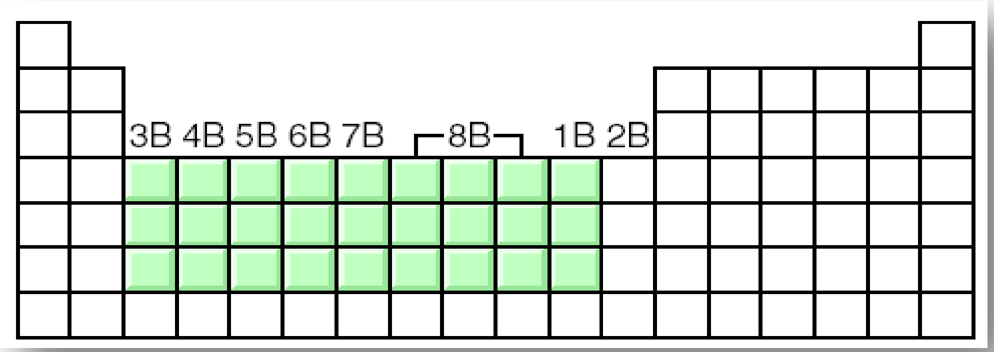
1.2 Naming Binary Ionic Compounds of Type II Cations

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

- Metals of this type are often transition metals.
- Some main group metals, such as Pb, Tl, and Sn, form more than one type of cation.

Example:

- FeO → iron is +2 cation (Fe^{2+}).
- Fe_2O_3 → iron is +3 cation (Fe^{3+}).
- Cu_2O → copper is +1 cation (Cu^{+}).
- CuO → copper is +2 cation (Cu^{2+}).



3.4. Ionic Compounds: formal and Naming

1.2 Naming Binary Ionic Compounds of Type II Cations

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

we distinguish between Fe^{2+} and Fe^{3+} as follows:

Fe^{2+} Iron(II)

Fe^{3+} Iron(III)

Roman numerals	
I	One
II	two
III	Three
IIII	Four
IV	Five

Example:

$FeCl_2$ → Iron (II) chloride

$FeCl_3$ → Iron (III) chloride

Cr_2S_3 → chromium(III) sulfide

3.4. Ionic Compounds: formal and Naming

1.2 Naming Binary Ionic Compounds of Type II Cations

Table 3.33.4 Some Metals That Form Cations with Different Charges

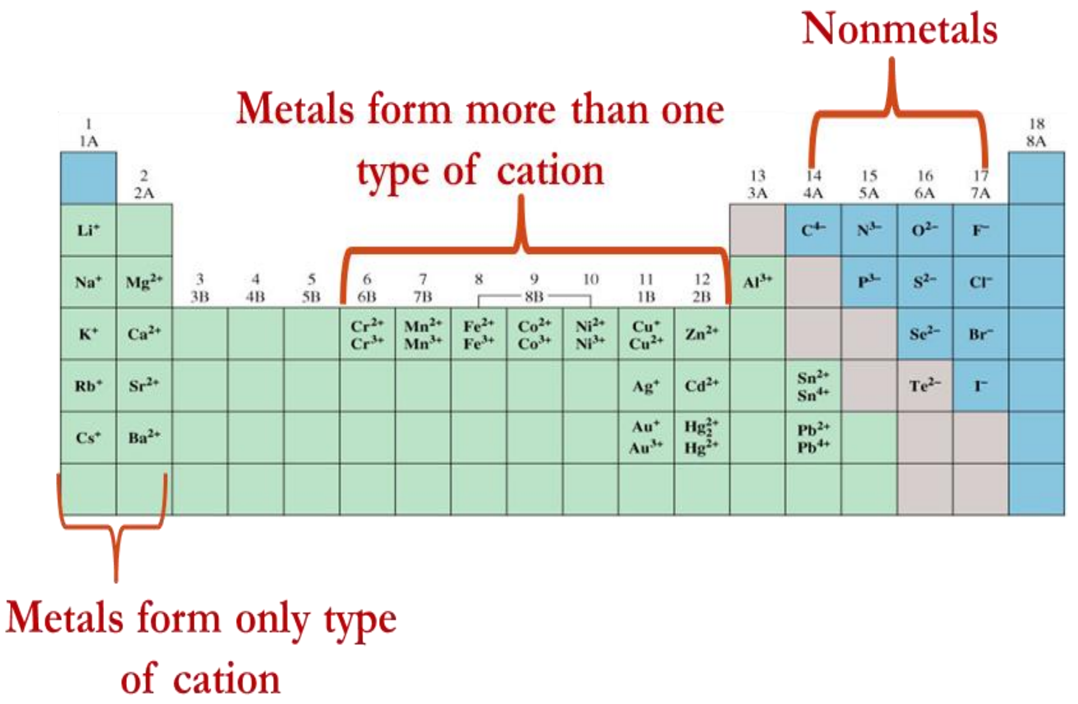
Metal	Ion	Name	Older Name*
Chromium	Cr^{2+}	Chromium(II)	Chromous
	Cr^{3+}	Chromium(III)	Chromic
Iron	Fe^{2+}	Iron(II)	Ferrous
	Fe^{3+}	Iron(III)	Ferric
Cobalt	Co^{2+}	Cobalt(II)	Cobaltous
	Co^{3+}	Cobalt(III)	Cobaltic
Copper	Cu^{+}	Copper(I)	Cuprous
	Cu^{2+}	Copper(II)	Cupric
Tin	Sn^{2+}	Tin(II)	Stannous
	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

3.4. Ionic Compounds: formal and Naming

Ionic Binary compounds

1- Metals form **only one** type of cation:
- Alkali metals.
- Alkali earth metals.
- Ag^{+}, Al^{3+}, Cd^{2+}, Zn^{2+}

2- Metals form **more than one** type of cation:
- Transition Metals.
- other metal cations.



3.4. Ionic Compounds: formal and Naming

2. Naming Ionic Compounds Containing Polyatomic Ions

Ionic compounds that contain a polyatomic ion is name in the same way as other ionic compounds, **except** that we use the name of the polyatomic ion whenever it occurs.

Example:

$NaNO_2$ is named according to

its cation, Na^+ → sodium

its polyatomic anion, NO_2^- → nitrite

$NaNO_2$ → sodium nitrite

3.4. Ionic Compounds: formal and Naming

2. Naming Ionic Compounds Containing Polyatomic Ions

Most polyatomic ions are **oxyanions anions** containing oxygen and another element.

1-If there are **two** ions in the series,

- The one with **more** O_2 atoms has the ending -**ate**,
  e.g. NO^{3-} → nitrate
- The one with **fewer** has the ending **-ite**.
  e.g. NO^{2-} → nitrite

2- If there are more than two ions in the series

- the prefixes **per-**, meaning **more** than,
- the prefixes **hypo-**, meaning **less** than,

→

ClO^- → hypochlorite
ClO_2^- → chlorite
ClO_3^- → chlorate
ClO_4^- → perchlorate

3.4. Ionic Compounds: formal and Naming

2. Naming Ionic Compounds Containing Polyatomic Ions

Table 3.45 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^+		

3.4. Ionic Compounds: formal and Naming

2. Naming Ionic Compounds Containing Polyatomic Ions

Example:

Name the following compounds:

(a) $Cu(NO_3)_2$

1. Cation: Copper cation (can form two types of cation) → Copper (II)

2. Anion: NO_3^- anion has a common name Nitrate

Thus: the name of the compound is: **Copper (II) nitrate**

(b) KH_2PO_4

1. Cation: K form only one type of cation→ Potassium Note: not potassium (I)

2. Anion: H_2PO^{4-} has a common name dihydrogen phosphate

Thus: the name of the compound is: **Potasium dihydrogen phosphate**

3.4. Ionic Compounds: formal and Naming

Example :

Determine the name for the compound $Li_2Cr_2O_7$?

4. Classify each compound as ionic or molecular.

a. CO_2
b. $NiCl_2$
c. NaI
d. PCl_3

A-molecular لان كل مكوناته لافلزات	B-ionic لانه تكون من فلز +للافلز
C-ionic	D-molecular

12. Write a formula for each ionic compound.

a. sodium hydrogen sulfite
b. lithium permanganate
c. silver nitrate
d. potassium sulfate
e. rubidium hydrogen sulfate
f. potassium hydrogen carbonate

A-NaHSO3	B-LiMnO4
C-AgNO3	D-K2SO4
E-RbHSO4	F-KHSO4

3.5. Molecular Compounds: Formulas and Names

- **Molecular compounds** are composed of two or more **nonmetals**.
- The formula for a molecular compound **cannot** readily be determined from its constituent elements (**why?**)
- because the same combination of elements may form many different molecular compounds, each with a different formula.

Example: Nitrogen and oxygen form all of the following unique molecular compounds:

NO NO_2 N_2O N_2O_3 N_2O_4 N_2O_5

3.5. Molecular Compounds: Formulas and Names

How to write the name of Molecular Compounds?

1- write the name of the element with the **smallest group** number first.

e.g. HCl → **hydrogen chloride**

- If the two elements lie in the same group, then write the element with the **greatest row** number first.

e.g. SO_2 → **sulphur dioxide**

2- The prefixes given to each element indicate the number of atoms present.

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

Greek Prefixes Used in Naming Molecular Compounds

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

3.5. Molecular Compounds: Formulas and Names

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112	(113)	114	(115)	116	(117)	118
				58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

3.5. Molecular Compounds: Formulas and Names

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

Example:

HCl	Hydrogen chloride
SiC	Silicone carbide
CO	Carbon monoxide
CO_2	Carbon dioxide
SO_2	Sulphur dioxide
SO_3	Sulphur trioxide
NO_2	Nitrogen dioxide
N_2O_4	Dinitrogen tetroxide

Mono- is omitted if its in the first element.

e.g. Monocarbon monoxide

→ Carbon monoxide

monooxide → monoxide the o is omitted

tetraoxide → tetroxide the a is omitted

3.5. Molecular Compounds: Formulas and Names

Example:

Name the following molecular compounds:

decaoxide the a is omitted

(a) $SiCl_4$ One silicon atom & four chlorine atoms

Thus: **Silicon tetrachloride**

(b) P_4O_{10} Four phosphourus atoms & ten oxygen atoms

Thus: **Tetraphosphourus decoxide**

3.5. Molecular Compounds: Formulas and Names

Example:

Write chemical formulas for the following molecular compounds:

(a) carbon disulphide

There is one carbon atom & two sulphur

Atoms, Thus: $\mathbf{CS_2}$

(b) Disilicon hexabromide

There is two silicon atoms & six bromine

Atoms, Thus: $\mathbf{Si_2Br_6}$

3.5. Molecular Compounds: Formulas and Names

Acids: are molecular compounds that form H^+ when dissolved in water.

- To indicate the compound is dissolved in water (*aq*) is written after the formula.

e.g. HCl → dissolved in water, → $H^+_{(aq)}$ and $Cl^-_{(aq)}$ ions

A compound is not considered an acid if it does not dissolve in water.

Characters of acids:

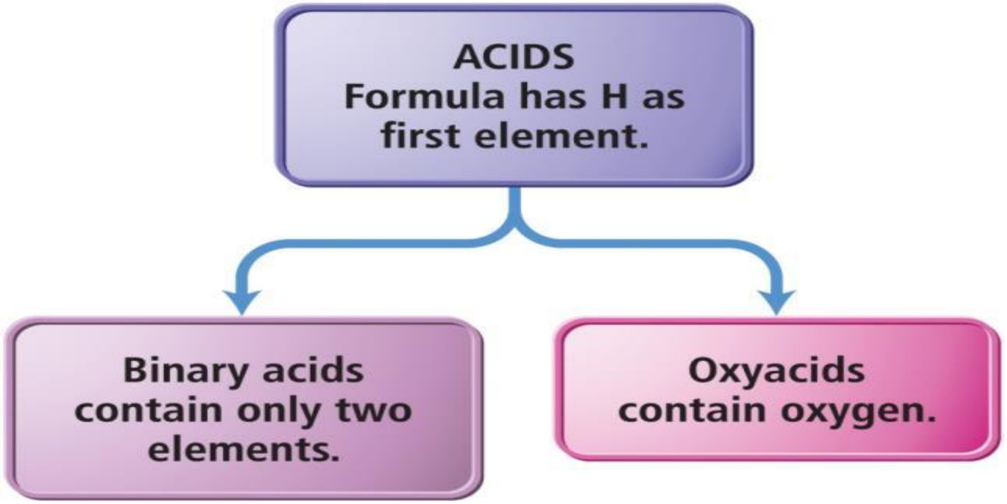
1. Sour taste
2. Dissolve many metals

   such as Zn, Fe, Mg; but not Au, Ag, Pt
3. Formula generally starts with H

   e.g., HCl, H_2SO_4

3.5. Molecular Compounds: Formulas and Names

Acids can be categorized into **two types**



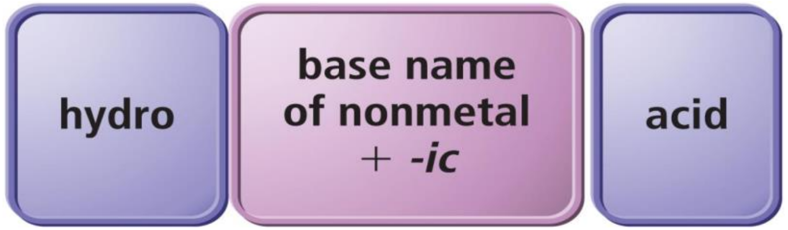
have H^{+1} cation and nonmetal anion.

have H^{+} cation and polyatomic anion.

3.5. Molecular Compounds: Formulas and Names

How to Naming Binary Acids?

1. Write a **hydro-** prefix.
2. Follow with the nonmetal name.
3. Change ending on nonmetal name to **–ic**.
4. Write the word **acid** at the end of the name.



Example: HCl → Hydrochloric acid
HBr → Hydrobromic acid

3.5. Molecular Compounds: Formulas and Names

Example

1- Determine the name for the compound HI(aq) ?

2-Determine the name for the compound HF(aq) ?

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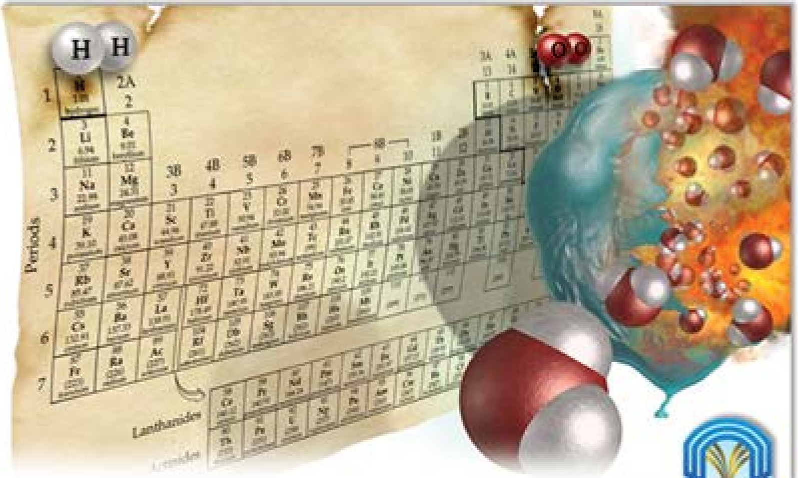

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Chapter 3

Molecules, Compounds, and Chemical Equations

Topic 10

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

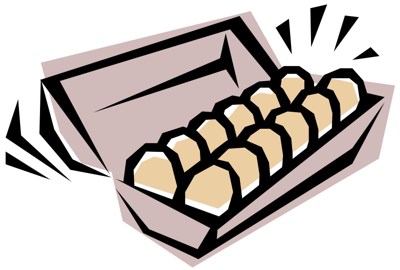
** Molar mass: Counting Atoms by Weighing Them

- **Mole (mol)** is the SI unit of the amount of substance.

- **Avogadro's Number (N_A):**
- one mole of each substance contains **6.022 x 10^{23} particles**
  - (atoms or molecules or ions etc.).

Thus 6.022 x 10^{23} **is the Avogadro's Number (N_A)**

$$1 \text{ mol} = N_A = 6.0221367 \times 10^{23}$$



=Dozen = 12

1 dozen = 12 Anything

1 mol = 6.022 x10^{23} particles

- Chemistry is quantitative in nature

- **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 (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**

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

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**

- **Molar mass (M):** the mass (in g or kg) of one mole of a substance;

M = mass/mol = g/mol

For **1 MOLE: 1 amu = 1 g**

e.g. The atomic mass of ^{12}C is 12.00 amu = 12.00 g

1 mole of ^{12}C = 12.00 amu = 12.00 g = has 6.022 x 10^{23} (N_A)atoms

- Thus:

The Molar Mass (M) of ^{12}C = 12.00 g/mol

Molar Mass (g/mol)= Atomic Mass (amu)

Examples:

1. The atomic mass of Na = 22.99 amu

The molar mass of Na = 22.99 g/mol

Element	(Formula mass)	(M_m) Molar Mass
Na	23 amu	23 g/mole
H	1 amu	1 g/mole from atoms
H_2	2 amu	2 g/mole from molecules
Fe	56 amu	56 g/mole
Al	27 amu	27 g/mole

Compound	Formula mass	M_m. Molar Mass
H_2O	1 × 2 + 16×1 = 18 amu	18 g/mole
NaOH	23×1 + 16×1 + 1×1 =40 amu	40 g/mole
CH_3OH	12×1 + 1 × 3 + 16×1×1 + 1 =32 amu	32 g/mole
H_2CO_3	1 × 2 + 12×1 + 16 × 3 = 62 amu	62 g/mole

(Ex-)-What is the molar mass of Na_2SO_4?

(a) 134 g/mole (b) 148 g/mole

(c) 158 g/mole (d) 142 g/mole

(Ex-)-What is the molar mass of nicotine, $C_{10}H_{14}N_2$?

(a) 134 g/mole (b) 148 g/mole

(c) 158 g/mole (d) 162 g/mole

(Ex-)-What is the molar mass of H_2SO_4?

(a) 134 g/mole

(b) 148 g/mole

(c) 158 g/mole

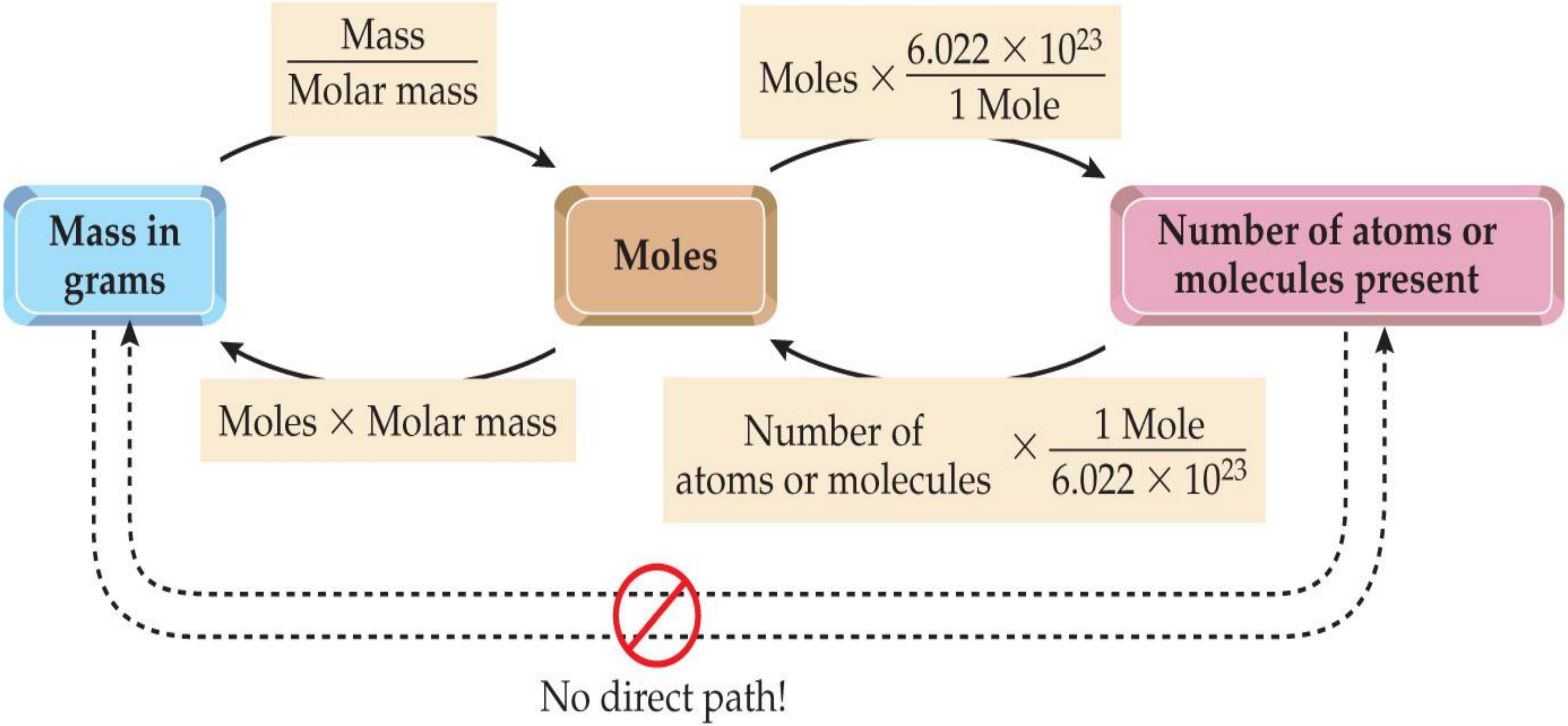
(d) 98 g/mole

(Ex:) -Penicillin is $C_{16}H_{18}N_2O_4S$. What is the molecular weight of penicillin?

✓Exercises: Calculate the molar mass (g/mole) 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}**
- **$Al_2(CO_3)_3$**

Conversions in Stoichiometry Calculations



n = number of moles

m = mass (atom or molecule)

M = molar mass (atomic mass or molecular mass)

What is the relation between them?

$$n = \frac{m}{M} = \frac{g}{g/mol} = mol$$

n = number of moles

N = number of atoms or molecules

N_A = Avogadro's number

(atoms or molecules/mol)

What is the relation between them?

$$n = \frac{N}{N_A} = \frac{\text{atoms (or molecules)}}{\text{atoms (or molecules) / mol}} = \text{mol}$$

Example:

How many moles of He atoms are in 6.46 g of He?

Solution:

- Number of moles = n
- **The atomic mass of He = The Molar mass of He**
- **From the periodic Table:**

The atomic mass of He = The molar mass of He = 4.003 g/mol

Thus: 4.003 g → 1 mole of He

6.46 g → ? moles of He

Thus: there is **1.61 moles of He atoms in 6.46 g of He**

How many the amount (in moles) of He atoms are in 6.46 g of He?

$$n(He) = \frac{m}{M} = \frac{6.46g}{4.003g/mol} = 1.61mol$$

Example: How many S atoms are in 16.3 g of S?

$$n(S) = \frac{m}{M} = \frac{16.3\,\text{g}}{32.07\ \text{g/mol}} = 0.508\,\text{mol}$$

$$n(S) = \frac{N}{N_A} \Rightarrow$$

$$N = nxN_A$$

$$= 0.508\ \text{mol}\ x\ 6.022x\,10^{23}\ \text{atoms/mol}$$

$$= 3.06x\,10^{23}\ \text{atoms}$$

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?

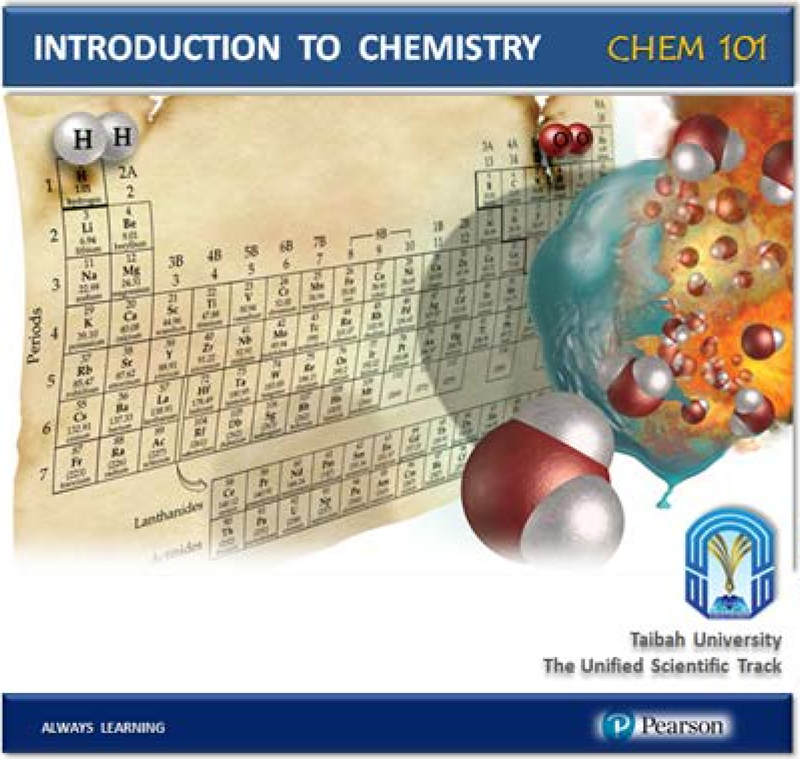
4- How many **moles** of NH_3 are there in **0.2 Kg** of NH_3?

x

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 $Al_2(CO_3)_3$?



Lecture Presentation

Chapter 3

Molecules, Compounds, and Chemical Equations

Topic 11

- Calculating Mass Percent
- Calculating Empirical and Molecular Formulas
- Balancing Chemical Equations

3.6. Composition of Compounds

- **Percent Composition by Mass:** is the percent by mass of each element in a compound

- percent composition of an element in a compound = $\frac{\text{mass of element X in 1 mol of compound}}{\text{mass of 1 mol of the compound}} \times 100$

أو بتعبير أخر

- percent composition of an element in a compound = $\frac{n \text{ x molar mass of element}}{\text{molar mass of compound}} \text{ x } 100\%$

Where n is the number of moles of the element in 1 mole of the compound

3.6. Composition of Compounds

- **Example:**

Calculate the mass percent of carbon in ethane (C_2H_6)?

$$\text{Mass percent of C in } C_2H_6 = \frac{(2)\text{x}(\text{molar mass of C})}{\text{molar moss of } C_2H_6} \times 100$$

$$\text{Molar moss of } C_2H_6 = (2 \ \text{x} \ (\text{molar mass of C})) + (6 \ \text{x molar mass of H})$$

$$= (2 \ \text{x} 12.011 \text{g/mol}) + (6 \ \text{x} \ 1.007 \text{g/mol}) = 30.070 \ \text{g/mol}$$

$$\%C = \frac{2 \ \text{x} \ (12.011 \text{g/mol})}{30.070 \ \text{g/mol}} \times 100$$

$$= 79.887\%$$

3.6. Composition of Compounds

- **Example:**

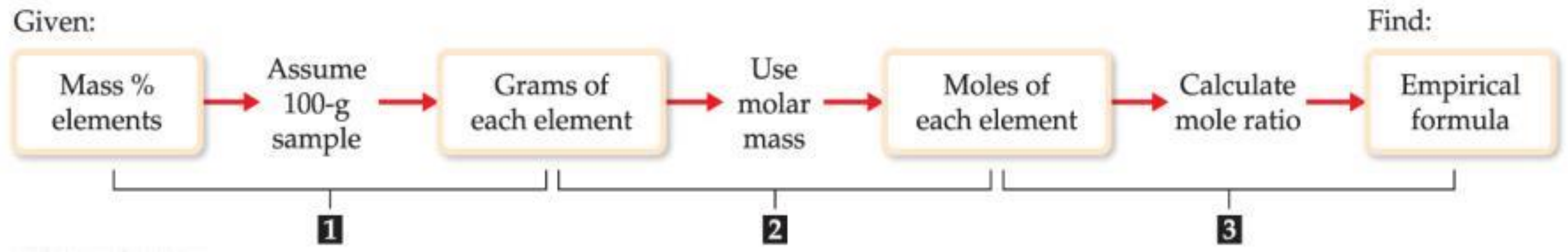
Calculate the mass percent of Cl in CCl_2F_2?

3.7. Determining a Chemical Formula from Experimental Data

Can be determined chemical formal from elemental analysis (experimental معمليا) by:

1. Masses of elements formed
2. Or Percent composition

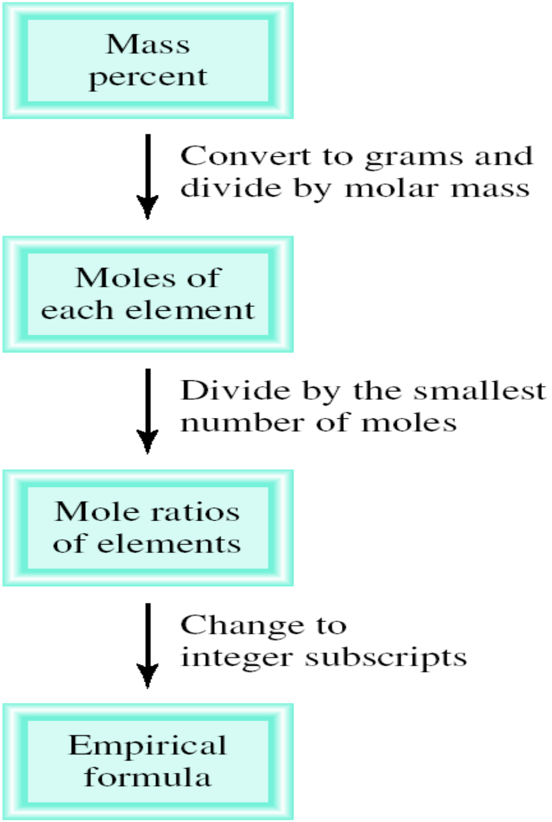
How ?



3.7. Determining a Chemical Formula from Experimental Data

How ?

1. Convert the percentages to grams.
   a) Assume you start with 100 g of the compound.
   b) Skip if already grams.
2. Convert grams to moles.
   a) Use molar mass of each element.
3. Write a pseudo formula using moles as subscripts.
4. Divide all by 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 0.5, multiply all by 2.
   b) if ratio 0.33 or 0.66, multiply all by 3.
   c) If ratio 0.25 or 0.75, multiply all by 4.
   d) If ratio 0.20 or 0.40 or 0.80 multiply all by 5
   d) Skip if already whole numbers.



3.7. Determining a Chemical Formula from Experimental Data

Q: Determine the empirical formula of a compound that has the following percent composition by mass:
K 24.75, Mn 34.77, O 40.51 percent.

	K	Mn	O
% →100g	24.75g	34.77g	40.51g
n=m/M	24.75/39.10 =0.633mol	34.77/54.94 =0.6329mol	40.51/16.00 = 2.532mol
÷ on smallest no. of mole	0.633/0.632 =1	0.6329/0.632 = 1	2.532/0.632 =4
The empirical formula is	K_1	Mn_1	O_4
	$KMnO_4$		

خطوات الحل

١. ننشأ جدول نضع فيه العناصر المذكورة في السؤال
٢. نعتبر أن النسبة المئوية معبر عنها بالجرام فلو كان عندنا ١٠٠ جرام من المركب فهذه ال ١٠٠ جرام موزعة على العناصر حسب نسبتها.
٣. نوجد عدد المولات nلكل عنصر باستخدام القانون n=m/M .
٤. نقسم عدد المولات على أصغر مول من العناصر.
٥. الأرقام التي نحصل عليها تمثل empirical formulaبشرط أن تكون أعداد صحيحة كما في المثال السابق.
٦. في حالة ظهور أعداد عشرية نقوم بضرب الأرقام التي في الأسفل الموجودة في الصيغة بأعداد بدأ من ٢، ٣ .... حتى نحصل على أعداد صحيحة .

3.7. Determining a Chemical Formula from Experimental Data

Q/ Ascorbic acid composed of 40.92% C, 4.58% H, and 54.50% O by mass. Determine its empirical formula.

	C	H	O
% →100g			
n=m/M			
÷ on smallest no. of mole			
Convert into integer x 3			
The empirical formula is			

3.7. Determining a Chemical Formula from Experimental Data

Determining Molecular Formulas For Compounds

- The molecular formula is a 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.

- The molar mass is a whole-number multiple of the **empirical formula molar mass**, the sum of the masses of all the atoms in the empirical formula:

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

3.7. Determining a Chemical Formula from Experimental Data

Determining Molecular Formulas For Compounds

Example:

A sample compound contains 1.52g of N and 3.47g of O. The molar mass of this compound is between 90g/mol and 95g/mol. Determine the molecular formula.

Solution:

1. $$n_N = \frac{1.52}{14.01} = 0.108 \text{ mol of N}$$

2. $$n_O = \frac{3.47}{16.00} = 0.217 \text{ mol of O}$$

3. $$N : \frac{0.108}{0.108} = 1 \qquad O = \frac{0.217}{0.108} = 2$$

Thus the empirical formula is: NO_2

Present Composition by Mass

↓

Empirical Formula

↓

Molecular Formula

3.7. Determining a Chemical Formula from Experimental Data

Determining Molecular Formulas For Compounds

4. The molar mass of the empirical formula NO_2 = 14.01 + (2x16.00) = 46.01g

5. The ratio between the empirical formula and the molecular formula:

$$\text{Ratio} = \frac{\text{molar mass of compound}}{\text{empirical molar mass}}$$

$$\text{Ratio} = \frac{90}{46.01} \approx 2$$

6. The molecular formula is $(NO_2)_2$ = N_2O_4

3.8 Writing and Balancing Chemical Equations

Chemical Reactions

- Reactions involve chemical changes in matter resulting in new substances.
- Reactions involve rearrangement and exchange of atoms to produce new molecules.
- Elements are not transmuted during a reaction.

Reactants → Products

3.8 Writing and Balancing Chemical Equations

Chemical equations: it shorthand way of describing a reaction

1. provide information about the reaction
2. formulas of reactants and products
3. states of reactants and products
4. relative numbers of reactant and product molecules that are required
5. can be used to determine weights of reactants used and products that can be made

TABLE 3.6 States of Reactants and Products in Chemical Equations

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

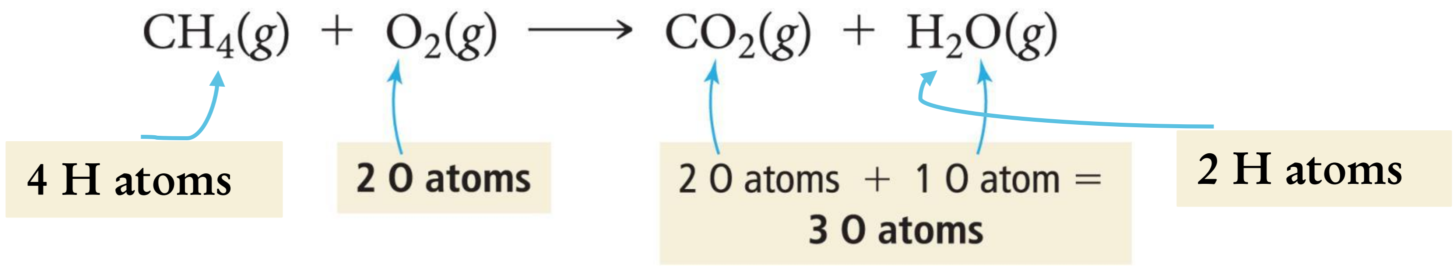
3.8 Writing and Balancing Chemical Equations

Combustion of Methane

- Methane gas ($CH_4(g)$)burns to produce carbon dioxide gas ($CO_2(g)$) and gaseous water ($H_2O(g)$).
  - Whenever something burns it combines with $O_2(g)$.

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

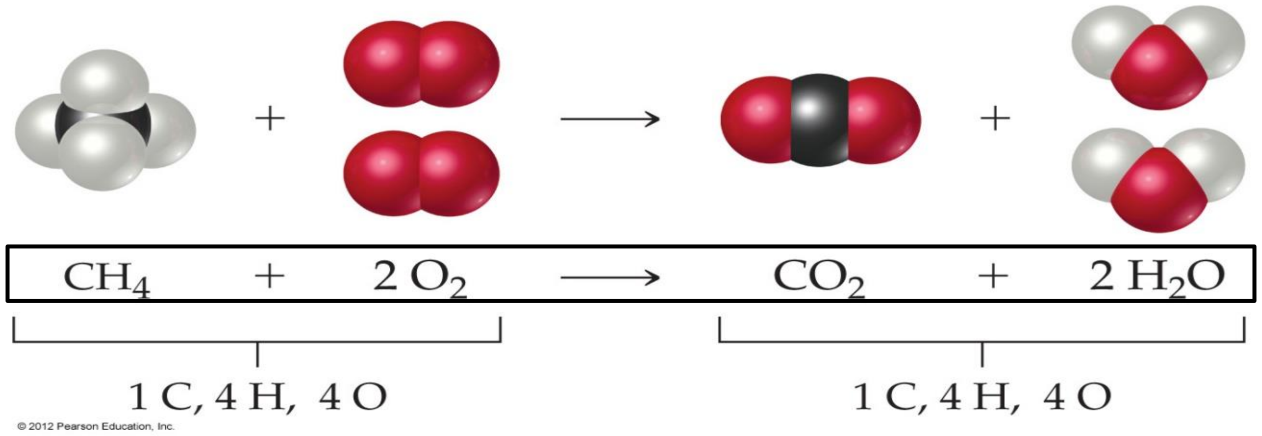
- If you look closely, you should immediately spot a **problem.**



Chapter 3/Stoichiometry

3.8 Writing and Balancing Chemical Equations

Combustion of Methane



Reactants appear on the left side of the equation.

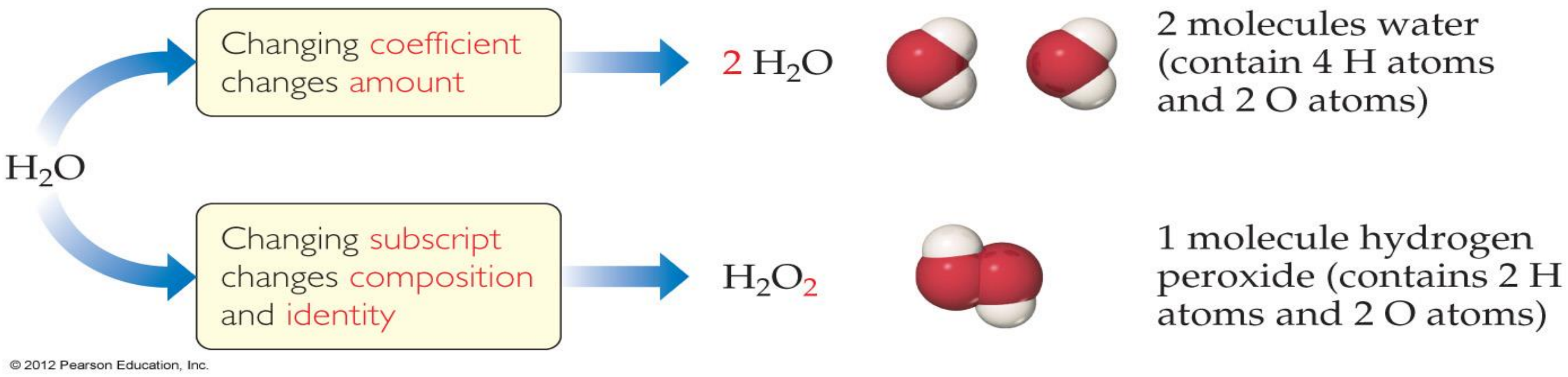
Products appear on the right side of the equation.

The **states** (*g*, *aq*, *s and l*) of the reactants and products are written in parentheses to the right of each compound.

Coefficients are inserted to balance the equation.

3.8 Writing and Balancing Chemical Equations

Subscripts and Coefficients Give Different Information



- Subscripts (الرمز السفلي) tell the number of atoms of each element in a molecule.
- Coefficients (المعاملات) tell the number of molecules.

3.8 Writing and Balancing Chemical Equations

Balancing Chemical Equations

- The principle that lies at the heart of the balancing process is that atoms are conserved in a chemical reaction.
- Atoms are neither created nor destroyed.
- The same number of each type of atom is found among the reactants and among the products.
- The identities (formulas) of the compounds must never be changed in balancing a chemical equation.

3.8 Writing and Balancing Chemical Equations

How to write and balance equations?

1. Read the description of the chemical reaction. What are the reactants, the products, and their states? Write the appropriate formulas.

   Hydrogen gas (H_2) and oxygen gas (O_2) combine to form liquid water (H_2O).

2. Write the **unbalanced** equation that summarizes the information from step 1.

   $H_2(g) + O_2(g) \rightarrow H_2O(l)$

3.8 Writing and Balancing Chemical Equations

How to write and balance equations?

3. Balance the equation by inspection, starting with the most complicated molecule.

$$H_2(g) + O_2(g) \longrightarrow H_2O(l)$$

$$H_2(g) + O_2(g) \longrightarrow H_2O(l) + H_2O(l)$$

2 H   2 O   4 H   2 O

$$H_2(g) + H_2(g) + O_2(g) \longrightarrow H_2O(l) + H_2O(l)$$

4 H   2 O   4 H   2 O

We must balance the equation by adding **more** molecules of reactants and/or products.

3.8 Writing and Balancing Chemical Equations

How to write and balance equations?

4. Check to see that the coefficients used give the **same number** of each type of atom on both sides of the arrow.
5. Also check to see that the coefficients used are the **smallest integers** that give the balanced equation.

The balanced equation is:

$$\mathbf{2}H_2(g) + O_2(g) \rightarrow \mathbf{2}H_2O(l)$$

3.8 Writing and Balancing Chemical Equations

EXAMPLE :

Balance the following equation in standard form (lowest multiple integers) and determine the sum of the coefficients?

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

3.8 Writing and Balancing Chemical Equations

EXAMPLE :

Balance the following equation in standard form (lowest multiple integers) and determine the sum of the coefficients?

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

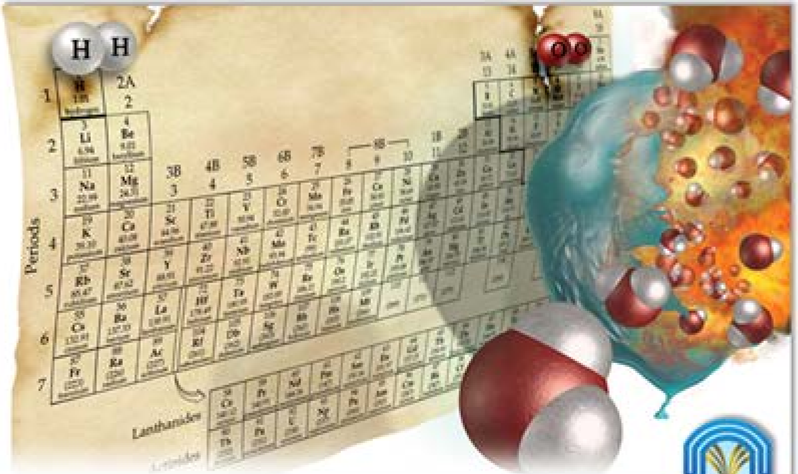
3.8 Writing and Balancing Chemical Equations

EXAMPLE:

Balance the following equation in standard form and determine the sum of the coefficients?

$$FeO(s) + O_2(g) \rightarrow Fe_2O_3(s)$$

INTRODUCTION TO CHEMISTRY CHEM 101



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Lecture Presentation

Chapter 3

Molecules, Compounds, and Chemical Equations

Topic 12

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

Chapter 4: Chemical Bonding and Chemical Reactions

4.9 Types of Chemical Bonds

Chemical bond is the force that holds the atoms together in a molecule,

Types of Chemical Bonds

- ionic bond
- covalent bond
- metallic bonding

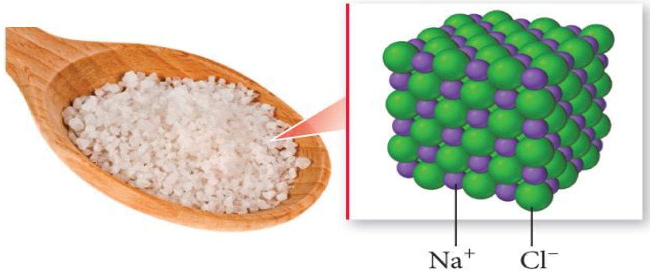
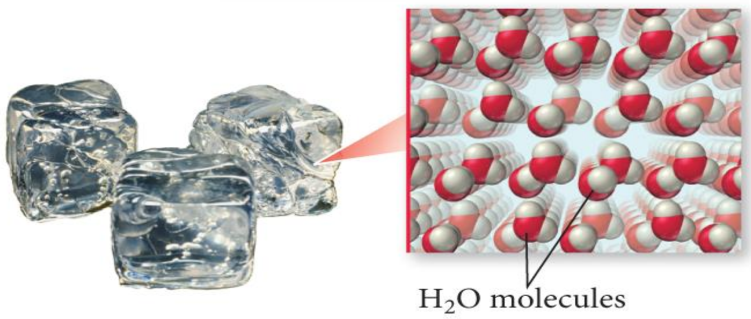
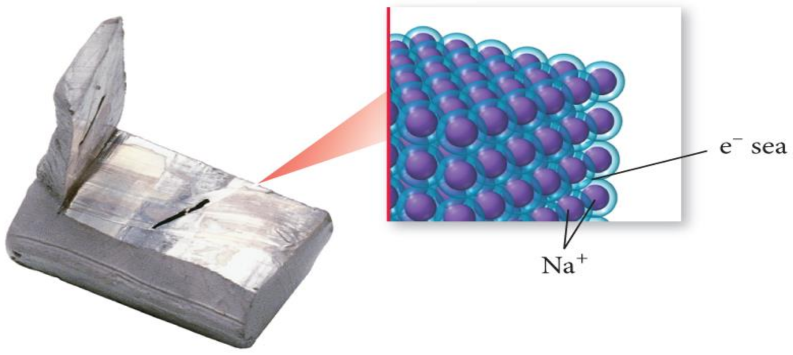


Table salt, NaCl(*s*)



Ice, $H_2O(s)$



Sodium metal, Na(*s*)

4.9 Types of Chemical Bonds

Type of Bond	Types of Atoms	Bond Characteristic
Ionic	Metals + Nonmetals	Electrons transferred
Covalent	Nonmetals + Nonmetals Nonmetals + Metalloid	Electrons shared
Metallic	Metal + Metal	Electron pooled

4.9 Types of Chemical Bonds

1-Ionic bond

When a metal atom loses electrons it becomes a **cation**.

Metals have low ionization energy, making it relatively easy to remove electrons from them.

When a nonmetal atom gains electrons it becomes an anion.

Nonmetals have **high electron affinities**, making it advantageous to add electrons to these atoms.

The oppositely charged ions are then attracted to each other, resulting in **an ionic bond.**

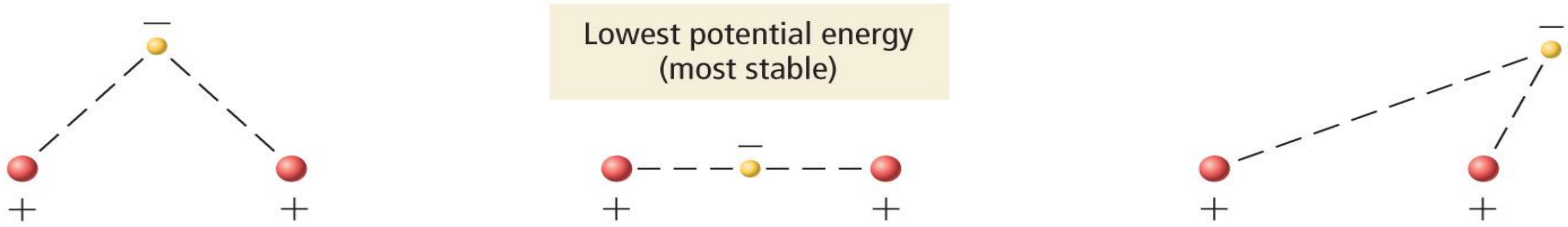
4.9 Types of Chemical Bonds

2- covalent bond:

- Nonmetal atoms have relatively high ionization energies, so it is difficult to remove electrons from them.

- When nonmetals bond together, it is better in terms of potential energy for the atoms to share valence electrons.
  - Potential energy is **lowest** when the electrons are between the nuclei.

- Shared electrons hold the atoms together by attracting nuclei of both atoms.



4.9 Types of Chemical Bonds

3- Metallic bond:

The relatively low ionization energy of metals allows them to lose electrons easily.

The simplest theory of metallic bonding involves the metal atoms releasing their valence electrons to be shared as a pool by all the atoms/ions in the metal.

An organization of metal cation islands in a sea of electrons

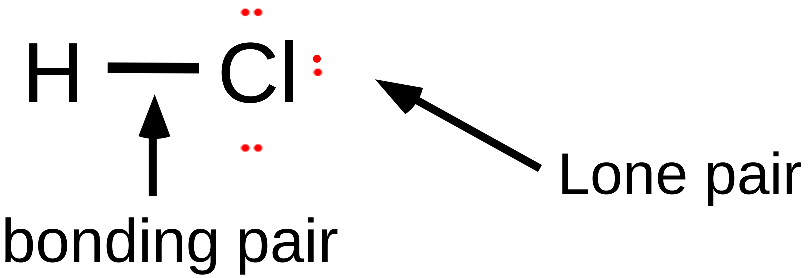
Electrons delocalized throughout the metal structure

Bonding results from attraction of cation for the delocalized electrons.

4.10 Lewis Dot Symbols نظام النقط للويس

Lewis Dot Representation: the valence electrons of an atom are represented by **dots.**

The Octet Rule القاعدة الثمانية : atoms usually gain, lose or share electrons until they have 8 in the outer shell to reach the same electronic configuration of the noble gasses (ns2 np6).



4.10 Lewis Dot Symbols نظام النقط للويس

Exception:

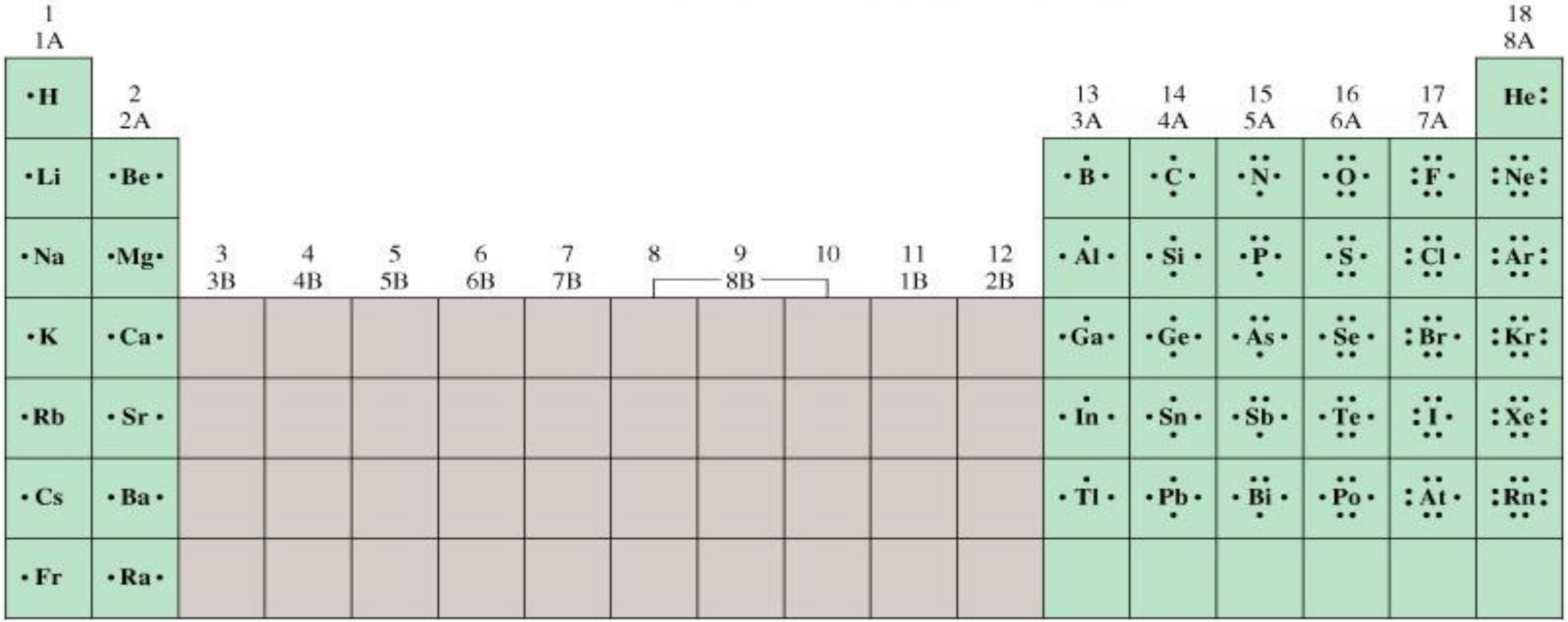
He = two valence electrons, **a duet**.

He :

Lewis structural of He is always paired

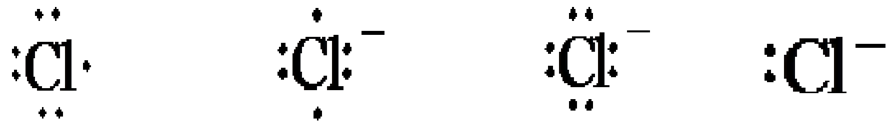
Because n=1 fills with 2 electrons.

4.10 Lewis Dot Symbols نظام النقط للويس

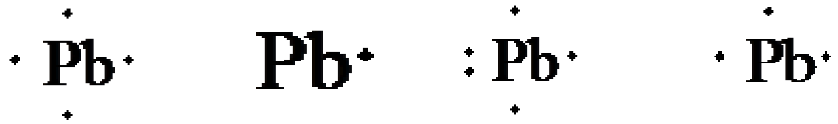


4.10 Lewis Dot Symbols نظام النقط للويس

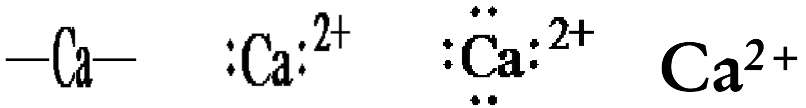
1-The Lewis dot symbol for the chloride ion is



2-The Lewis dot symbol for the lead atom is



3-The Lewis dot symbol for the calcium ion is



4-The Lewis dot symbol for the S^{2-} ion is

write the electron configuration of atom or ion

↓

determine the valence electrons

↓

draw the Lewis dot structure

Ionic bonding:

In Lewis theory, we represent ionic bonding by **moving** electron dots **from the metal to the nonmetal.**

Example, potassium and chlorine have the Lewis structures:

When potassium and chlorine bond, potassium **transfers** its valence electron to chlorine

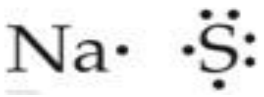
K $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

K^+ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0$

Octet in previous level

Ionic bonding:

Example: The Lewis structures for sodium and sulfur are:

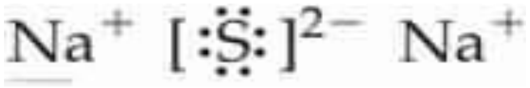


Sodium must **lose one valence electron** to obtain an octet,

while sulfur must **gain two** electrons to obtain an octet.

The compound that forms between Na and S requires two Na atoms to every one S atom.

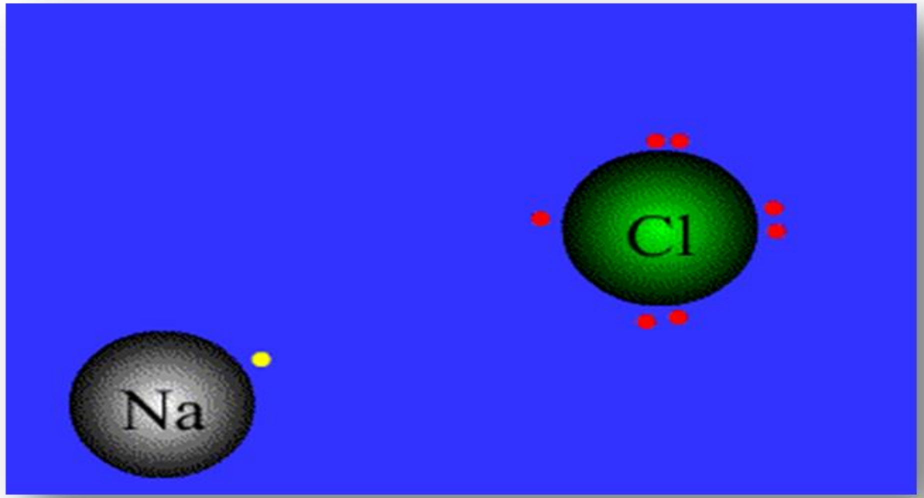
The Lewis structure is:



The correct chemical formula is: Na_2S.

Ionic bonding:

- The ionization energy of the metal is **endothermic**.
  - $Na(s) \rightarrow Na^{+}(g) + 1\ e^{-}$ $\Delta H° = +496$ kJ/mol

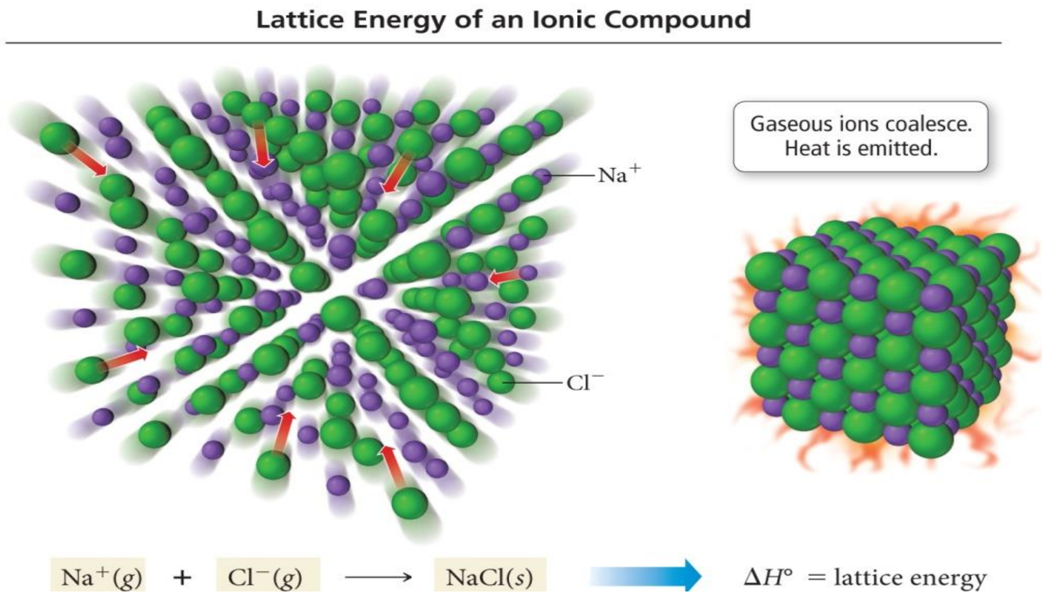


- The electron affinity of the nonmetal is **exothermic**.
  - $\frac{1}{2}Cl_2(g) + 1\ e^{-} \rightarrow Cl^{-}(g)$ $\Delta H° = -244$ kJ/mol

- The heat of formation of most ionic compounds is exothermic and generally large.
  - $Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$ $\Delta H°_f = -411$ kJ/mol

Ionic bonding:

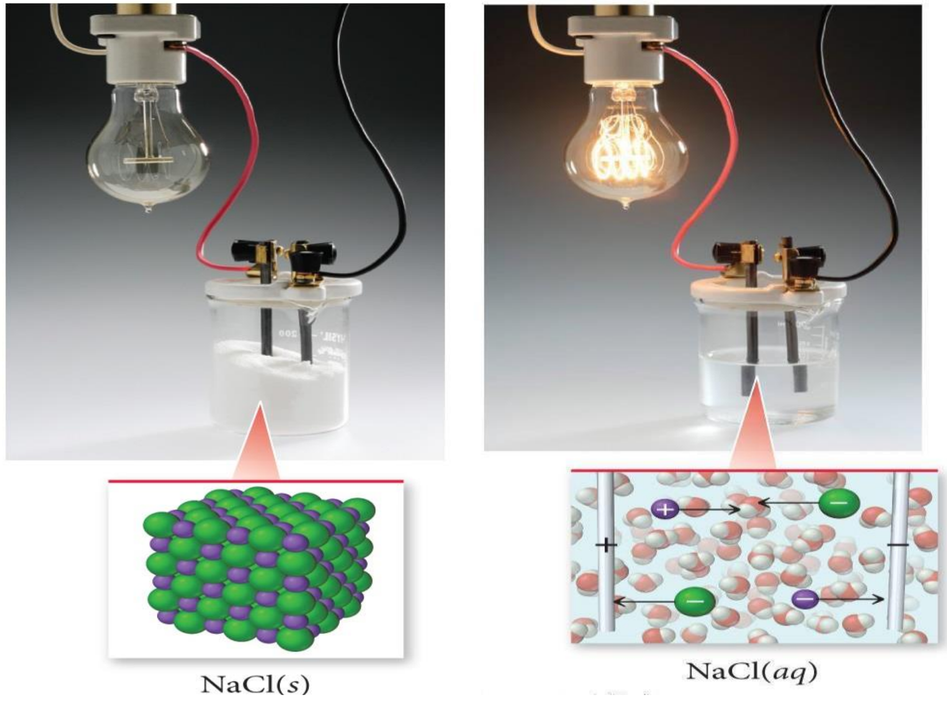
- The extra energy that is released comes from the formation of a structure in which every cation is surrounded by anions, and vice versa.
- This structure is called a **crystal lattice**.
- **The crystal lattice** is held together by the electrostatic attraction of the cations for all the surrounding anions.
- Electrostatic attraction is nondirectional



Ionic bonding:

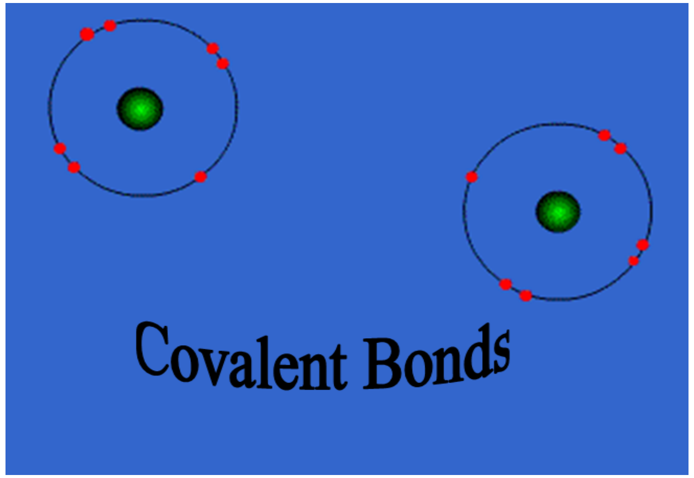
Conductivity of NaCl

In NaCl(*s*), the ions are stuck in position and not allowed to move to the charged rods.



In NaCl(*aq*), the ions are separated and allowed to move to the charged rods.

Covalent Bonding



Valence electrons is to share their valence electrons with other atoms.

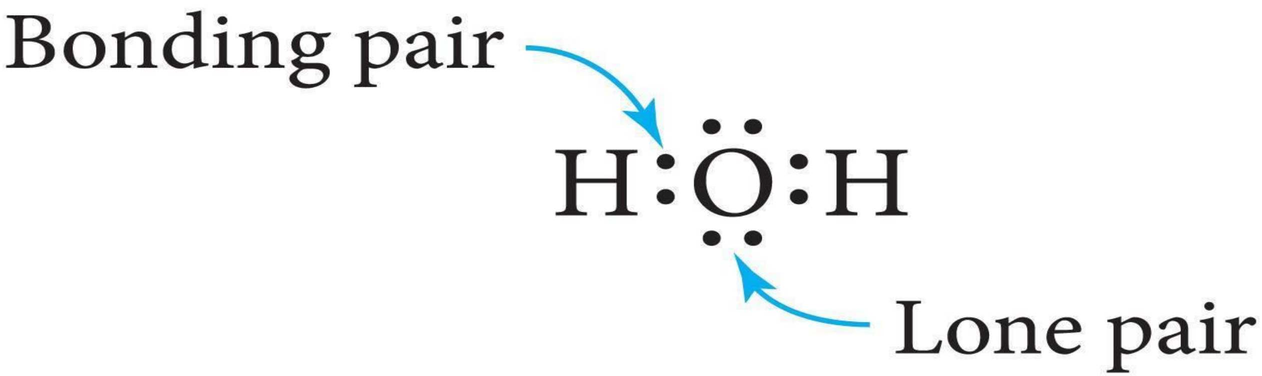
The shared electrons would then count toward each **atom's octet.**

The sharing of valence electrons is called **covalent bonding.**

Electrons that are shared by atoms are called bonding pairs.

Electrons that are not shared by atoms but belong to a particular atom are called lone pairs.

- ✓Also known as **nonbonding pairs**



Covalent Bonding

When two atoms share one pair of electrons, it is called a **single covalent bond**.

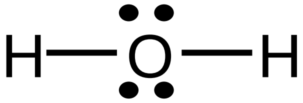
✓Two electrons

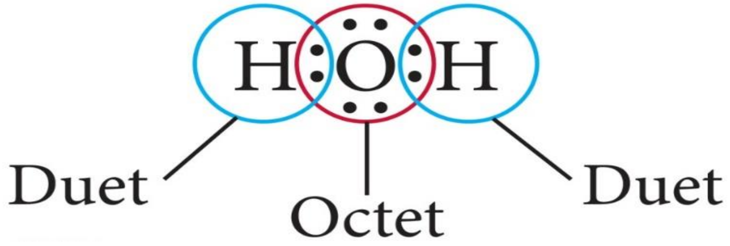
One atom may use more than one single bond to fulfill its octet.

✓To different atoms

✓H only duet

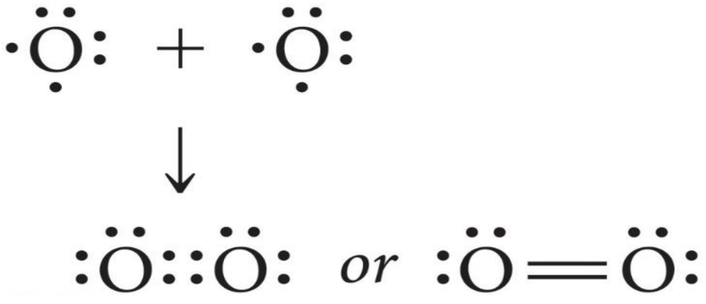
or

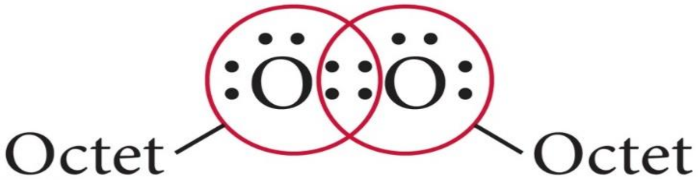




Covalent Bonding

- When two atoms share two pairs of electrons the result is called a **double covalent bond**.
  - Four electrons





- When two atoms share three pairs of electrons the result is called a **triple covalent bond**.
  - Six electrons

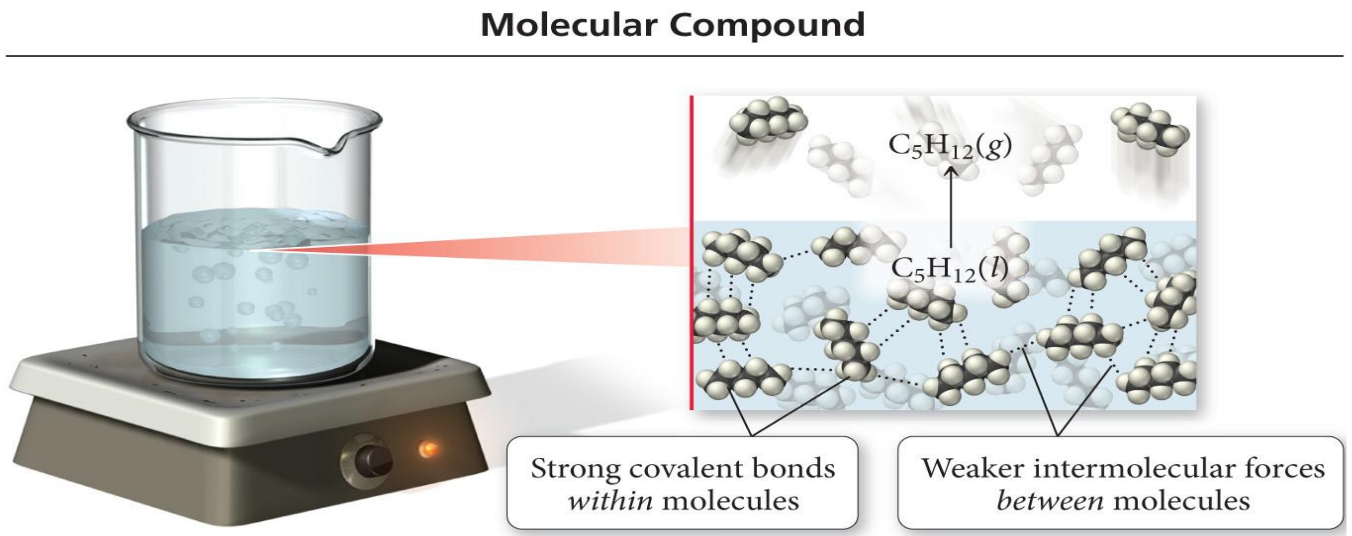
N = 2 x 5ve = 10 ve

Covalent Bonding

- Lewis theory of covalent bonding implies that the attractions between atoms are **directional**.
  - The shared electrons are most stable between the bonding atoms.

- Therefore, Lewis theory predicts covalently bonded compounds will be found as individual molecules.
  - Rather than an array like ionic compounds

- Compounds of nonmetals are made of individual molecule units.



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

a. Al   b. Na^+ c. Mg^{2+} d. Cl^- e. Ne

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

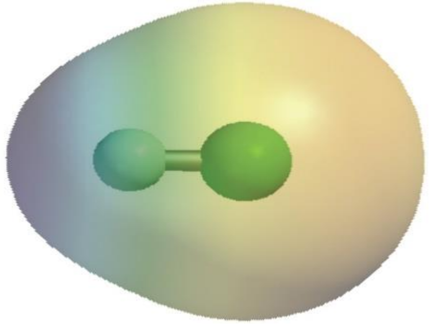
a. hydrogen   b. bromine   c. oxygen   d. nitrogen

3- Write the Lewis structure for each compound:

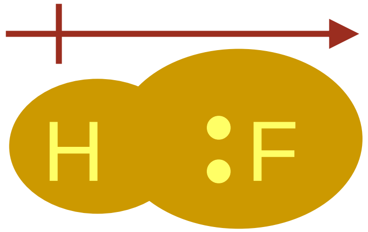
a. PH_3	b. SCl_2	c. HI	d. CH_4
e. NaF	f. CaO	g. $SrBr_2$	h. K_2O

Electronegativity and Bond Polarity:

- **Covalent bonding between unlike atoms** results in unequal sharing of the electrons.
  - One atom pulls the electrons in the bond closer to its side.
  - **One end of the bond has larger electron density** than the other.



- The result is a polar covalent bond.
  - Bond polarity
  - The end with the **larger electron density gets a partial negative charge.**
  - The end that is electron deficient gets a partial positive charge.



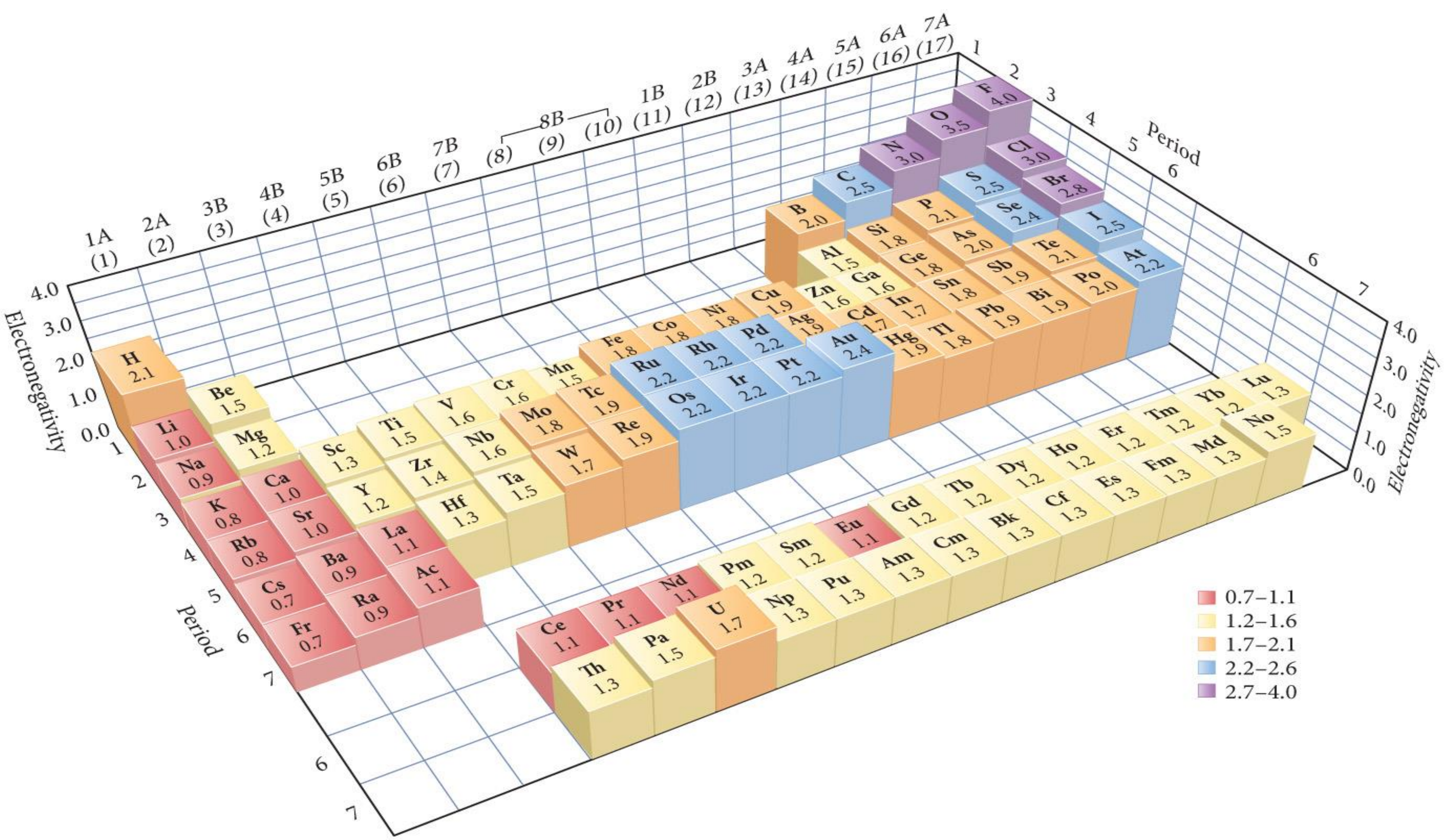
or $\overset{\delta+}{H}-\overset{\delta-}{F}$

Electronegativity and Bond Polarity:

- The ability of an atom to attract bonding electrons to itself is called electronegativity.
- **Increases** across period (left to right) and **decreases** down group (top to bottom)
  - Fluorine is the most electronegative element.
  - Francium is the least electronegative element.
  - Noble gas atoms are not assigned values.
  - Opposite of atomic size trend.
- The larger the difference in electronegativity, the more polar the bond.
  - Negative end toward more electronegative atom.

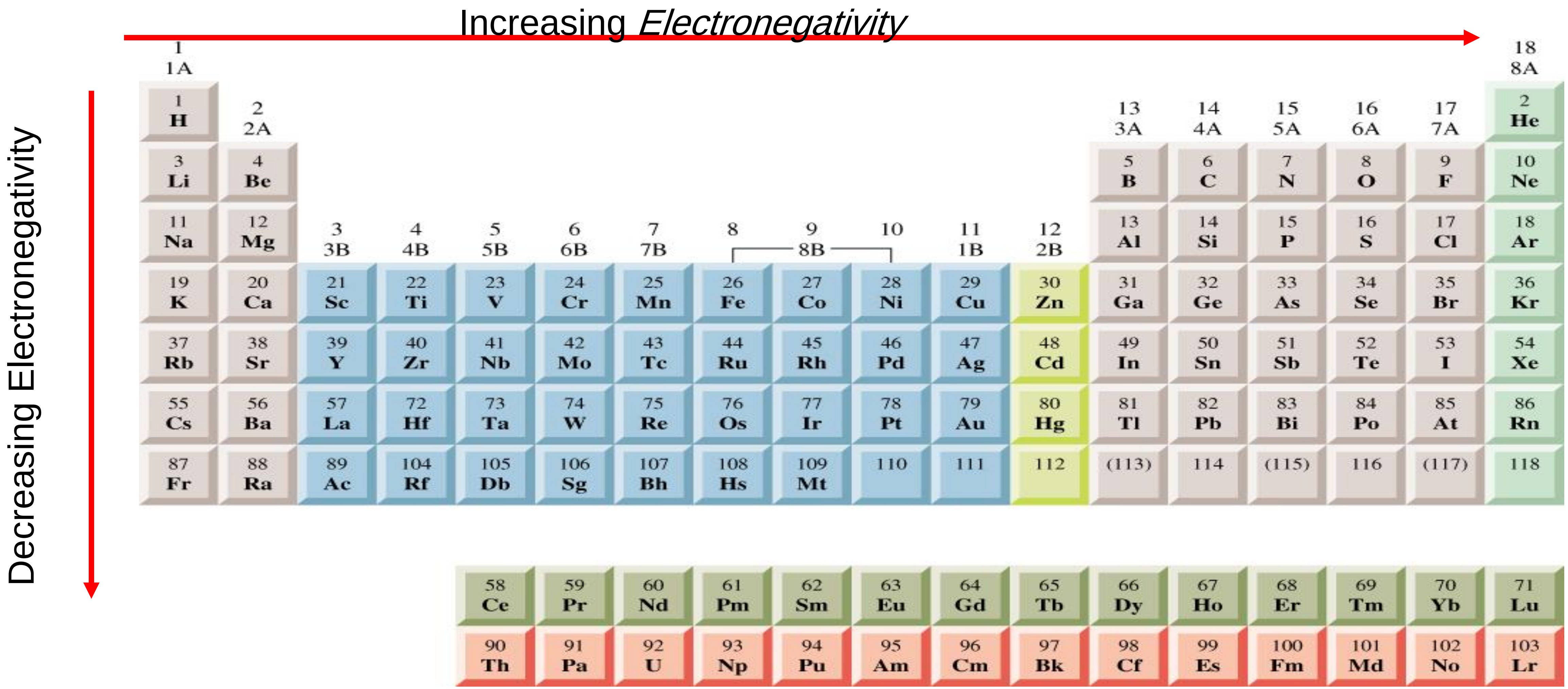
Chapter 4: Chemical Bonding and Chemical Reactions

Trends in Electronegativity



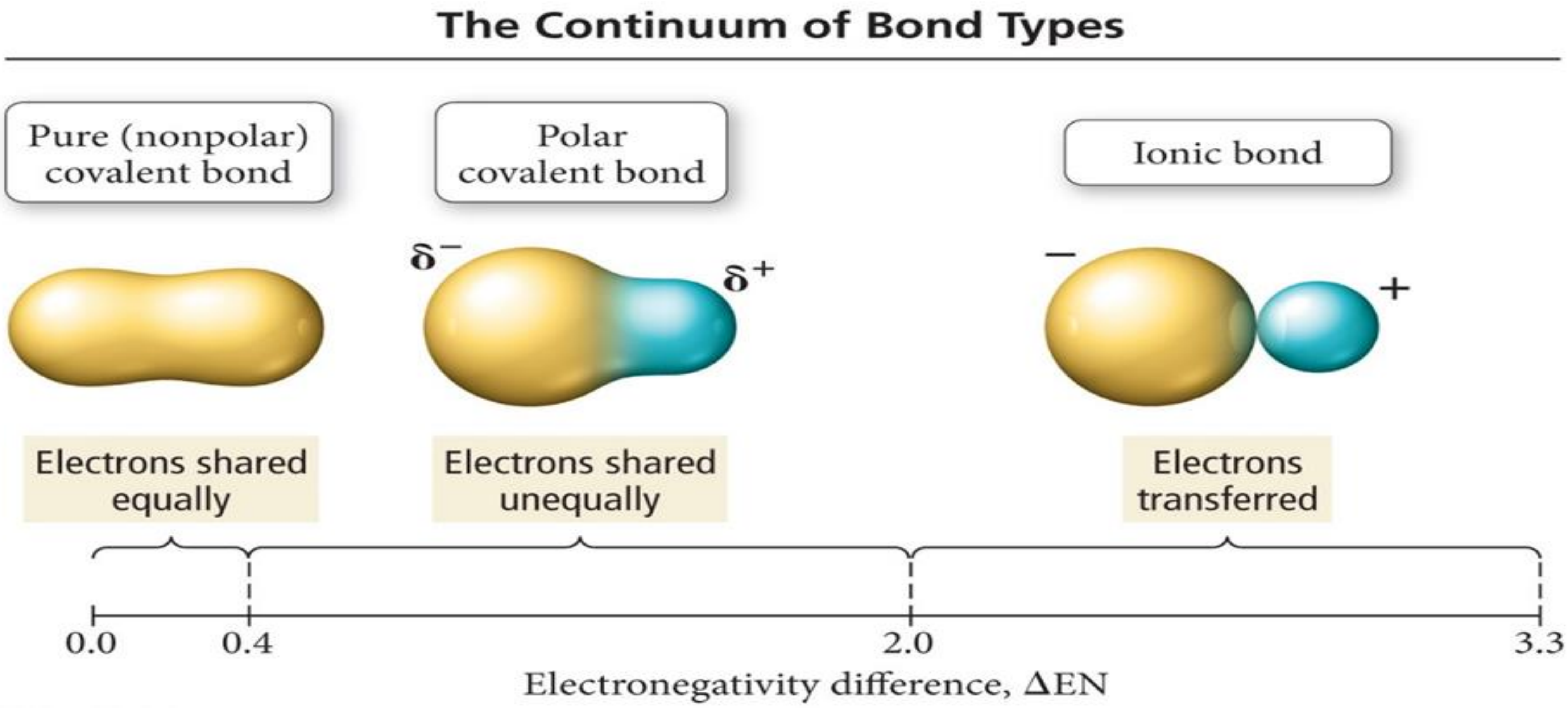
Chapter 4: Chemical Bonding and Chemical Reactions

General Trend in *Electronegativity*



Electronegativity and Bond Polarity:

- If the difference in electronegativity between bonded atoms is 0, the bond is pure covalent.
- If the difference in electronegativity between bonded atoms is 0.1 to 0.4, the bond is nonpolar covalent.
- If the difference in electronegativity between bonded atoms is 0.5 to 1.9, the bond is polar covalent.
- If difference in electronegativity between bonded atoms is larger than or equal to 2.0, it will be ionic.



The Continuum of Bond Types
Pure (nonpolar) covalent bond
Polar covalent bond
Ionic bond
δ−
δ+
−
+
Electrons shared equally
Electrons shared unequally
Electrons transferred
0.0
0.4
2.0
3.3
Electronegativity difference, ΔEN

Electronegativity and Bond Polarity:

TABLE 9.1 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

Electronegativity and Bond Polarity:

- Dipole moment, μ, is a measure of bond polarity.
  - A dipole is a material with a + and − end.
  - it is directly proportional to the size of the partial charges and **directly** proportional to the distance between them.
    - $\mu = (q)(r)$
    - Measured in Debyes, D
- Generally, the more electrons two atoms share and the larger the atoms are, the larger the dipole moment.

TABLE 9.2 Dipole Moments of Several Molecules in the Gas Phase

Molecule	ΔEN	Dipole Moment (D)
Cl_2	0	0
ClF	1.0	0.88
HF	1.9	1.82
LiF	3.0	6.33

Example

Classify the following bonds as ionic, polar covalent, or covalent:

1-The bond in CsCl:

ΔEN of Cs= – 0.7	ΔEN of Cl= – 3.0	3.0 – 0.7 = 2.3	Ionic

2- The bond in H_2S

ΔEN of H = – 2.1	ΔEN of S= – 2.5	2.5 – 2.1 = 0.4	Non-Polar Covalent

3- The NN bond in H_2NNH_2

ΔEN of N= – 3.0	ΔEN of N= – 3.0	3.0 – 3.0 = 0	Covalent

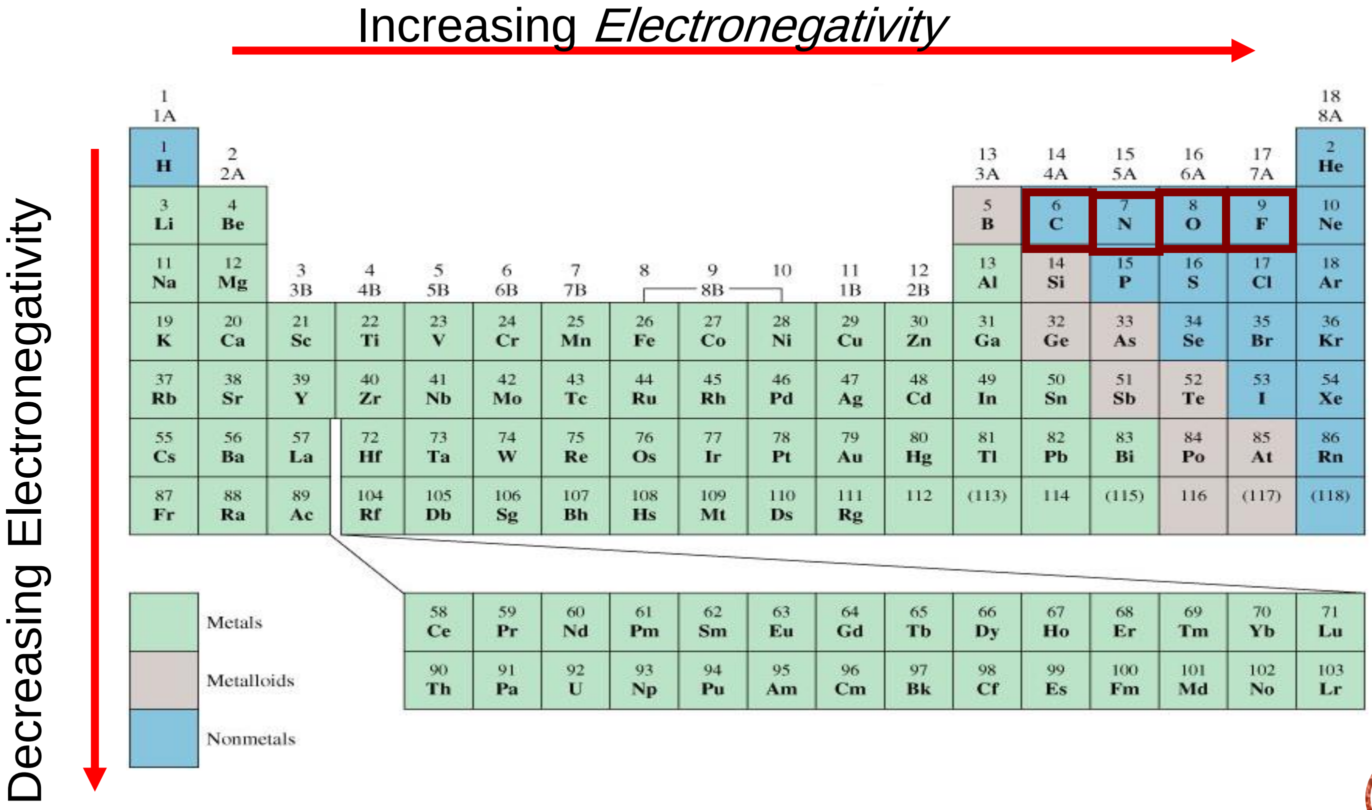
Which choice below correctly lists the elements in order of increasing electronegativity?

a. C < N < O < F

b. N < C < O < F

c. N < C < F < O

d. C < N < F < O



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

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

. Bond energies and bond lengths

- Chemical reactions involve breaking bonds in reactant molecules and making new bonds to create the products.
- The $\Delta H°_{reaction}$ can be estimated by comparing the cost of breaking old bonds to the income from making new bonds.
- The amount of energy it takes to break one mole of a bond in a compound
- is called the **bond energy**.

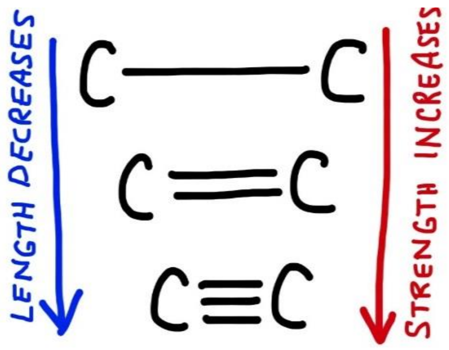
$$Cl_2(g) \rightarrow 2Cl(g) \qquad \Delta H = 243 \text{ kJ/mol}$$

$$C{\equiv}C \text{ (837 kJ)} > C{=}C \text{ (611 kJ)} > C{-}C \text{ (347 kJ)}$$

. Bond energies and bond lengths

TRENDS IN BOND ENERGIES

- In general, the more electrons two atoms share, the stronger the covalent bond.
  - Must be comparing bonds between like atoms
  - C≡C (837 kJ) > C═C (611 kJ) > C—C (347 kJ)
  - C≡N (891 kJ) > C ═ N (615 kJ) > C—N (305 kJ)



- In general, the **shorter the covalent bond, the stronger the bond.**
  - Must be comparing similar types of bonds
  - Br—F (237 kJ) > Br—Cl (218 kJ) > Br—Br (193 kJ)
  - Bonds **get weaker down the column.**
  - Bonds **get stronger across the period.**

. Bond energies and bond lengths

- **Bond length** is the distance between the nuclei of two bonded toms.
- **Bond length depends on**

1- the size of the atoms in the bond.

- Length is measured in Angstroms. (1 Å = 1 × 10^{-2} pm)

2- The electronegativity of the atoms in the bond.

	H - H	<	S - H	<	Cl - H	<	O - H	<	F - H
	2.1 2.1		2.5 2.1		3.0 2.1		3.5 2.1		4.0 2.1
Electronegativity Difference (Δ)	0		0.4		0.9		1.4		1.9

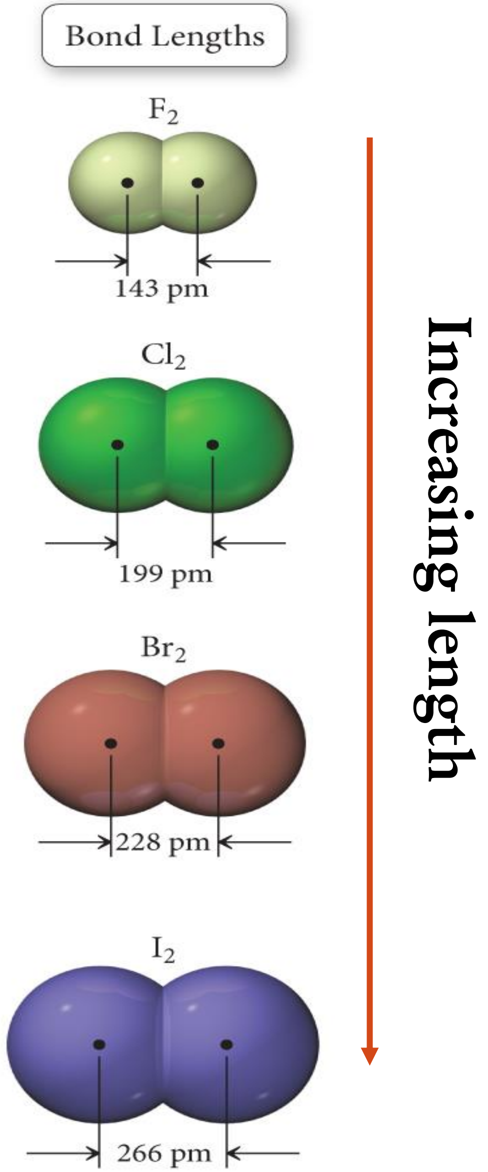
Covalent → Polar Covalent

increasing polarity

Decreasing length

3-The number of electrons between the atoms.

Single > double > triple



. Bond energies and bond lengths

TABLE 9.4 Average Bond Lengths

Bond	Bond Length (pm)	Bond	Bond Length (pm)	Bond	Bond Length (pm)
H—H	74	C—C	154	N—N	145
H—C	110	C═C	134	N═N	123
H—N	100	C≡C	120	N≡N	110
H—O	97	C—N	147	N—O	136
H—S	132	C═N	128	N═O	120
H—F	92	C≡N	116	O—O	145
H—Cl	127	C—O	143	O═O	121
H—Br	141	C═O	120	F—F	143
H—I	161	C—Cl	178	Cl—Cl	199
				Br—Br	228
				I—I	266

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

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