Chapter 1 - Fundamentals

Introduction to matter
Properties of matter
    Physical and chemical properties
    Extensive and intensive properties
    Temperature and density

Chapter 2 Atoms, Molecules, and Ions

2.3 Dalton's Atomic Theory
2.5 Early Experiments to Characterize the Atom
2.6 the modern view of Atomic Structure: An Introduction
2.7 Molecules and Ions
2.8 An Introduction to the Periodic Table
2.9 Naming Simple Compounds

Chapter 3 Stoichiometry

3.1 Atomic Masses
3.2 The Mole
3.3 Molar Mass
3.5 Percent Composition of Compounds

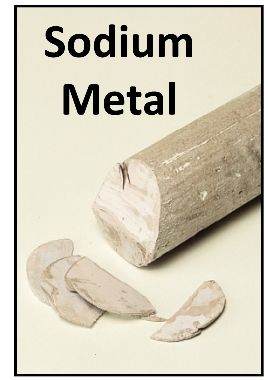
Introduction to Matter

• Physical material of the universe

• Anything that occupies space and has mass

• Exists in **three physical states** (solid, liquid, gas)

• **Matter is made up of atoms**

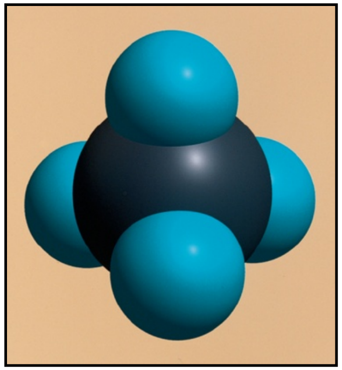


- **Atom:** Basic unit of any chemical element
- **Element:**
    - substance made up of atoms of the same kind (same atomic number)
    - Represented by **<u>symbols</u> of 1 or 2 letters (Co, Cu, H, O…)**
    - To date: 118 elements (periodic table)
    - **Allotropes** are <u>two or more distinct forms of an element</u> (O, O_2, O_3)

Introduction to Matter

- **Compound:** substance made up of atoms of 2 or more elements chemically united (H_2O, CO_2, NaCl…)
- **Substance:** a form of matter (element or compound) having a fixed composition and distinct identity

E.g.: water, iron, glucose…

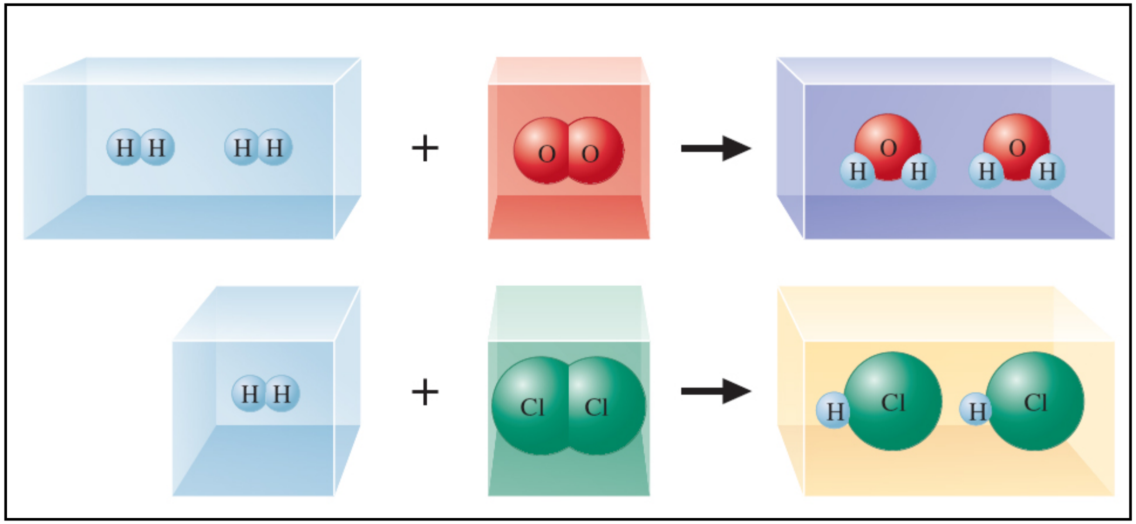


- **Mixture:**
  - Combination of 2 or more substances (tea, salted water, vinegar, mixture of sand + iron filling, air…)
  - **Homogenous mixture (or solution) :** composition is uniform throughout the sample (e.g., salted water, coca-cola, Tea…)
  - **Heterogeneous mixture:** composition is not uniform throughout the sample (e.g., Sand + iron filling, Water + sand…)

Physical and Chemical Properties of matter

- Physical property:
  - Can be measured without changing the identity
    - E.g.: Color, melting point, boiling point, optical density....

- Chemical property (Reactivity):
  - Describes the way a substance may change into another
    - E.g.: CH_4 + $2O_2$ ---> CO_2 + $2H_2O$



Extensive and intensive properties

- Extensive property:
  - Is additive ➔ depends on the amount of matter
    - E.g.: mass, volume, length…

- Intensive property:
  - Not additive ➔ does not depend on the amount of matter
    - E.g.: density, concentration, pressure, viscosity…

Temperature

A measure of the motion of particles in a system

Three systems for measuring the temperature:

- The Celsius scale (^oC): under 1 atm:
  - Zero: the freezing point of water
  - 100: boiling point of water

- The Kelvin scale (K):
  - Zero is the lowest temperature that can be attained theoretically = **- 273.15ºC**

$$\mathbf{T_k = T_C + 273.15}$$

Celsius and Kelvin scales have the same degree size but differ in the zero point.

- *Remark: The Fahrenheit scale: under 1 atm:*

  32 is freezing point of water; 212: boiling point of water ($T_F = (T_C \times 9/5) + 32$)

  Used in engineering sciences

  Differs from the Celsius and Kelvin scale in the zero point and in the degree size

Density

The mass of a substance per unit of volume of the substance:

$d = m/V$

Unit: **g/cm^3** or (**g/L for gases)**

Remark: The mass of an object is measured by comparing it to a standard mass of 1 kg, which is the basic SI unit for mass.

1 kg is the mass of I liter of water at 4°C.

The weight is a measure of the gravitational force (pull) on a given mass by the gravity.

Dalton's Atomic Theory

Dalton's Model

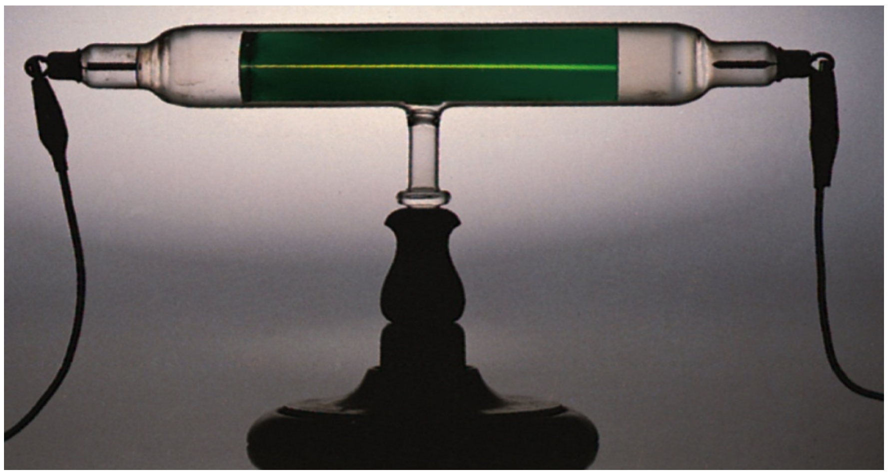
1. Each element is made up of tiny particles called atoms.
2. The atoms of a given element are identical; the atoms of different elements are different in some fundamental way or ways.
3. Chemical compounds are formed when atoms combine with one another. A given compound always has the same relative numbers and types of atoms.
4. Chemical reactions involve reorganization of the atoms – changes in the way they are bound together. The atoms themselves are not changed in a chemical reaction.

"Atomos" = Indivisible

Definition: the smallest particle of an element that retains the properties of that element

Early experiments to Characterize the Atom
The Electron

Thomson studied *cathode-ray tubes* and reasoned that all atoms must contain negatively charges particles called **electrons**.



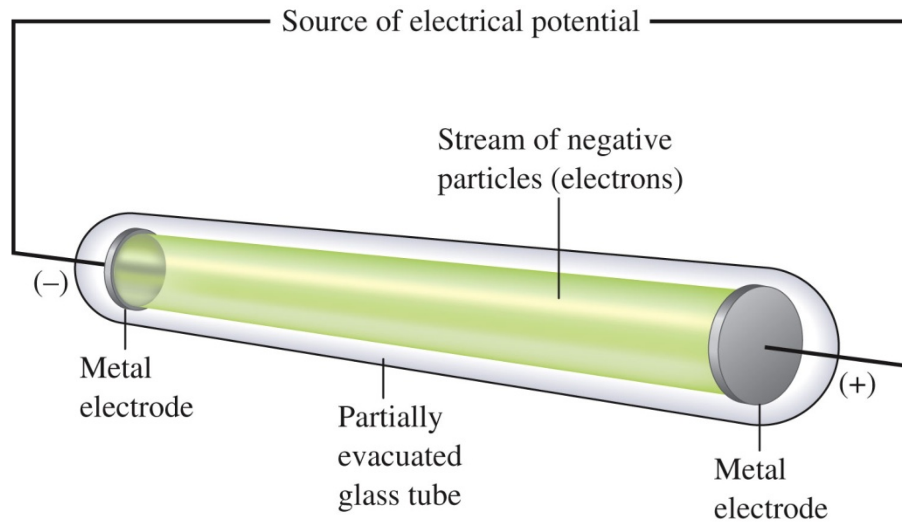


Figure 2.6 - A cathode-ray tube. The fast-moving electrons excite the gas in the tube, causing a glow between the electrodes. The green color in the photo is due to the response of the screen (coated with zinc sulfide) to the electron beam.

Since atoms were known to be electrically neutral, he further assumed that atoms also must contain positively charged particles.

Early experiments to Characterize the Atom
The Nuclear Atom

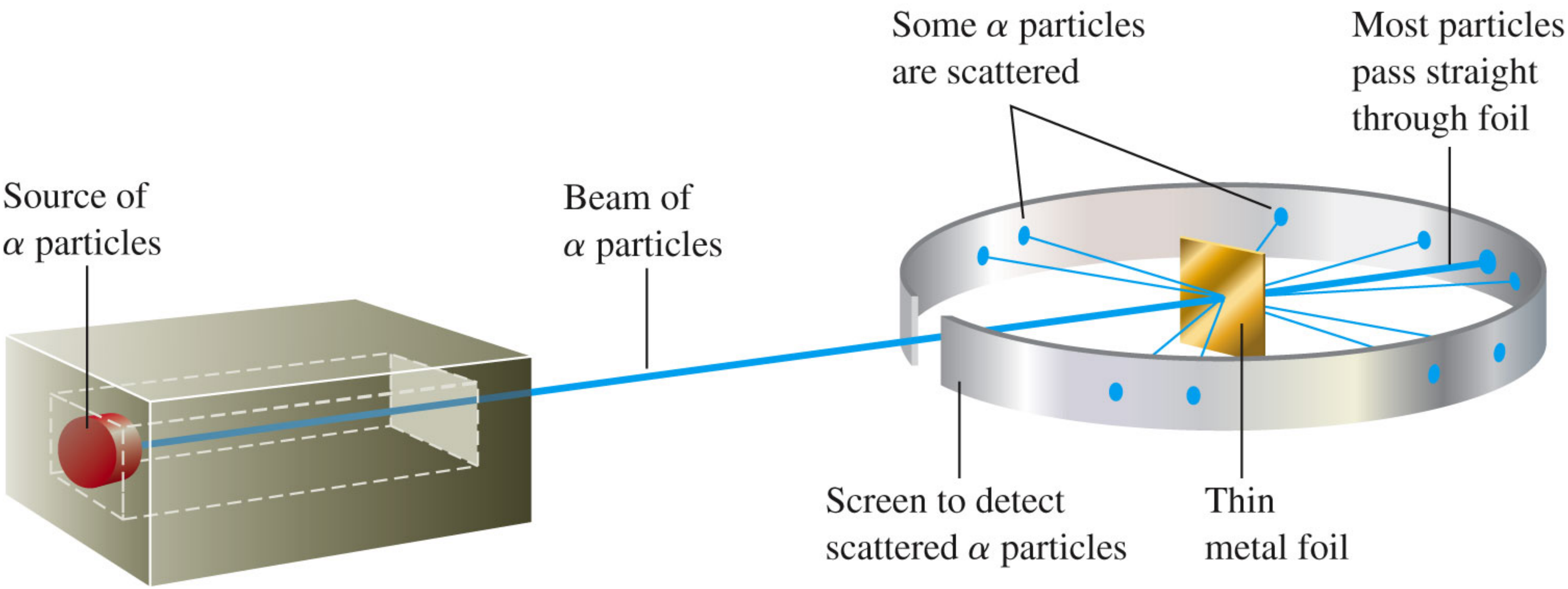
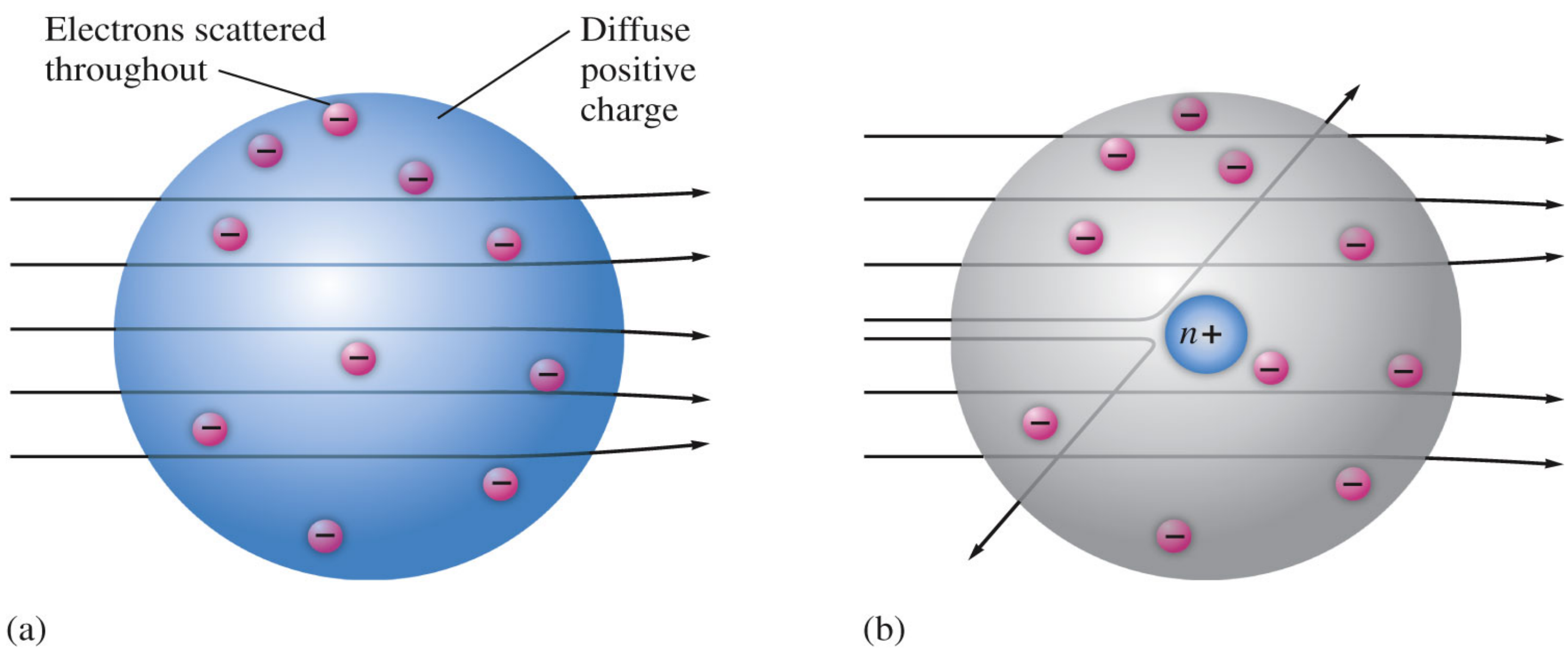


Figure 2.11 - Rutherford's experiment on α-particle bombardment of metal foil. (Gold foil was used in the original experiments because it can be hammered into extremely thin sheets.)

Early experiments to Characterize the Atom
The Nuclear Atom



__Plum pudding model:__ diffuse cloud of positive charge with the negative electrons embedded randomly in it.

__Actual atom model:__ dense center of positive charge (nucleus) and electrons moving around it.

Figure 2.11 - Rutherford's experiment on *a*-particle bombardment of metal foil. (Gold foil was used in the original experiments because it can be hammered into extremely thin sheets.)

Atomic Structure

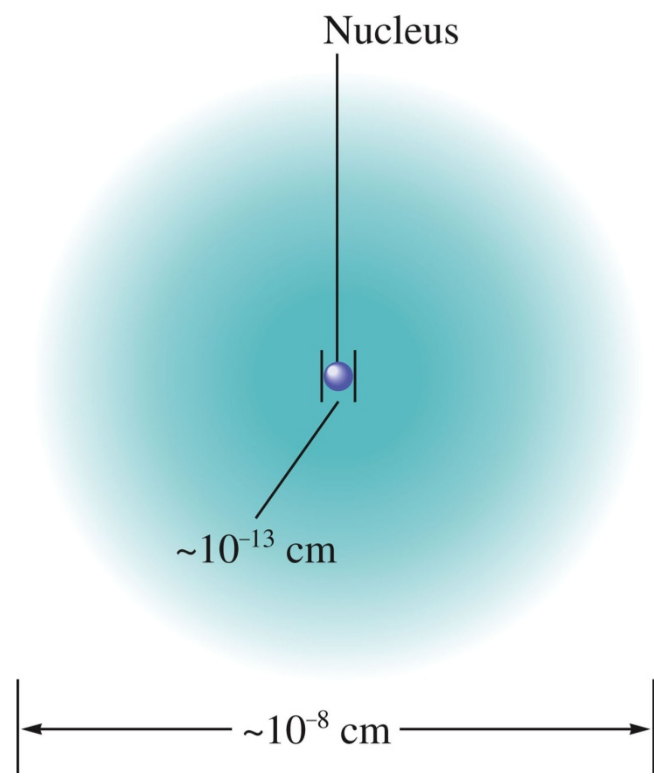
Atoms are composed from a nucleus (protons + neutrons) surrounded by an electronic cloud

Table 2.2

The Mass and Charge of the Electron, Proton, and Neutron

Particle	Mass	Charge*
Electron	9.11×10^{-31} kg	1−
Proton	1.67×10^{-27} kg	1+
Neutron	1.67×10^{-27} kg	None

*The magnitude of the charge of the electron and the proton is 1.60×10^{-19} C.



Representation of atoms:

$^{A}_{Z}\mathbf{X}$

A ------> <u>atomic mass</u> (nb of p^+ + nb of n^o)

Z ------> <u>atomic number</u> (nb of p^+)

Isotopes

Isotopes: atoms of a given element having the <u>same Z but # A</u>

same number of protons but different numbers of neutrons

E.g.: C-12 and C-13 ; H-1, H-2 and H-3; Na-23, Na-24

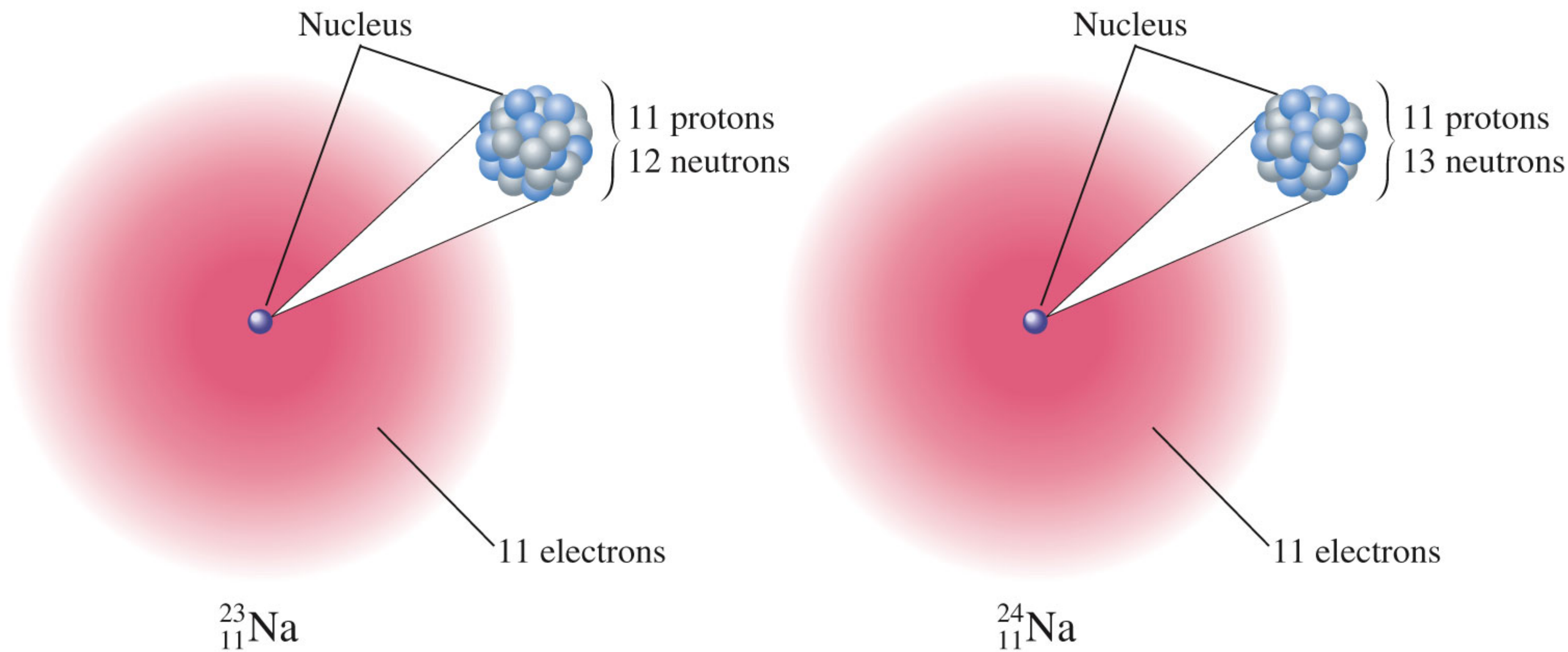


Figure 2.14 - Two isotopes of sodium. Both have 11 protons and 11 electrons, but they differ in the number of neutrons in their nuclei. Sodium-23 is the only naturally occurring form of sodium. Sodium-24 does not occur naturally but can be made artificially.

Molecules and ions

Dalton first recognized that chemical compounds were collections of atoms.

During the 20th century, scientists have learned that atoms have electrons and that these electrons participate in the bonding of one atom to another **(Chemical bonds)**.

- **Covalent bonds:** atoms are sharing electrons

- **Ionic bonds:** attraction between oppositely charged ions

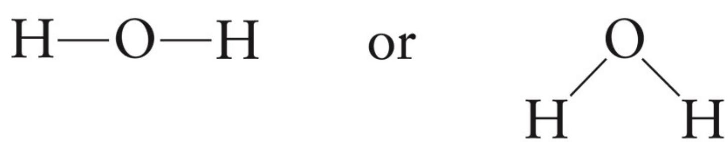
Molecules and ions

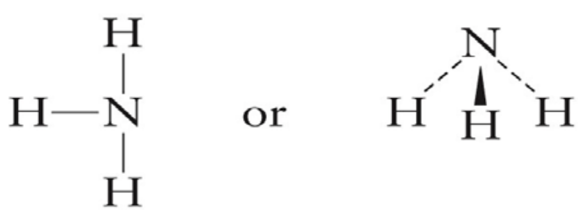
Molecules: collection of atoms

Molecules can be presented by a **chemical formula** or a **structural formula**.

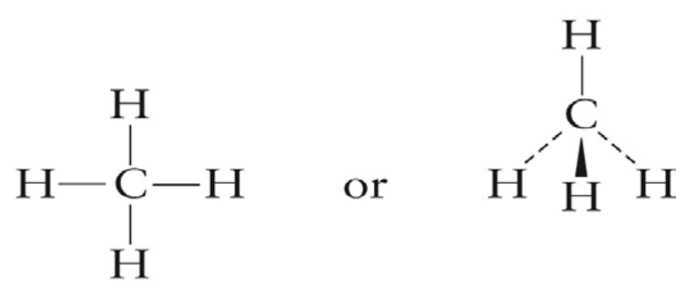
Chemical formula: water (H_2O), ammonia (NH_3), methane (CH_4)

Structural formula:





Ammonia



Methane

Molecules and ions

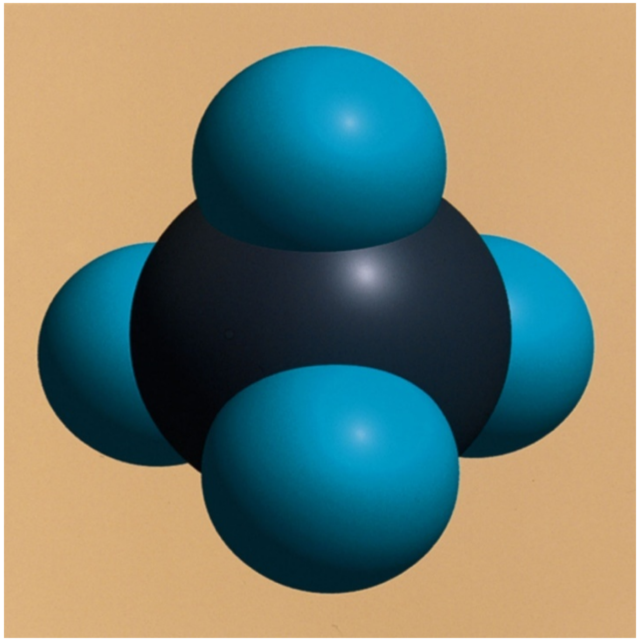
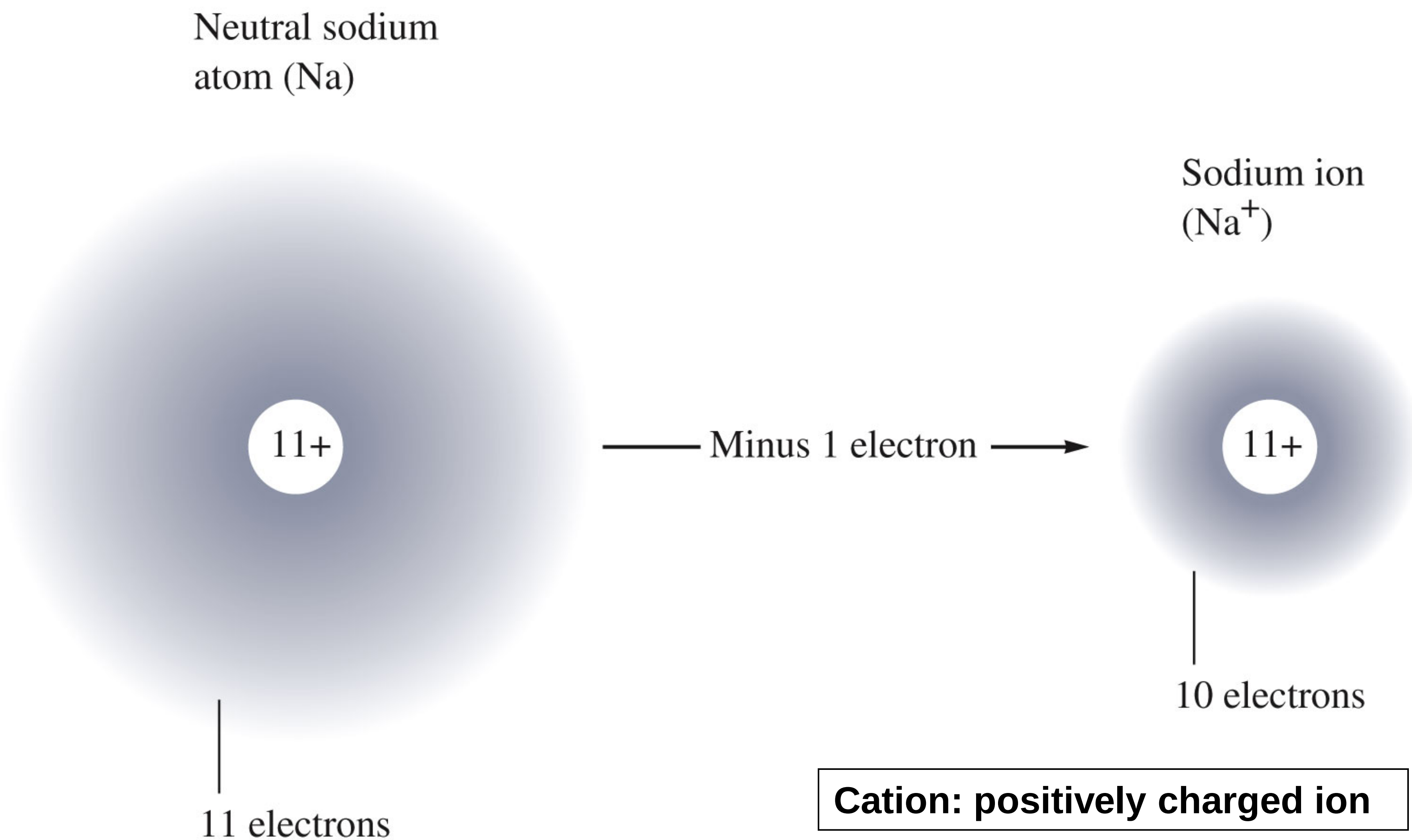


Figure 2.15 - Space-filling model of the methane molecule. This type of model shows both the relative sizes of the atoms in the molecule and their spatial relationships.

Figure 2.17 - Ball-and-stick model of methane.

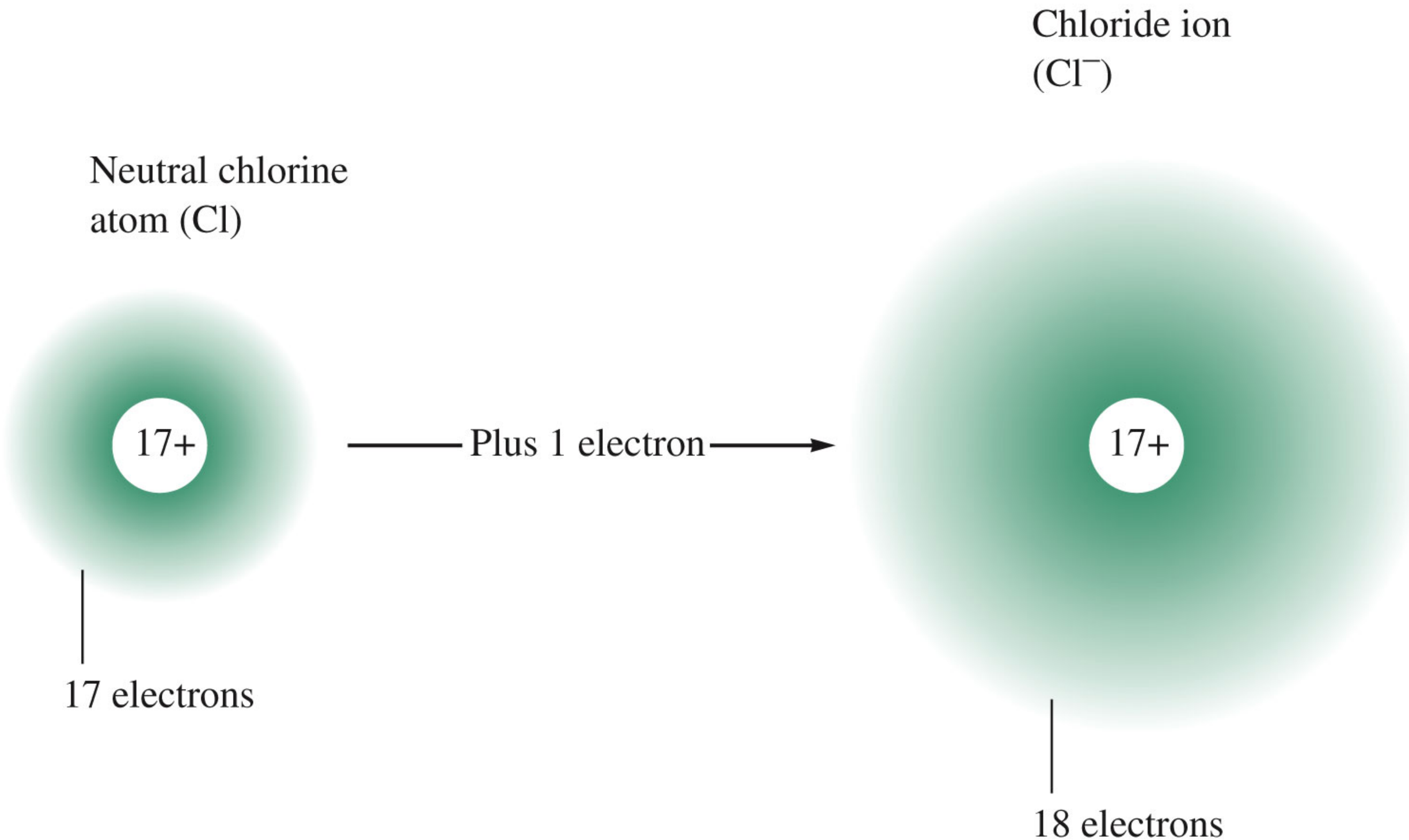
Molecules and ions

Ion: an atom or group of atoms that has a net positive or negative charge.



Cation: positively charged ion

Molecules and ions



Anion: negaively charged ion

Molecules and ions

Anions and cations attract each other ➡ ionic bond

The resulting solid is called **ionic solid or salt**.
Salts can be formed from simple ions (NaCl) or polyatomic ions (NH_4NO_3)

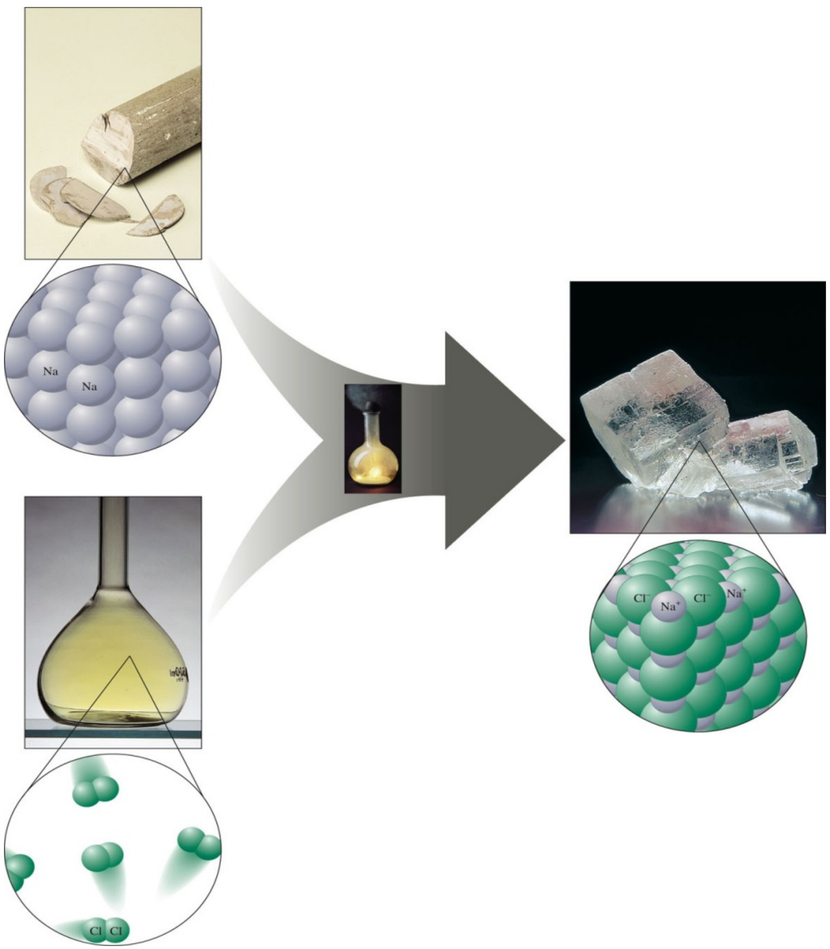


Figure 2.19 - Sodium metal (which is so soft it can be cut with a knife and which consists of individual sodium atoms) reacts with chlorine gas (which contains Cl_2 molecules) to form solid sodium chloride (which contains Na^+ and Cl^- ions packed together).

An Introduction to the Periodic Table

Alkali metals (group 1) · Alkaline earth metals (group 2) · Transition metals (groups 3–12) · Halogens (group 17) · Noble gases (group 18)

1 1A	2 2A	3	4	5	6	7	8	9	10	11	12	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 Ds	111 Rg	112 Cn	113 Uut	114 Uuq	115 Uup		117 Uus	118 Uuo

*Lanthanides	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
†Actinides	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

An Introduction to the Periodic Table

Metals: left-side of the Al-Po diagonal except hydrogen

- good conductor of electricity and heat, ductile, malleable and lustrous
- tend to lose electrons to form positive ions.

Nonmetals: right-side of the Al-Po diagonal

- poor conductor of electricity and heat, not ductile nor malleable, non-lustrous
- tend to gain electrons to form anions.
- bond to each other by forming covalent bonds.

Metalloides (elements found along the Al – Po diagonal):

- (Al, Si, Ge, As, Sb, Te, Po, At)
- have properties of both metals and nonmetals

An Introduction to the Periodic Table

- Elements in the same **vertical** columns (**groups**) have similar chemical properties.

Group	**Name**	**Elements**	**Properties**
1A	Alkali metals (wood ashes)	Li, Na, K, Rb, Cs, Fr	Monovalent cations
2A	Alkaline earth metals	Be, Mg, Ca, Sr, Ba, Ra	Divalent cations
6A	Chalcogens ("chalk formers")	O, S, Se, Te, Po	Divalent anions
7A	Halogens ("salt formers")	F, Cl, Br, I, At	Monovalent anions
8A	Noble gases (rare gases)	He, Ne, Ar, Kr, Xe, Rn	Inert

- *Group I A: very active, form ions with +1 charge, react with nonmetals.*
- *Group 7A: form diatomic molecules, react with metals to form salts.*
- *Group 8A: little chemical reactivity, exist as monoatomic gases.*

- The **horizontal** row of elements are called **periods**.

Naming Simple Compounds

Binary Compounds (Type I; Ionic)

Binary ionic compounds contain a positive ion (cation), always written first in the formula , and a negative ion.

- The cation is always named before the anion
- A monoatomic cation takes its name from the name of the element.
- A monoatomic anion is named by taking the first part of the element name and adding *–ide*.

Table 2.3

Common Monatomic Cations and Anions

Cation	Name	Anion	Name
H^+	Hydrogen	H^-	Hydride
Li^+	Lithium	F^-	Fluoride
Na^+	Sodium	Cl^-	Chloride
K^+	Potassium	Br^-	Bromide
Cs^+	Cesium	I^-	Iodide
Be^{2+}	Beryllium	O^{2-}	Oxide
Mg^{2+}	Magnesium	S^{2-}	Sulfide
Ca^{2+}	Calcium	N^{3-}	Nitride
Ba^{2+}	Barium	P^{3-}	Phosphide
Al^{3+}	Aluminum		
Ag^+	Silver		
Zn^{2+}	Zinc		

Naming Simple Compounds

Binary Compounds (Type I; Ionic)

A type I binary compound contains a metal that form only one type of cation.

Compound	*Ions Present*	*Name*
$NaCl$	Na^{+}, Cl^{-}	Sodium chloride
KI	K^{+}, I^{-}	Potassium iodide
CaS	Ca^{2+}, S^{2-}	Calcium sulfide
Li_3N	Li^{+}, N^{3-}	Lithium nitride
$CsBr$	Cs^{+}, Br^{-}	Cesium bromide
MgO	Mg^{2+}, O^{2-}	Magnesium oxide

Naming Simple Compounds

Binary Compounds (Type II; Ionic)

A type II binary compound contains a metal that form more than one type of positive ion and thus more than one type of ionic compound with a given anion.

- Fe^{2+} is iron(II) or Ferr**ous** ion

  → $FeCl_2$ Iron (II) chloride or Ferrous chloride

- Fe^{3+} is iron(III) or Ferr**ic** ion

  → $FeCl_3$ Iron (III) chloride or Ferric chloride

<u>*Common metals that do not require a Roman numeral are:*</u>

Group 1A, Group 2A, Aluminum (form only Al^{3+})

<u>Common transition metals that do not require a Roman numeral are:</u>

Zinc (only Zn^{2+}) and silver (only Ag^{+})

Naming Simple Compounds

Binary Compounds (Type II; Ionic)

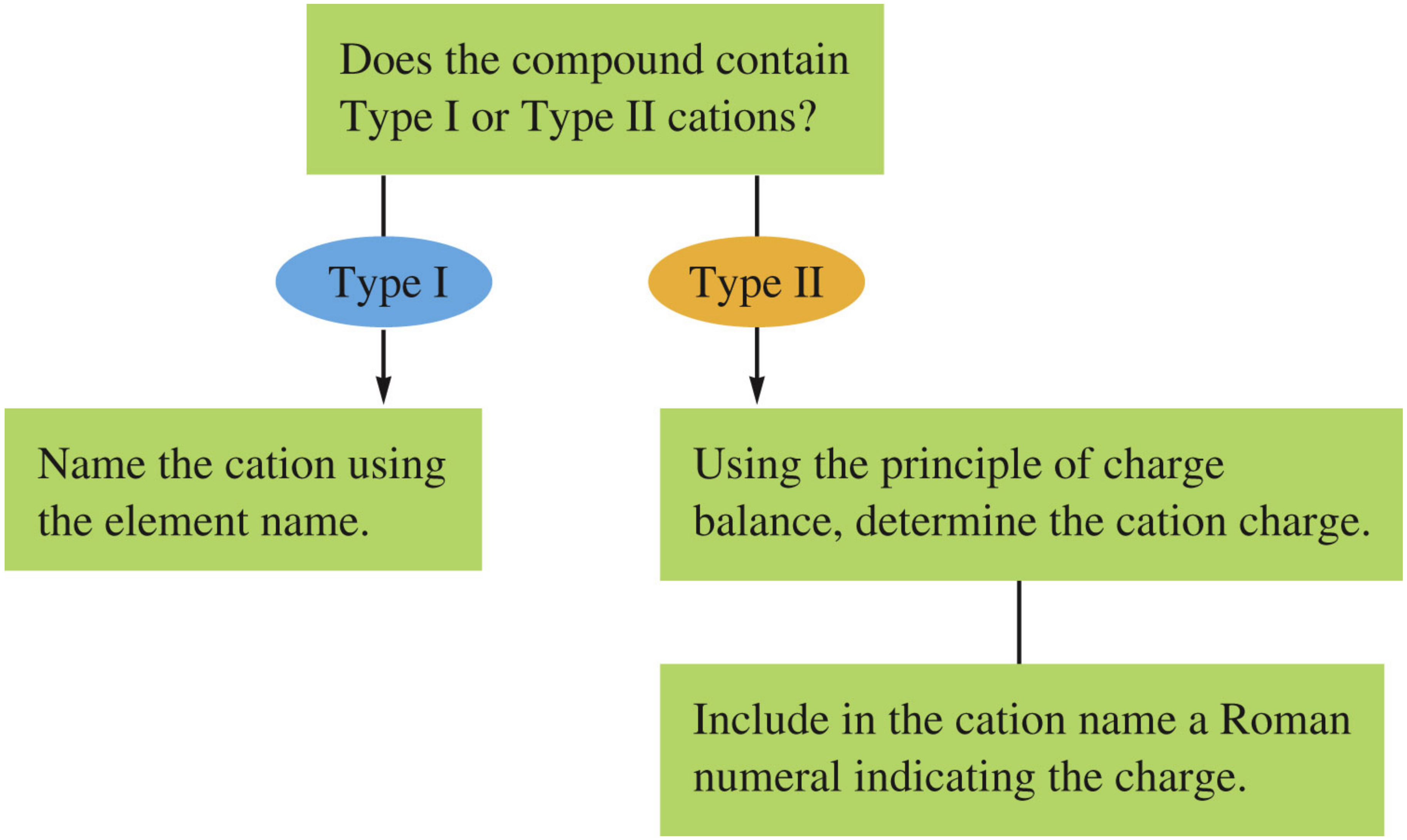
Table 2.4

Common Type II Cations

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

*Note that mercury(I) ions always occur bound together to form Hg_2^{2+}.

Naming Simple Compounds



Naming Simple Compounds

Example 2.2

Give the systematic name of each of the following compounds:

a. $CoBr_2$
b. $CaCl_2$
c. Al_2O_3
d. $CrCl_3$

Naming Simple Compounds

Example 2.2 - SOLUTION

a.	$CoBr_2$	Cobalt (II) bromide
b.	$CaCl_2$	Calcium chloride
c.	Al_2O_3	Aluminum oxide
d.	$CrCl_3$	Chromium (III) chloride

Naming Simple Compounds

Ionic Compounds with Polyatomic Ions

<u>Oxyanions:</u> Anions containing an atom of a given element and different numbers of oxygen atoms.

The name of the one with the larger number of oxygen ends in –***ate***.

The name of the one with the smaller number of oxygen ends in –***ite***.

SO_4^{2-} is the sulf**ate** ion

PO_4^{3-} is the phosph**ate** ion

SO_3^{2-} is the sulf**ite** ion

PO_3^{3-} is the phosph**ite** ion

When more than two oxyanions make-up a series, ***hypo***- and ***per***- are used as prefixes to name the members of the series with the fewest and the most oxygen atoms, respectively.

ClO_4^- is the **per**chlorate ion

ClO_3^- is the chlor**ate** ion

ClO_2^- is the chlor**ite** ion

ClO^- is the **hypo**chlorite ion

Naming Simple Compounds

Ionic Compounds with Polyatomic Ions

Table 2.5

Common Polyatomic Ions

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

Naming Simple Compounds

Binary Compounds (Type III; Covalent-contain 2 nonmetals)

Binary covalent compounds are formed between 2 nonmetals

- The first element in the formula is named first, using the full element name.
- The second element is named as if it were an anion.
- Prefixes are used to denote the numbers of atoms present.
- The prefix mono- is never used for naming the first element (CO: carbon monoxide)

Table 2.6

Prefixes Used to Indicate Number in Chemical Names

Prefix	Number Indicated
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8

Naming Simple Compounds

Binary Compounds (Type III; Covalent-contain 2 nonmetals)

Compound	*Systematic name*	*Common name*
N_2O	Dinitrogen monoxide	Nitrous oxide
NO	Nitrogen monoxide	Nitric oxide
NO_2	Nitrogen dioxide	
N_2O_3	Dinitrogen trioxide	
N_2O_4	Dinitrogen tetroxide	
N_2O_5	Dinitrogen pentoxide	

Some common names: H_2O (water), NH_3 (ammonia)

Naming Simple Compounds

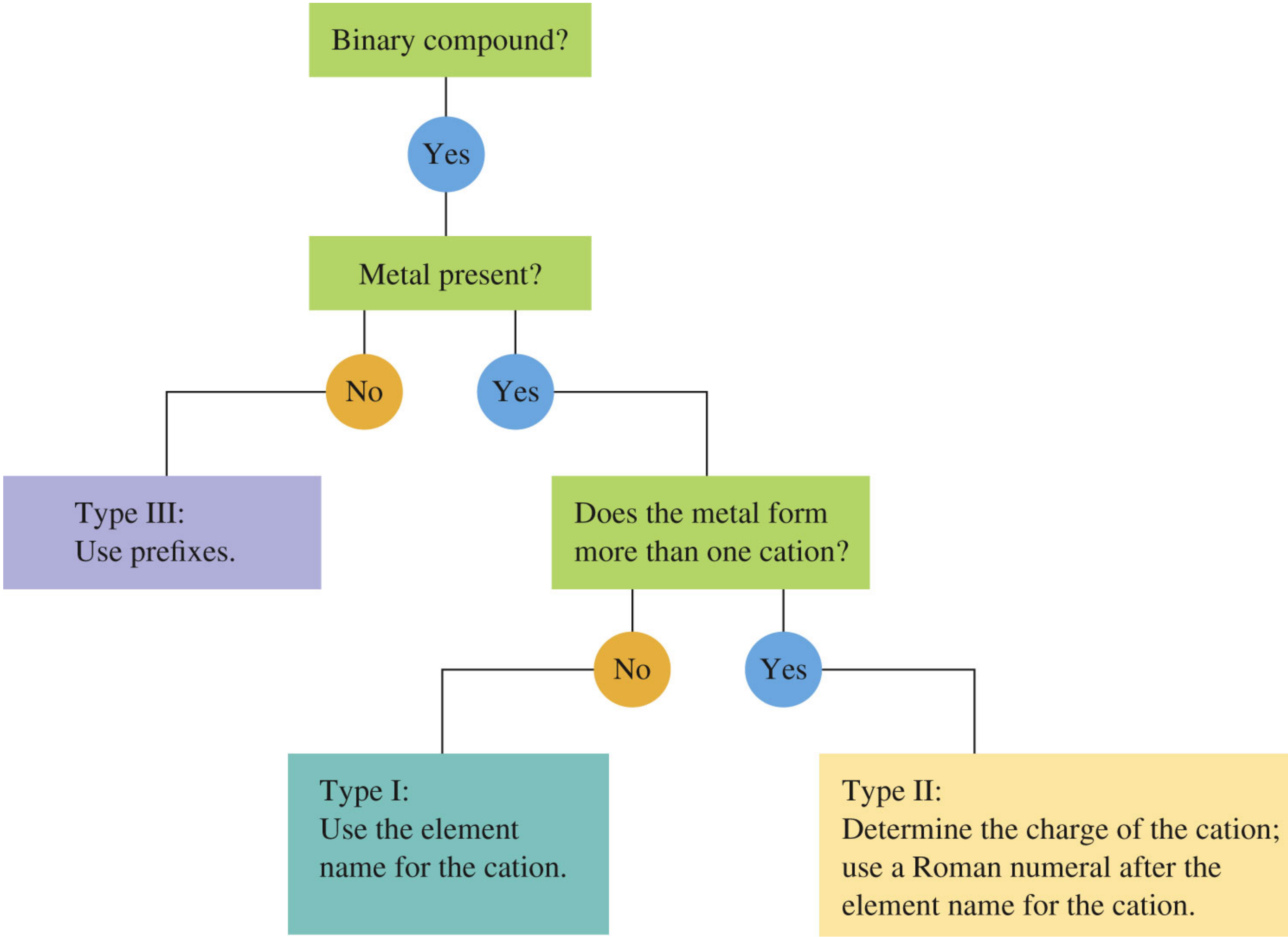


Figure 2.21 - A flowchart for naming binary compounds.

Naming Simple Compounds

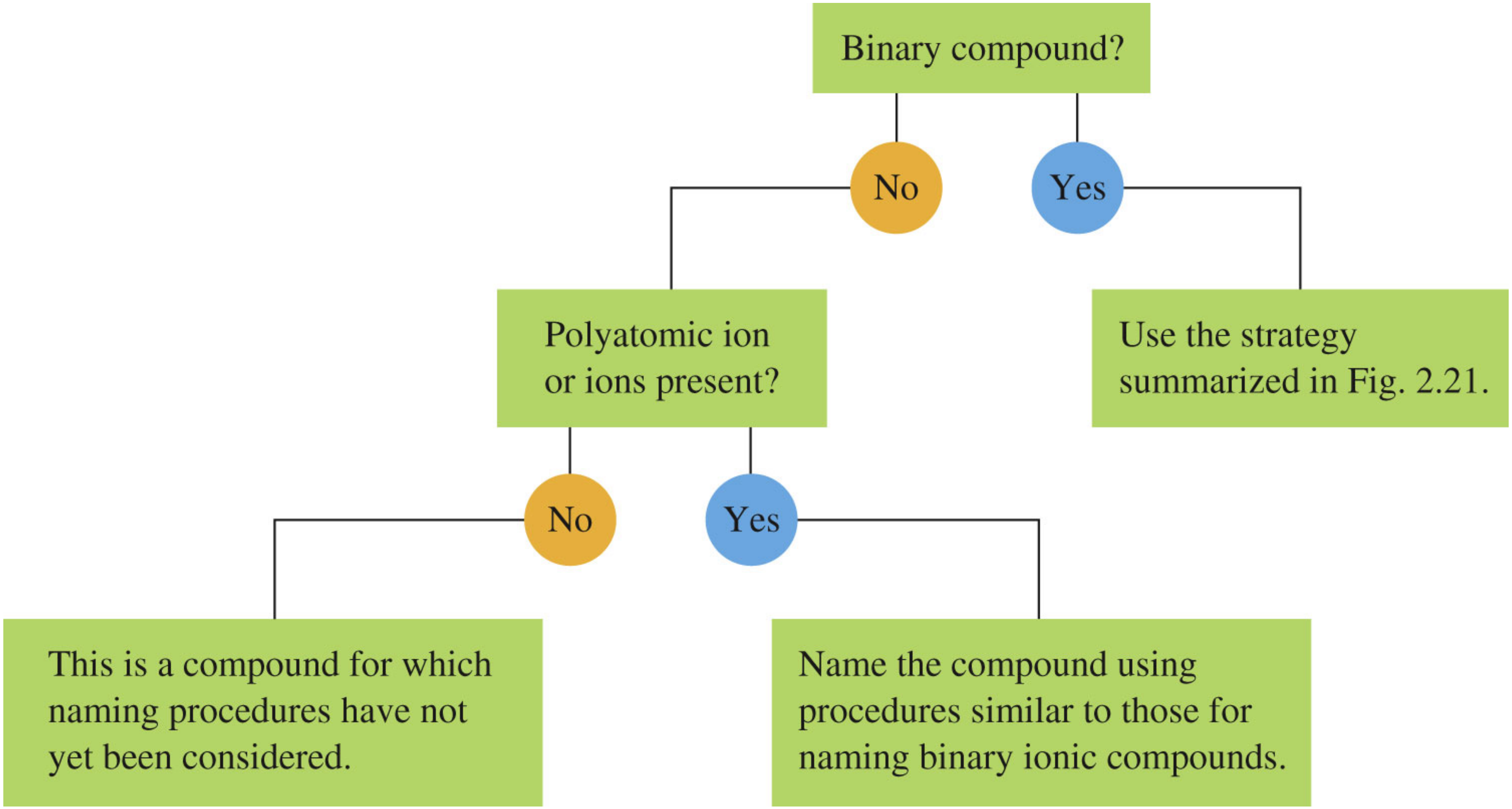


Figure 2.22 - Overall strategy for naming chemical compounds.

Naming Simple Compounds

Example 2.3

Give the systematic name of each of the following compounds:

a. Na_2SO_4
b. KH_2PO_4
c. $Fe(NO_3)_3$
d. $Mn(OH)_2$
e. Na_2SO_3
f. Na_2CO_3
g. $NaHCO_3$
h. $CsClO_4$
i. $NaOCl$
j. Na_2SeO_4
k. $KBrO_3$

Naming Simple Compounds

Example 2.3 - SOLUTION

Give the systematic name of each of the following compounds:

a. Na_2SO_4	sodium sulfate
b. KH_2PO_4	Potassium dihydrogen phosphate
c. $Fe(NO_3)_3$	Iron (III) nitrate
d. $Mn(OH)_2$	Manganese (II) hydroxide
e. Na_2SO_3	Sodium sulfite
f. Na_2CO_3	Sodium carbonate
g. $NaHCO_3$	Sodium hydrogen carbonate
h. $CsClO_4$	Cesium perchlorate
i. $NaOCl$	Sodium hypochlorite
j. Na_2SeO_4	Sodium selenate
k. $KBrO_3$	Potassium bromate

Naming Simple Compounds

Example 2.4

Given the following systematic names, write the formula for each compound.

a. Ammonium sulfate
b. Vanadium (V) fluoride
c. Dioxygen difluoride
d. Rubidium peroxide
e. Gallium oxide

Naming Simple Compounds

Example 2.4- SOLUTION

Given the following systematic names, write the formula for each compound.

a. Ammonium sulfate	$(NH_4)_2SO_4$
b. Vanadium (V) fluoride	VF_5
c. Dioxygen difluoride	O_2F_2
d. Rubidium peroxide	Rb_2O_2
e. Gallium oxide	Ga_2O_3

Naming Simple Compounds

Acids

An acid can be viewed as a molecule with one or more H^+ ions attached to an anion.

- If the anion **does not contain oxygen**, the acid is named with the prefix ***hydro***- and the suffix –***ic***.

Table 2.7

Names of Acids That Do Not Contain Oxygen

Acid	Name
HF	Hydrofluoric acid
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HI	Hydroiodic acid
HCN	Hydrocyanic acid
H_2S	Hydrosulfuric acid

Naming Simple Compounds

Acids

- If the anion **contains oxygen**, the acid name is formed from the root name of the anion with a suffix of –***ic*** or –***ous***.

  - if anion name ends in –***ate***, the acid name ends with –***ic***.
  - if anion has –***ite*** ending, the acid name ends with –***ous***.

Table 2.8

Names of Some Oxygen-Containing Acids

Acid	Name
HNO_3	Nitric acid
HNO_2	Nitrous acid
H_2SO_4	Sulfuric acid
H_2SO_3	Sulfurous acid
H_3PO_4	Phosphoric acid
$HC_2H_3O_2$	Acetic acid

Naming Simple Compounds

Acids

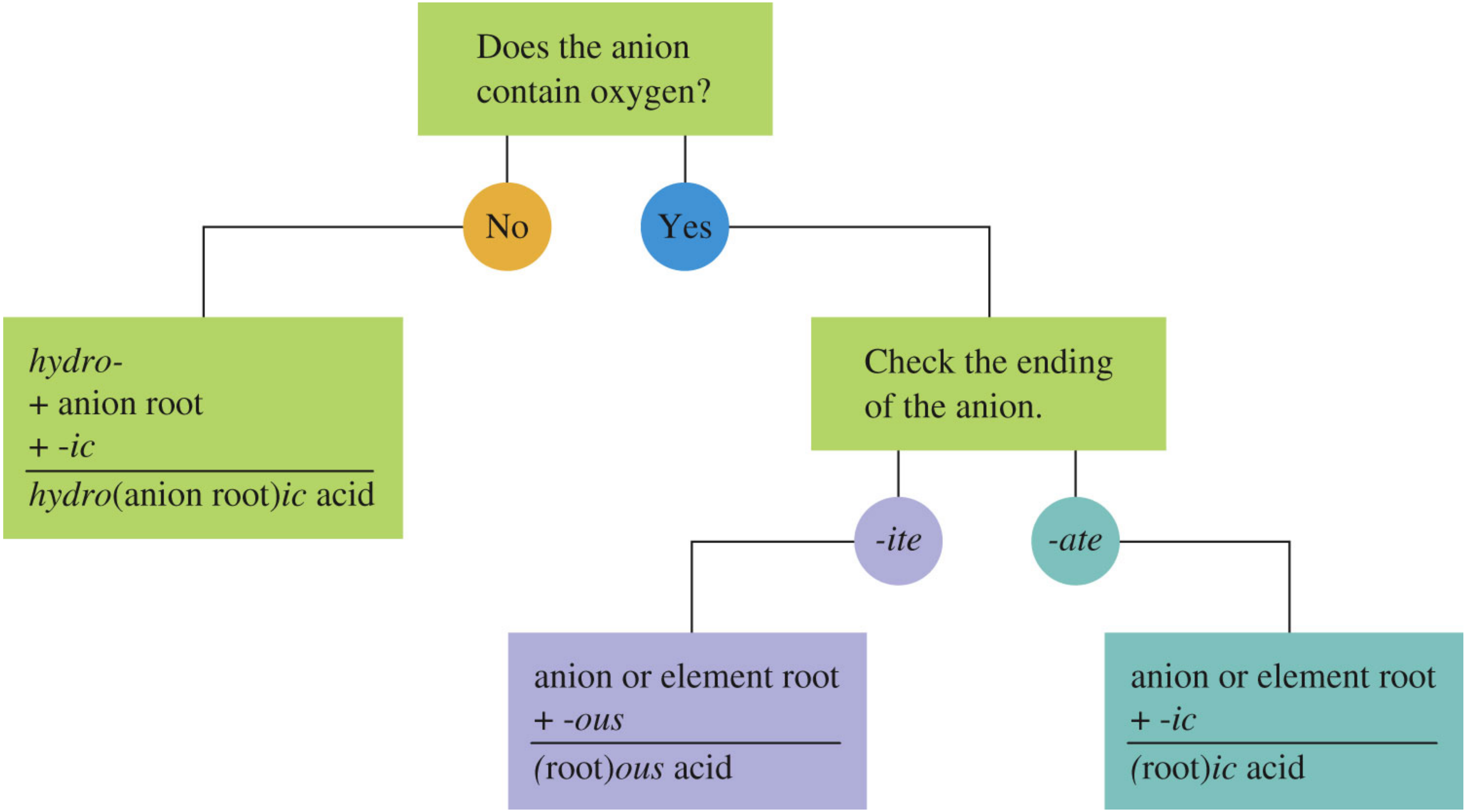


Figure 2.23 - A flowchart for naming acids. The acid has one or more H^+ ions attached to an anion.

Acids containing ions ending with **ide** often become **hydro -ic** acid

Cl^- (chloride)	=>	HCl	hydrochloric acid
F^- (fluoride)	=>	HF	hydrofluoric acid
S^{2-} (sulfide)	=>	H_2S	hydrosulfuric acid
CN^- (cyanide)	=>	HCN	hydrocyanic acid

Acids containing ions ending with **ate** usually become **-ic** acid

$CH_3CO_2^-$	(acetate)	=>	CH_3CO_2H	acetic acid
CO_3^{2-}	(carbonate)	=>	H_2CO_3	carbonic acid
BO_3^{3-}	(borate)	=>	H_3BO_3	boric acid
NO_3^-	(nitrate)	=>	HNO_3	nitric acid
SO_4^{2-}	(sulfate)	=>	H_2SO_4	sulfuric acid
ClO_4^-	(perchlorate)	=>	$HClO_4$	perchloric acid
PO_4^{3-}	(phosphate)	=>	H_3PO_4	phosphoric acid
MnO_4^-	(permanganate)	=>	$HMnO_4$	permanganic acid
CrO_4^{2-}	(chromate)	=>	H_2CrO_4	chromic acid
ClO_3-	(chlorate)	=>	$HClO_3$	chloric acid

Acids containing ions ending with **ite** usually become **-ous** acid

ClO_2^-	(chlorite)	=>	$HClO_2$	chlorous acid
NO_2^-	(nitrite)	=>	HNO_2	nitrous acid
SO_3^{2-}	(sulfite)	=>	H_2SO_3	sulfurous acid

Atomic Masses

- Based on the mass of ^{12}C as the standard

- C-12 is assigned a mass of exactly 12 atomic mass units (amu)

- Masses of other atoms are given relative to C-12 ➔ Relative Atomic masses

- Relative atomic masses are easily calculated by mass spectrometry.

Atomic Masses

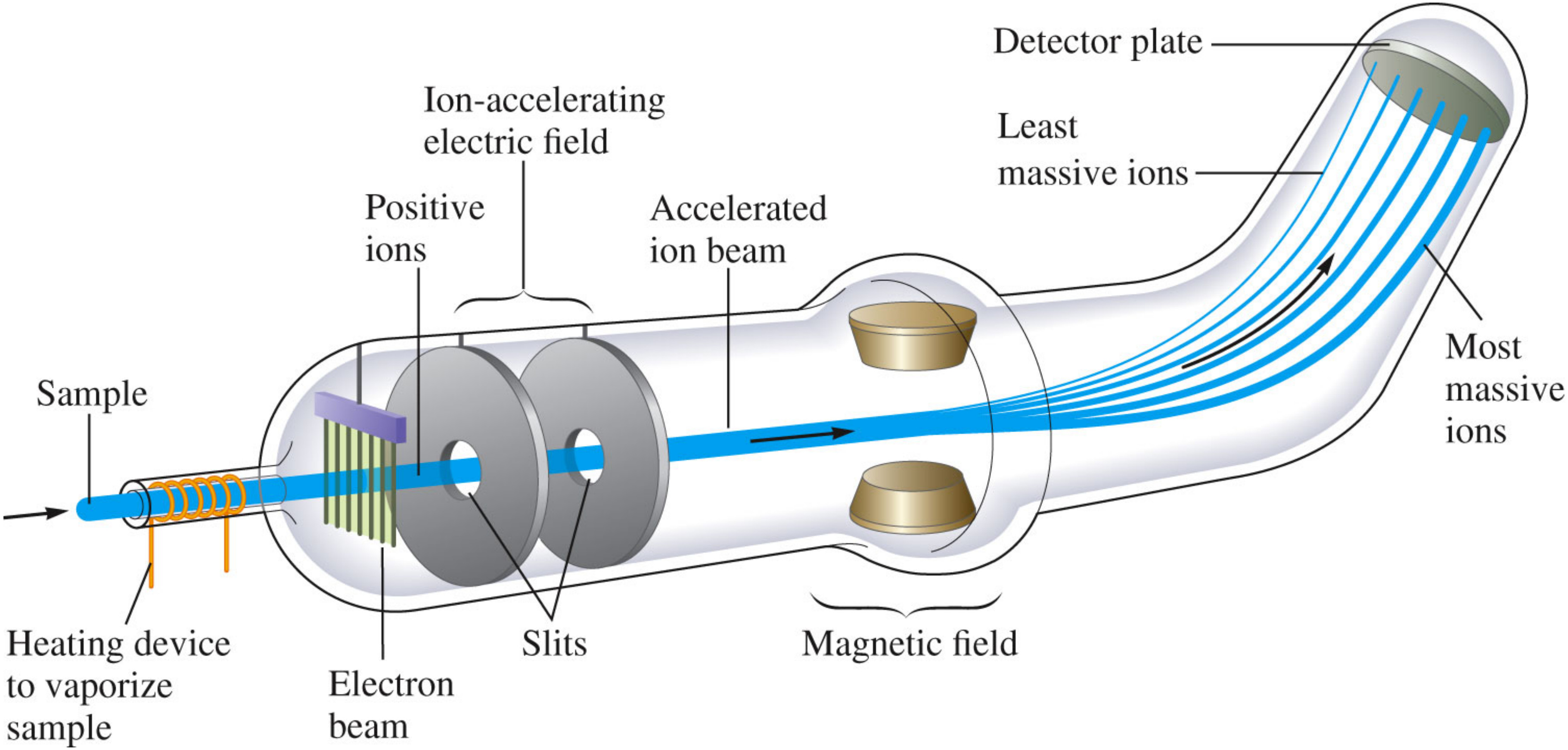


Figure 3.1 - Schematic diagram of a mass spectrometer.

<u>Finding the mass of an element:</u>

From mass spectrometery : mass of ^{13}C / mass of ^{12}C = 1.0836129

So the mass of a ^{13}C atom is (1.0836129) (12 amu) = 13.003355 amu

Average Atomic Mass

- Elements occur in nature as mixtures of isotopes

- Atomic mass is based on the relative abundance of isotopes

E.g.: carbon in nature:
98.89% ^{12}C, 1.11% ^{13}C and <0.01% ^{14}C (negligable)

Average atomic mass of natural carbon =
(98.89 /100x 12 amu) + (1.11/100 x 13.003 amu) =
12.01 amu

- The average atomic mass is often called the atomic weight

Isotopic Composition of Ne

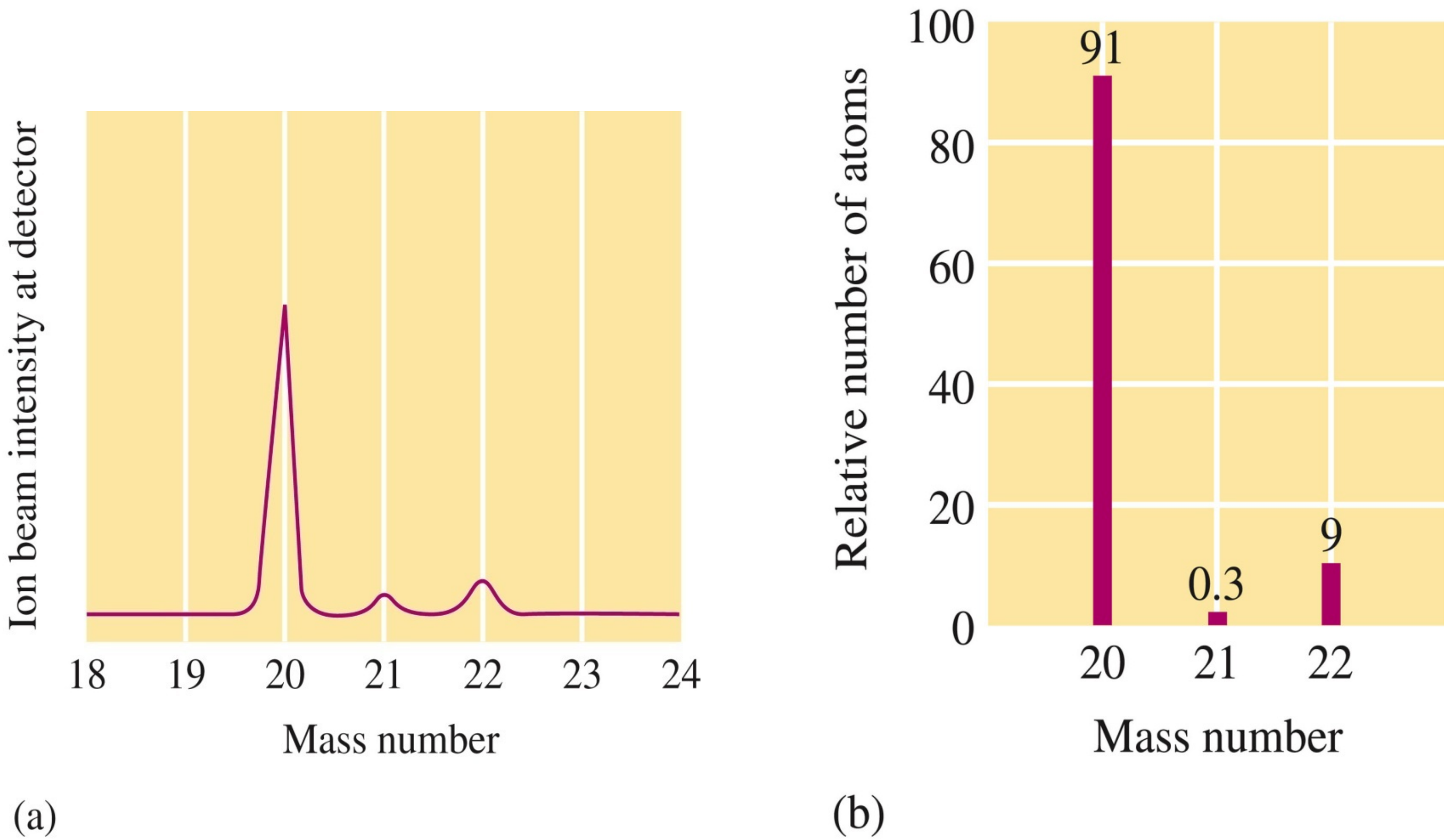
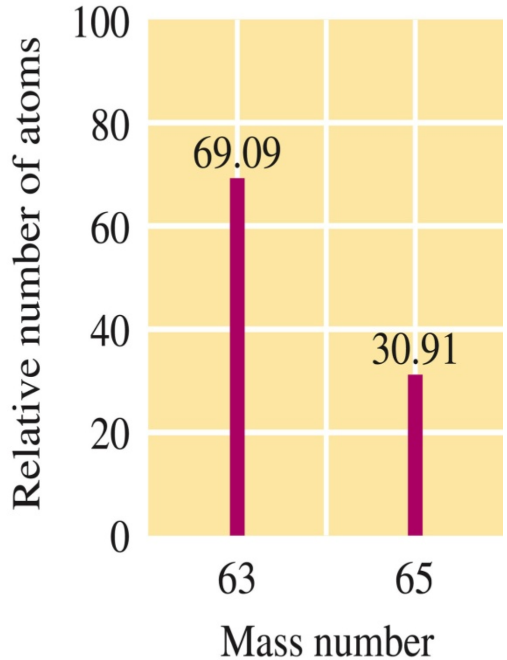


Figure 3.2 - The relative intensities of the signals recorded when natural neon is injected into a mass spectrometer, represented in terms of (a) "peaks" and (b) a bar graph. The relative areas of the peaks are 0.9092 (^{20}Ne), 0.00257 (^{21}Ne), and 0.0882 (^{22}Ne); natural neon is therefore 90.92% ^{20}Ne, 0.257% ^{21}Ne, and 8.82% ^{22}Ne.

Example 3.1

Copper is a very important metal used for water pipes, electrical wiring, roof coverings, and other materials. When a sample of natural copper is vaporized and injected into a mass spectrometer, the results shown in Fig. 3.3 are obtained. Use these data to compute the average mass of natural copper. (the mass values for ^{63}Cu and ^{65}Cu are 62.93 amu and 64.93 amu, respectively)



Example 3.1

Copper is a very important metal used for water pipes, electrical wiring, roof coverings, and other materials. When a sample of natural copper is vaporized and injected into a mass spectrometer, the results shown in Fig. 3.3 are obtained. Use these data to compute the average mass of natural copper. (the mass values for ^{63}Cu and ^{65}Cu are 62.93 amu and 64.93 amu, respectively)

Solution

$$\text{Average atomic mass} = (0.6909)\,(62.93) + (0.309)\,(64.93)$$
$$= 63.55 \text{ amu/atom}$$

The Mole

- Relates the atomic mass to a unit of gram for lab purposes

- The number of carbon atoms in exactly 12 grams of pure ^{12}C

- Modern techniques have been used to define this number as **6.022137×10^{23}.** This number is calle**d Avogadro's number** (N_A)

- One mole of something consists of 6.022×10^{23} units of that substance. Just like 1 dozen eggs is 12 eggs

1 mole of C is 12 grams and contains 6.022×10^{23} atoms of C

Table 3.1

Comparison of 1-Mole Samples of Various Elements

Element	Number of Atoms	Mass of Sample (g)
Aluminum	6.022×10^{23}	26.98
Gold	6.022×10^{23}	196.97
Iron	6.022×10^{23}	55.85
Sulfur	6.022×10^{23}	32.07
Boron	6.022×10^{23}	10.81
Xenon	6.022×10^{23}	131.30

The Mole

Thus the mole is defined such that a sample of a natural element with a mass equal to the element's atomic mass expressed in grams contains 1 mole of atoms.

$$(6.022\ 10^{23} \text{ atoms})\ (12 \text{ amu/atom}) = 12\text{g}$$

$$6.022\ 10^{23} \text{ amu} = 1 \text{ g}$$

Or

$$1 \text{ amu} = 1.66053886 \times 10^{-24} \text{ grams}$$

The Mole

Example 3.2

Americium (^{243}Am) is an element that does not occur naturally, it can be made in very small amounts in a device called a particle accelerator. Compute the mass in grams of a sample of americium containing 6 atoms.

The Mole

Example 3.2 - SOLUTION

^{243}Am

6 atoms x 243 amu/atom = 1.46×10^3 amu

Since 6.022×10^{23} amu = 1 g, the mass of 6 Am atom in grams is:

1.46×10^3 amu x 1g / 6.022×10^{23} amu = 2.42×10^{-21} g

The Mole

Example 3.3

A silicon ship used in an integrated circuit of a microcomputer has a mass of 5.68 mg. How many silicon (Si) atoms are present in this ship?

$^{28.09}Si$

The Mole

Example 3.3 –SOLUTION

5.68×10^{-3} g Si x 1 mol Si / 28.09 g Si = 2.02×10^{-4} mol Si

2.02×10^{-4} mol Si x $6.022 .10^{23}$ atoms / 1 mol Si = 1.22×10^{20} atoms

Molar Mass

A substance's molar mass (or molecular weight) is the mass in grams of one mole of the compound

Molar mass of CO_2?

Mass of 1 mole of C = 12.01

Mass of 2 mol of O = 2 x 16.00

Mass of 1 mol of CO_2 = 44.01

Molar Mass

Example 3.4

Isopentyl acetate ($C_7H_{14}O_2$), the compound responsible for the scent of bananas, can be produced commercially. Interestingly, bees release about 1 μg (10^{-6} g) of this compound when they sting to attract other bees to join the attack. How many molecules of isopentyl acetate are released in a typical bee sting? How many atom of carbon are present?

Molar Mass

Example 3.4 - SOLUTION

Molar mass of $C_7H_{14}O_2$

7 mol C x 12.011 g/mol	= 84.077 g C
14 mol H x 1.0079 g/mol	= 14.111 g H
2 mol O x 15.999 g/mol	= 31.998 g O
Mass of 1 mol of $C_7H_{14}O_2$	= 130.186 g

Thus 1 mole of isopentyl acetate (6.022×10^{23} molecules) has a mass of 130.186 g. to find the number of molecules released in a sting, we must first determine the number of moles of isopentyl acetate in 1×10^{-6} g:

1×10^{-6}g $C_7H_{14}O_2$ x 1mol $C_7H_{14}O_2$/130.186g $C_7H_{14}O_2$ = 8×10^{-9} mol $C_7H_{14}O_2$

Since 1 mol is 6.022×10^{23} units, we can determine the number of molecules:

8×10^{-9}mol $C_7H_{14}O_2$ x 6.022×10^{23} molecules / 1mol $C_7H_{14}O_2$ = 5×10^{15} molecules

Molar Mass

Example 3.4 – SOLUTION

To determine the number of carbon atoms present, we must multiply the number of molecules by 7 (each molecule of isopentyl acetate contains seven carbon atoms):

5×10^{15} molecules x 7 carbon atoms/molecule = 4×10^{16} carbon atoms

Percent Composition

Percent composition:

percentage by mass contributed by each element in a substance shows how many grams of each element exist in 100 g of a compound

Mass percent of an element

$$\boxed{\text{Mass\%} = \frac{\text{mass of element in compound}}{\text{mass of compound}} \times 100\%}$$

Percent Composition

Consider ethanol, which has the formula C_2H_5OH. The mass of each element present and the molar mass are obtained through the following procedure:

Mass of C = 2 mol x 12.011 g/mol = 24.022 g
Mass of H = 6 mol x 1.008 g/mol = 6.048 g
Mass of O = 1 mol x 15.999 g/mol = 15.999 g
Mass of 1 mol of C_2H_5OH = 46.069 g

Mass % of C = 24.022g/46.069g x 100% = 52.144%
Mass % of H = 6.048g/46.069g x 100% = 13.13%
Mass % of O = 15.999g/46.069g x 100% = 34.728%

HOMEWORK

Chap.2: 39, 41, 49, 53, 54, 63
Chap.3: 23, 37, 47

39

Would you expect each of the following atoms to gain or lose electrons when forming ions? What ion is the most likely in each case?

a. Ra	**b. In**
c. P	**d. Te**
e. Br	**f. Rb**

41

What is the symbol for an ion with 63 protons, 60 electrons, and 88 neutrons? If an ion contain 50 protons, 68 neutrons, and 48 electrons, what is its symbol?

Name the following compounds. Assume the potential acids are dissolved in water.

a. $HC_2H_3O_2$

b. NH_4NO_2

c. Co_2S_3

d. ICl

e. $Pb_3(PO_4)_2$

f. $KClO_3$

g. H_2SO_4

h. Sr_3N_2

i. $Al_2(SO_3)_3$

j. SnO_2

k. Na_2CrO_4

l. $HClO$

49

Name each of the following compounds.

a. CuI
b. CuI_2
c. CoI_2
d. Na_2CO_3
e. $NaHCO_3$
f. S_4N_4
g. $SeBr_4$
h. $NaOCl$
i. $BaCrO_4$
j. NH_4NO_3

53

Write formulas for the following compounds.

a. Sulfur dioxide
b. Sulfur trioxide
c. Sodium sulfite
d. Potassium hydrogen sulfite
e. Lithium nitride
f. Chromium (III) carbonate
g. Chromium (II) acetate
h. Tin (IV) fluoride
i. Ammonium hydrogen sulfate
j. Ammonium hydrogen phosphate
k. Potassium perchlorate
l. Sodium hydride
m. Hypobromous acid
n. Hydrobromic acid

54

Each of the following compounds is incorrectly named. What is wrong with each name, and what is the correct name for each compound?

a. **$FeCl_3$, iron chloride**
b. **NO_2, nitrogen (IV) oxide**
c. **CaO, calcium (II) oxide**
d. **Al_2S_3, dialuminum trisulfide**
e. **$Mg(C_2H_3O_2)_2$, manganese diacetate**
f. **$FePO_4$, iron (II) phosphide**
g. **P_2S_5, phosphorus sulfide**
h. **Na_2O_2, sodium oxide**
i. **HNO_3, nitrate acid**
j. **H_2S, sulfuric acid**

63

Elements in the same family often form oxyanions of the same general formula. The anions are named in a similar fashion. What are the names of the oxyanions of selenium and tellurium: SeO_4^{2-}, SeO_3^{2-}, TeO_4^{2-},TeO_3^{2-}?

23

The element rhenium (Re) has two naturally occurring isotopes, ^{185}Re and ^{187}Re, with an average atomic mass of 186.207 amu. Rhenium is 62.60% ^{187}Re, and the atomic mass of ^{187}Re is 186.956 amu. Calculate the mass of ^{185}Re.

Chloral hydrate ($C_2H_3Cl_3O_2$) is a drug formerly used as a sedative and hypnotic. It is the compound used to make "Mickey Finns" in detective stories.

a.Calculate the molar mass of chloral hydrate.

b. How many moles of $C_2H_3Cl_3O_2$ molecules are in 500.0 g of chloral hydrate?

c. What is the mass in grams of $2.0x10^{-2}$ mol chloral hydrate?

d. How many chlorine atoms are in 5.0g chloral hydrate?

e. What mass of chloral hydrate would contain 1.0 g Cl?

f. What is the mass of exactly 500 molecules of chloral hydrate?

47

Fungal laccase, a blue protein found in wood-rotting fungi, is 0.390% Cu by mass. If a fungal laccase molecule contains four copper atoms, what is the molar mass of fungal laccase?