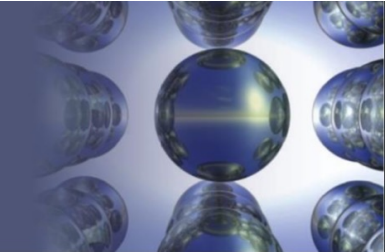


Chapter 2

Atoms, Molecules, and Ions

Section 2.1

The Early History of Chemistry

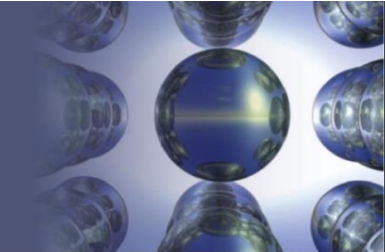


Early History of Chemistry

- Greeks were the first to attempt to explain why chemical changes occur.
- Alchemy dominated for 2000 years.
 - Several elements discovered.
 - Mineral acids prepared.
- Robert Boyle was the first “chemist”.
 - Performed quantitative experiments.
 - Developed first experimental definition of an element.

Section 2.2

Fundamental Chemical Laws



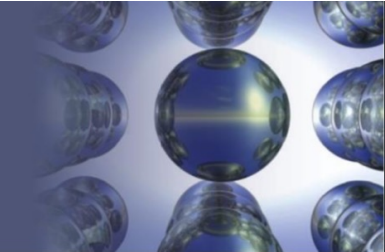
Three Important Laws

- Law of conservation of mass (Lavoisier):
 - Mass is neither created nor destroyed in a chemical reaction.

- Law of definite proportion (Proust):
 - A given compound always contains exactly the same proportion of elements by mass.

Section 2.2

Fundamental Chemical Laws

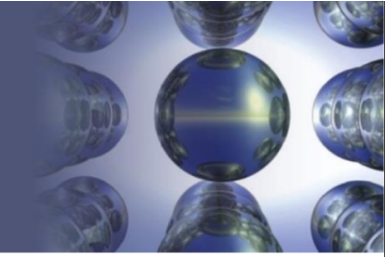


Three Important Laws (continued)

- Law of multiple proportions (Dalton):
 - When two elements form a series of compounds, the ratios of the masses of the second element that combine with 1 gram of the first element can always be reduced to small whole numbers.

Section 2.3

Dalton's Atomic Theory

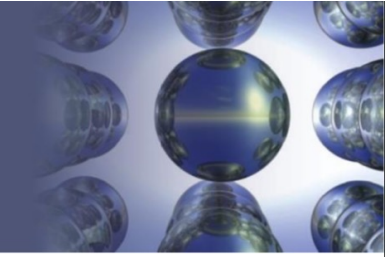


Dalton's Atomic Theory (1808)

- *Each element is made up of tiny particles called atoms.*

Section 2.3

Dalton's Atomic Theory

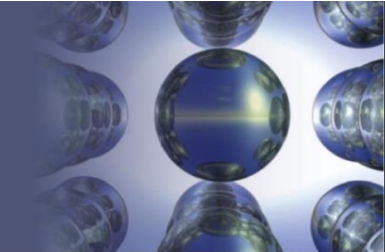


Dalton's Atomic Theory (continued)

- *The atoms of a given element are identical;* the atoms of different elements are different in some fundamental way or ways.

Section 2.3

Dalton's Atomic Theory

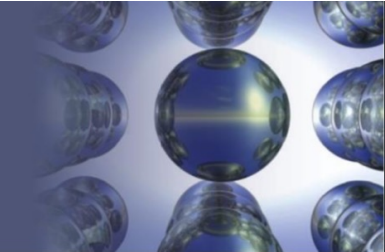


Dalton's Atomic Theory (continued)

- *Chemical compounds are formed when atoms of different elements combine with each other. A given compound always has the same relative numbers and types of atoms.*

Section 2.3

Dalton's Atomic Theory

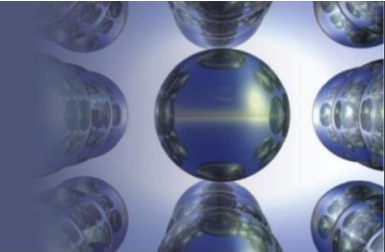


Dalton's Atomic Theory (continued)

- *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.

Section 2.3

Dalton's Atomic Theory

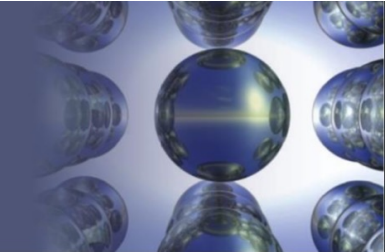


Gay-Lussac and Avogadro (1809—1811)

- Gay—Lussac
 - Measured (under same conditions of T and P) the volumes of gases that reacted with each other.
- Avogadro's Hypothesis
 - At the same T and P, equal volumes of different gases contain the same number of particles.
 - Volume of a gas is determined by the number, not the size, of molecules.

Section 2.3

Dalton's Atomic Theory

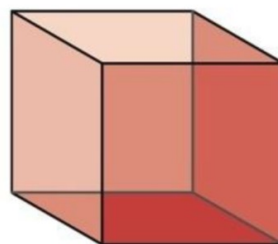


Representing Gay—Lussac's Results

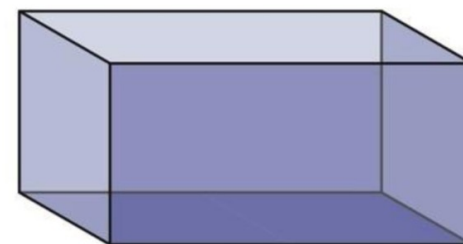


2 volumes hydrogen

+



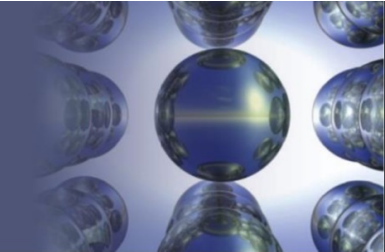
combines with 1 volume oxygen



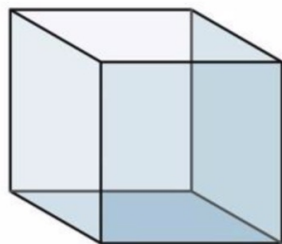
to form 2 volumes gaseous water

Section 2.3

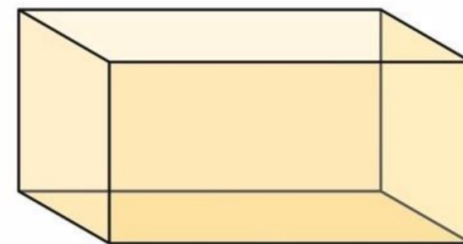
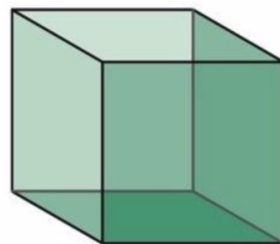
Dalton's Atomic Theory



Representing Gay—Lussac's Results



+

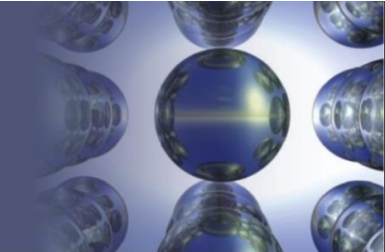


1 volume hydrogen

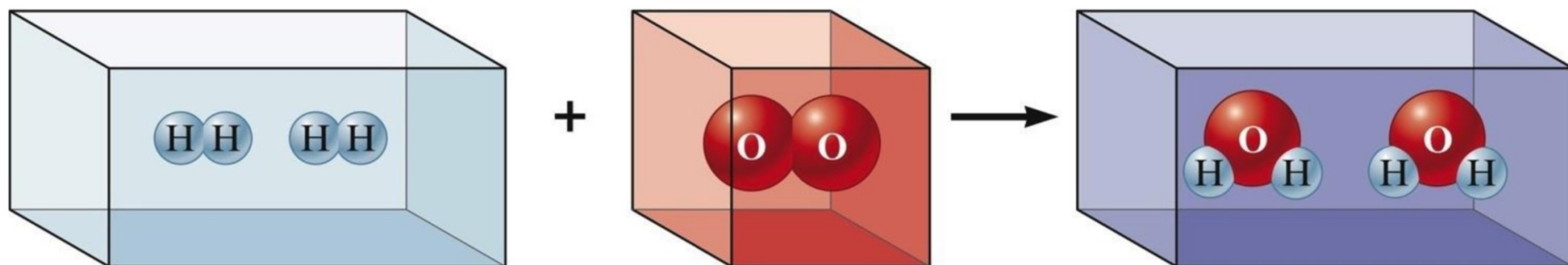
combines with 1 volume chlorine to form 2 volumes hydrogen chloride

Section 2.3

Dalton's Atomic Theory

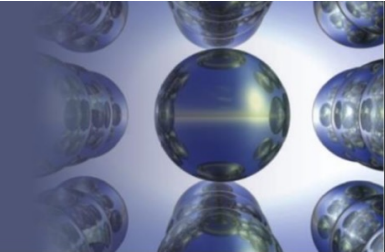


Representing Gay—Lussac's Results

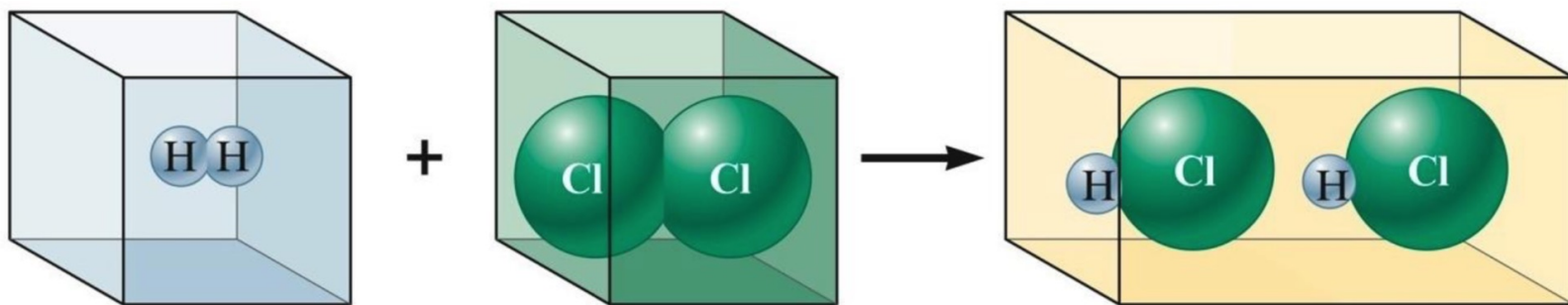


Section 2.3

Dalton's Atomic Theory

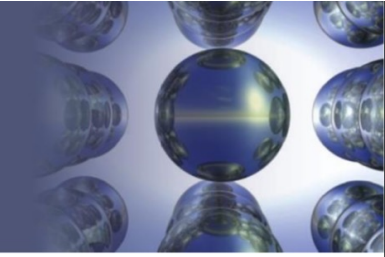


Representing Gay—Lussac's Results



Section 2.4

Early Experiments to Characterize the Atom

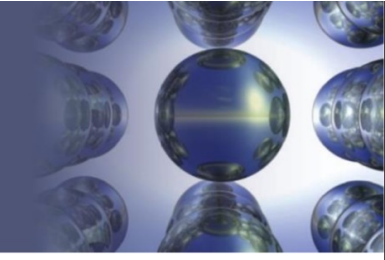


J. J. Thomson (1898—1903)

- Postulated the existence of negatively charged particles, that we now call electrons, using cathode-ray tubes.
- Determined the charge-to-mass ratio of an electron.
- The atom must also contain positive particles that balance exactly the negative charge carried by electrons.

Section 2.4

Early Experiments to Characterize the Atom

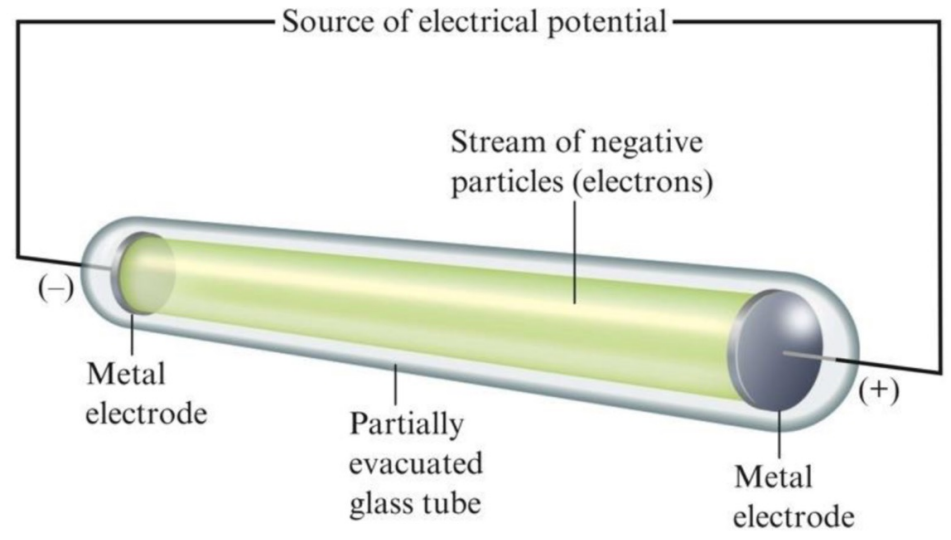


Cathode-Ray Tube



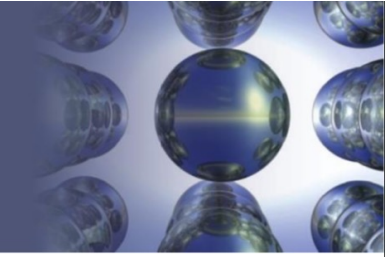
Richard Megna/Fundamental Photographs
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Section 2.4

Early Experiments to Characterize the Atom

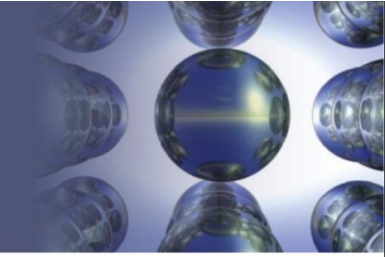


Robert Millikan (1909)

- Performed experiments involving charged oil drops.
- Determined the magnitude of the charge on a single electron.
- Calculated the mass of the electron
 - $(9.11 \times 10^{-31} \text{ kg})$.

Section 2.4

Early Experiments to Characterize the Atom

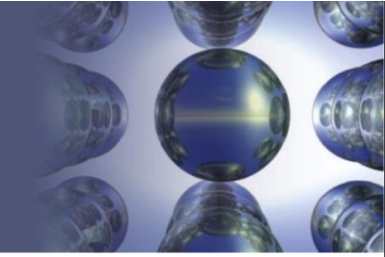


Henri Becquerel (1896)

- Discovered radioactivity by observing the spontaneous emission of radiation by uranium.
- Three types of radioactive emission exist:
 - Gamma rays (γ) – high energy light
 - Beta particles (β) – a high speed electron
 - Alpha particles (α) – a particle with a 2+ charge

Section 2.4

Early Experiments to Characterize the Atom

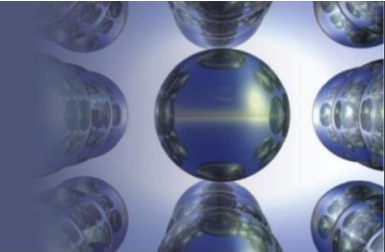


Ernest Rutherford (1911)

- Explained the nuclear atom.
- The atom has a dense center of positive charge called the nucleus.
- Electrons travel around the nucleus at a large distance relative to the nucleus.

Section 2.5

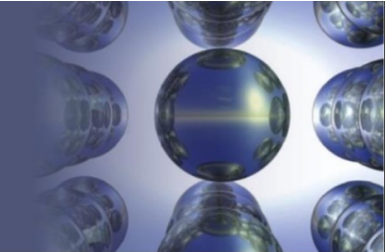
The Modern View of Atomic Structure: An Introduction



- The atom contains:
 - *Electrons* – found outside the nucleus; negatively charged.
 - *Protons* – found in the nucleus; positive charge equal in magnitude to the electron's negative charge.
 - *Neutrons* – found in the nucleus; no charge; virtually same mass as a proton.

Section 2.5

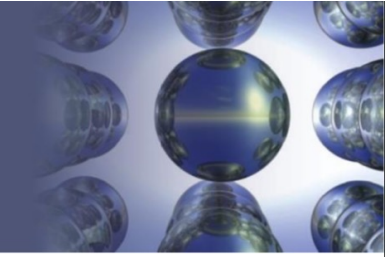
The Modern View of Atomic Structure: An Introduction



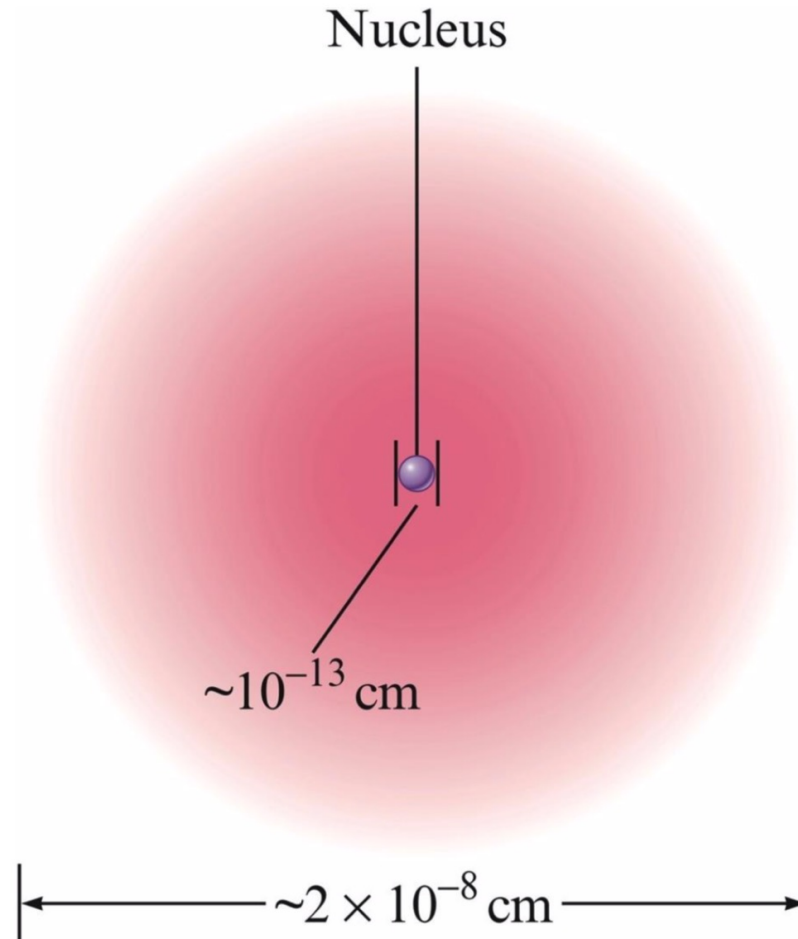
- The nucleus is:
 - Small compared with the overall size of the atom.
 - Extremely dense; accounts for almost all of the atom's mass.

Section 2.5

The Modern View of Atomic Structure: An Introduction

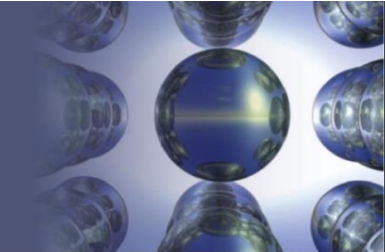


Nuclear Atom Viewed in Cross Section



Section 2.5

The Modern View of Atomic Structure: An Introduction



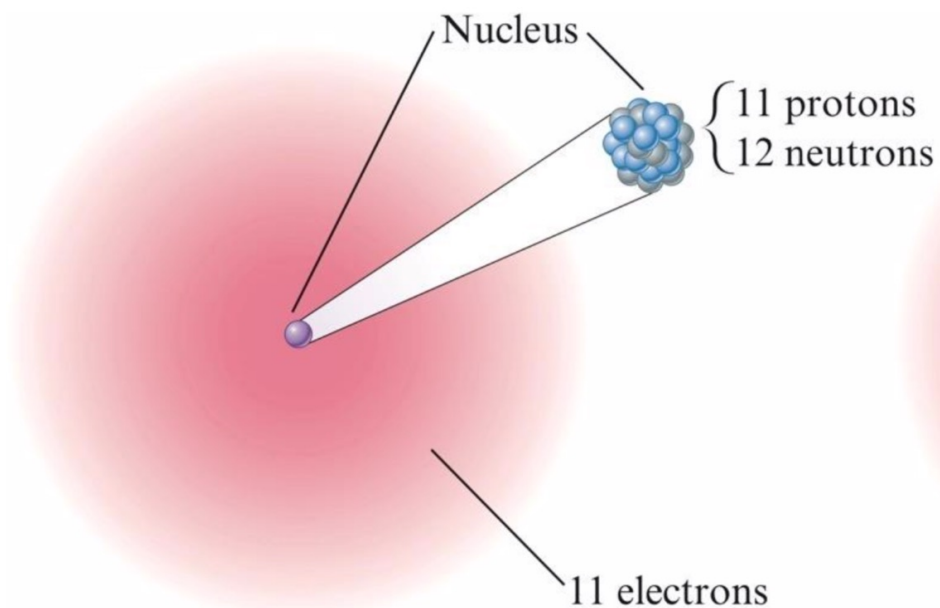
Isotopes

- Atoms with the same number of protons but different numbers of neutrons.
- Show almost identical chemical properties; chemistry of atom is due to its electrons.
- In nature most elements contain mixtures of isotopes.

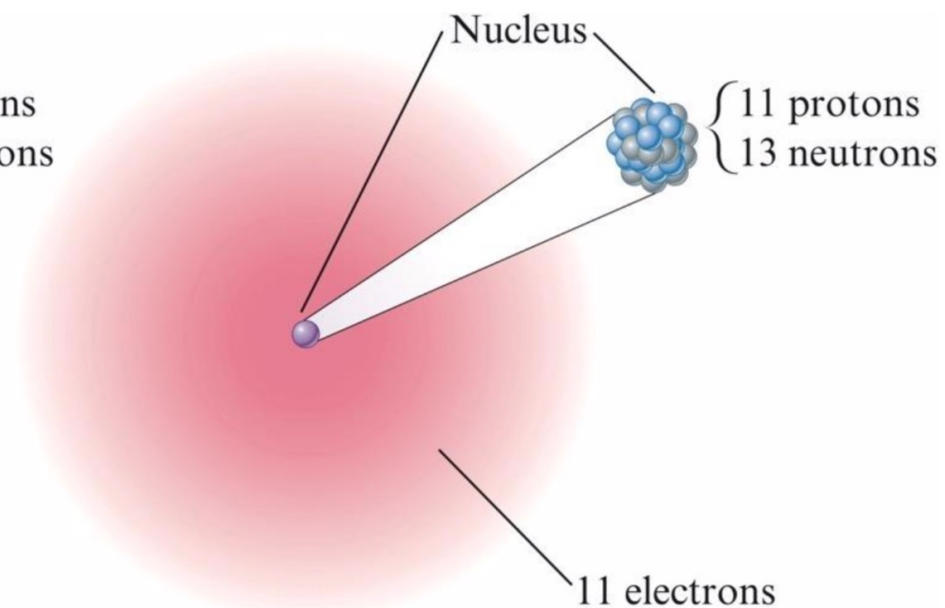
Section 2.5

The Modern View of Atomic Structure: An Introduction

Two Isotopes of Sodium

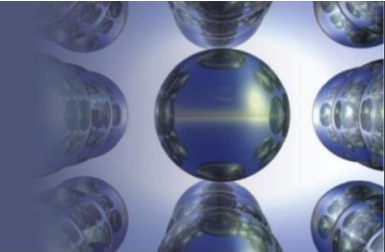


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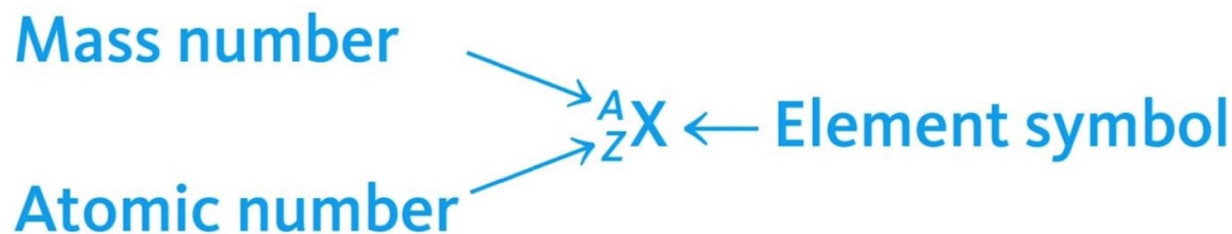


Section 2.5

The Modern View of Atomic Structure: An Introduction

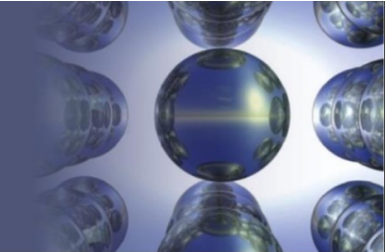


- Isotopes are identified by:
 - Atomic Number (Z) – number of protons
 - Mass Number (A) – number of protons plus number of neutrons



Section 2.5

The Modern View of Atomic Structure: An Introduction



EXERCISE!

A certain isotope X contains 23 protons and 28 neutrons.

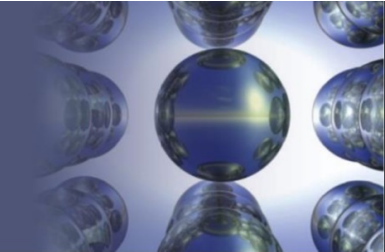
- What is the **mass number** of this isotope?
- Identify the **element**.

Mass Number = 51

Vanadium

Section 2.6

Molecules and Ions

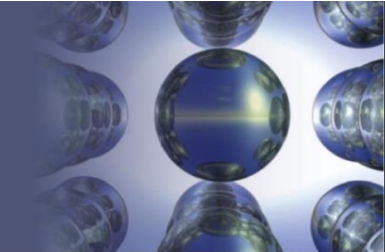


Chemical Bonds

- Covalent Bonds
 - Bonds form between atoms by sharing electrons.
 - Resulting collection of atoms is called a molecule.

Section 2.6

Molecules and Ions



Chemical Bonds

- Ionic Bonds
 - Bonds form due to force of attraction between oppositely charged ions.
 - *Ion* – atom or group of atoms that has a net positive or negative charge.
 - *Cation* – positive ion; lost electron(s).
 - *Anion* – negative ion; gained electron(s).

Section 2.6

Molecules and Ions

EXERCISE!

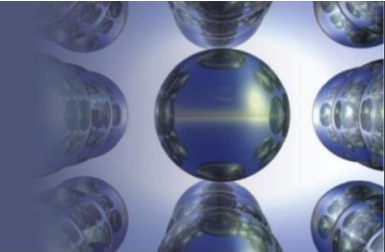
A certain isotope X^+ contains 54 electrons and 78 neutrons.

- What is the **mass number** of this isotope?

133

Section 2.6

Molecules and Ions



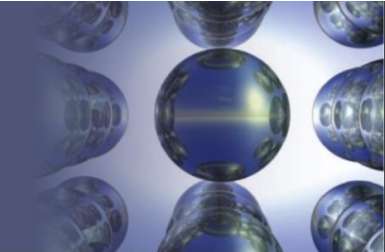
CONCEPT CHECK!

Which of the following statements regarding Dalton's atomic theory are still believed to be **true**?

- I. Elements are made of tiny particles called atoms.
- II. All atoms of a given element are identical.
- III. A given compound always has the same relative numbers and types of atoms.
- IV. Atoms are indestructible.

Section 2.6

Molecules and Ions



CONCEPT CHECK!

Which of the following statements regarding Dalton's atomic theory are still believed to be **true**?

I. Elements are made of tiny particles called atoms.

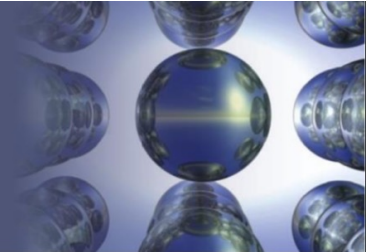
II. All atoms of a given element are identical.

III. A given compound always has the same relative numbers and types of atoms.

IV. Atoms are indestructible.

Section 2.7

An Introduction to the Periodic Table



The Periodic Table

- *Metals vs. Nonmetals*
- *Groups or Families* – elements in the same vertical columns; have similar chemical properties
- *Periods* – horizontal rows of elements

Section 2.7

An Introduction to the Periodic Table

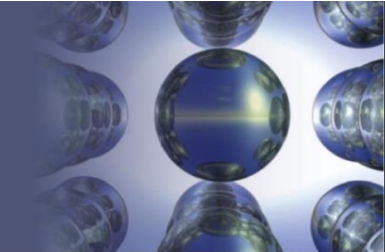
The Periodic Table

		Alkaline earth metals																	Noble gases		
		1 1A											2 2A								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	Transition metals										13 Al	14 Si	15 P	16 S	17 Cl	18 Ar		
Alkali metals	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 Fl	115 Uup	116 Lv	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

Section 2.7

An Introduction to the Periodic Table



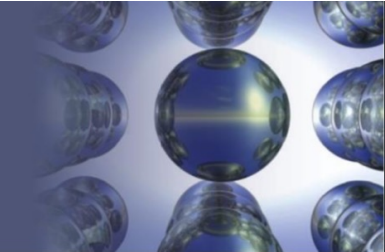
Groups or Families

- Table of common charges formed when creating ionic compounds.

Group or Family	Charge
Alkali Metals (1A)	1+
Alkaline Earth Metals (2A)	2+
Halogens (7A)	1-
Noble Gases (8A)	0

Section 2.8

Naming Simple Compounds

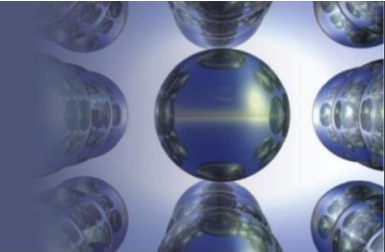


Naming Compounds

- Binary Compounds
 - Composed of two elements
 - Ionic and covalent compounds included
- Binary Ionic Compounds
 - Metal—nonmetal
- Binary Covalent Compounds
 - Nonmetal—nonmetal

Section 2.8

Naming Simple Compounds



Binary Ionic Compounds (Type I)

1. The cation is always named first and the anion second.
2. A monatomic cation takes its name from the name of the parent element.
3. A monatomic anion is named by taking the root of the element name and adding *-ide*.

Section 2.8

Naming Simple Compounds

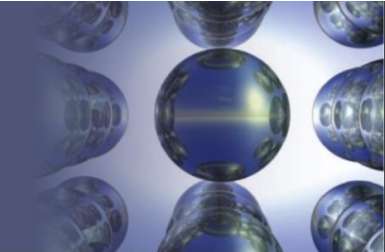
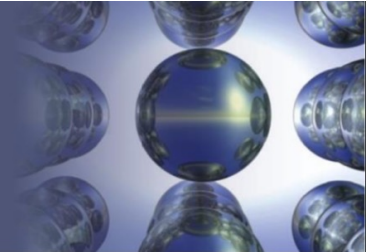


Table- Common Monatomic Cations and Anions

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		

Section 2.8

Naming Simple Compounds



Binary Ionic Compounds (Type I)

- Examples:



Potassium chloride



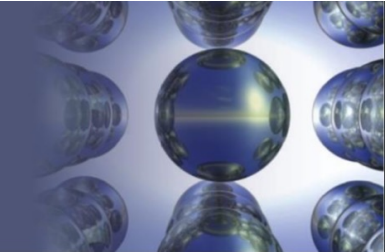
Magnesium bromide



Calcium oxide

Section 2.8

Naming Simple Compounds



Binary Ionic Compounds (Type II)

- Metals in these compounds form more than one type of positive ion.
- Charge on the metal ion must be specified.
- Roman numeral indicates the charge of the metal cation.
- Transition metal cations usually require a Roman numeral.
- Elements that form only one cation do not need to be identified by a roman numeral.

Section 2.8

Naming Simple Compounds

Binary Ionic Compounds (Type II)

- Examples:

CuBr Copper(I) bromide

FeS Iron(II) sulfide

PbO_2 Lead(IV) oxide

Table- Common Type II Cations

Ion	Systematic Name
Fe^{3+}	Iron(III)
Fe^{2+}	Iron(II)
Cu^{2+}	Copper(II)
Cu^{+}	Copper(I)
Co^{3+}	Cobalt(III)
Co^{2+}	Cobalt(II)
Sn^{4+}	Tin(IV)
Sn^{2+}	Tin(II)
Pb^{4+}	Lead(IV)
Pb^{2+}	Lead(II)
Hg^{2+}	Mercury(II)
Hg_2^{2+}	Mercury(I)
Ag^{+}	Silver [†]
Zn^{2+}	Zinc [†]
Cd^{2+}	Cadmium [†]

Section 2.8

Naming Simple Compounds

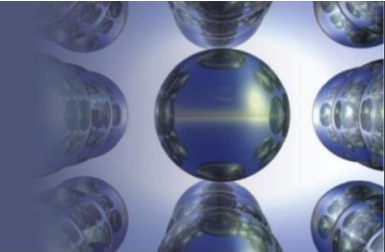
1A	2A										3A	4A	5A	6A	7A	8A
Li ⁺													N ³⁻	O ²⁻	F ⁻	
Na ⁺	Mg ²⁺										Al ³⁺			S ²⁻	Cl ⁻	
K ⁺	Ca ²⁺				Cr ²⁺ Cr ³⁺	Mn ²⁺ Mn ³⁺	Fe ²⁺ Fe ³⁺	Co ²⁺ Co ³⁺		Cu ⁺ Cu ²⁺	Zn ²⁺	Ga ³⁺			Br ⁻	
Rb ⁺	Sr ²⁺									Ag ⁺	Cd ²⁺		Sn ²⁺ Sn ⁴⁺		I ⁻	
Cs ⁺	Ba ²⁺										Hg ₂ ²⁺ Hg ²⁺		Pb ²⁺ Pb ⁴⁺			

Common Type I cations
 Common Type II cations
 Common monatomic anions

Figure- The common cations and anions.

Section 2.8

Naming Simple Compounds



Polyatomic Ions

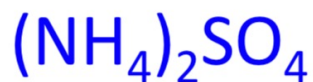
- Must be memorized (see Table 2.5 on pg. 65 in text).
- Examples of compounds containing polyatomic ions:



Sodium hydroxide



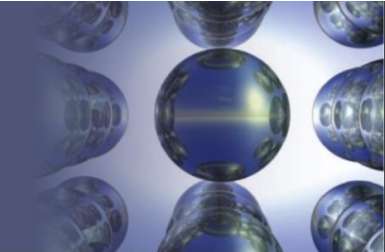
Magnesium nitrate



Ammonium sulfate

Section 2.8

Naming Simple Compounds

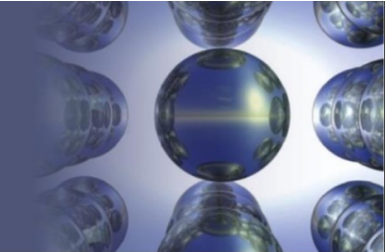


Polyatomic Ions

Ion	Name	Ion	Name
Hg_2^{2+}	Mercury(I)	NCS^- or SCN^-	Thiocyanate
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	ClO^- or OCl^-	Hypochlorite
HSO_4^-	Hydrogen sulfate (bisulfate is a widely used common name)	ClO_2^-	Chlorite
		ClO_3^-	Chlorate
OH^-	Hydroxide	ClO_4^-	Perchlorate
CN^-	Cyanide	$\text{C}_2\text{H}_3\text{O}_2^-$	Acetate
PO_4^{3-}	Phosphate	MnO_4^-	Permanganate
HPO_4^{2-}	Hydrogen phosphate	$\text{Cr}_2\text{O}_7^{2-}$	Dichromate
H_2PO_4^-	Dihydrogen phosphate	CrO_4^{2-}	Chromate
		O_2^{2-}	Peroxide
		$\text{C}_2\text{O}_4^{2-}$	Oxalate
		$\text{S}_2\text{O}_3^{2-}$	Thiosulfate

Section 2.8

Naming Simple Compounds



Binary Covalent Compounds (Type III)

- Formed between two nonmetals.
1. The first element in the formula is named first, using the full element name.
 2. The second element is named as if it were an anion.
 3. Prefixes are used to denote the numbers of atoms present.
 4. The prefix *mono-* is never used for naming the first element.

Section 2.8

Naming Simple Compounds

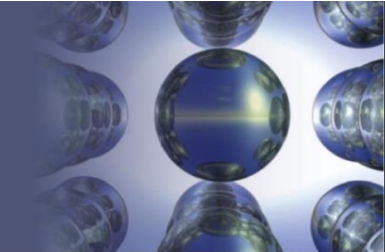
Prefixes Used to Indicate Number in Chemical Names

Table 2.6 | Prefixes Used to Indicate Number in Chemical Names

Prefix	Number Indicated
<i>mono-</i>	1
<i>di-</i>	2
<i>tri-</i>	3
<i>tetra-</i>	4
<i>penta-</i>	5
<i>hexa-</i>	6
<i>hepta-</i>	7
<i>octa-</i>	8
<i>nona-</i>	9
<i>deca-</i>	10

Section 2.8

Naming Simple Compounds



Binary Covalent Compounds (Type III)

- Examples:

CO_2 Carbon dioxide

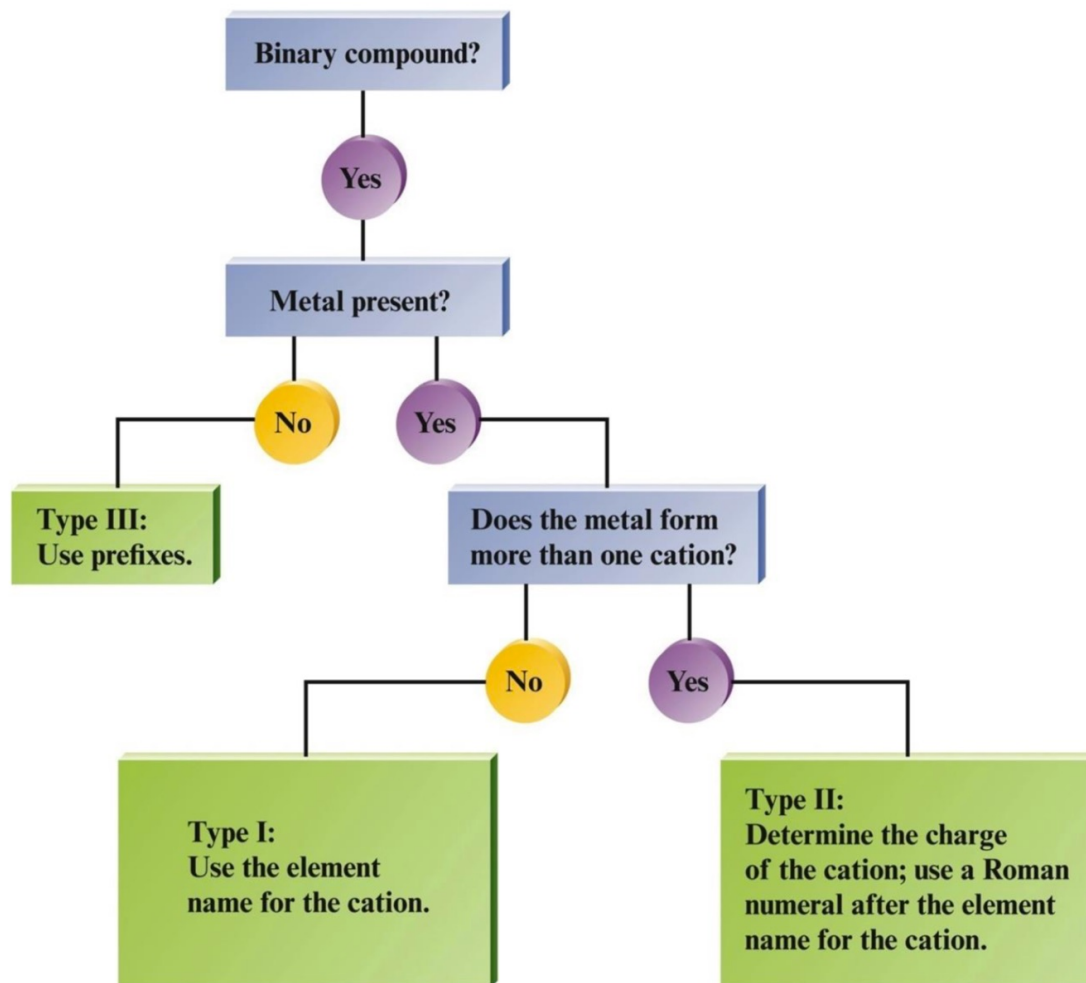
SF_6 Sulfur hexafluoride

N_2O_4 Dinitrogen tetroxide

Section 2.8

Naming Simple Compounds

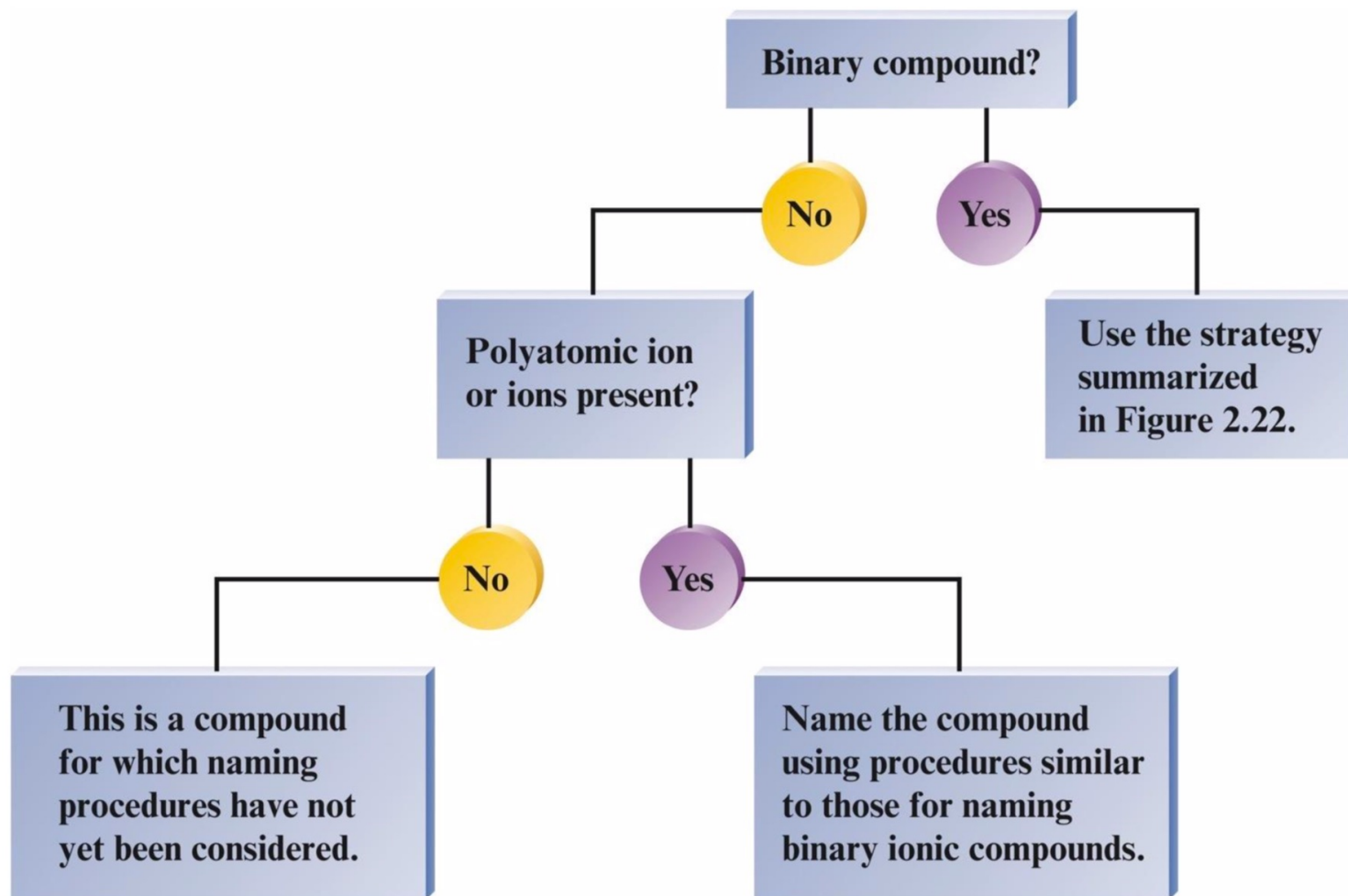
Flowchart for Naming Binary Compounds



Section 2.8

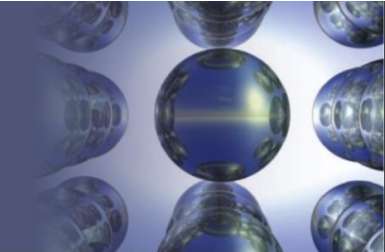
Naming Simple Compounds

Overall Strategy for Naming Chemical Compounds



Section 2.8

Naming Simple Compounds

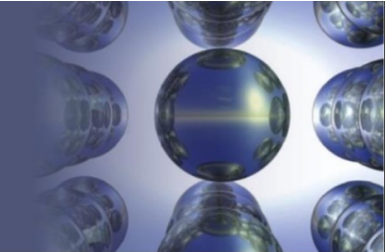


Acids

- Acids can be recognized by the hydrogen that appears first in the formula—HCl.
- Molecule with one or more H^+ ions attached to an anion.

Section 2.8

Naming Simple Compounds



Acids

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

- Examples:

HCl

Hydrochloric acid

HCN

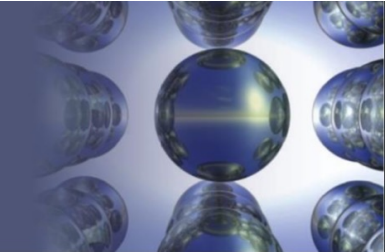
Hydrocyanic acid

H₂S

Hydrosulfuric acid

Section 2.8

Naming Simple Compounds



Acids

- If the anion *does* contain oxygen:
 - The suffix *-ic* is added to the root name if the anion name ends in *-ate*.

- Examples:

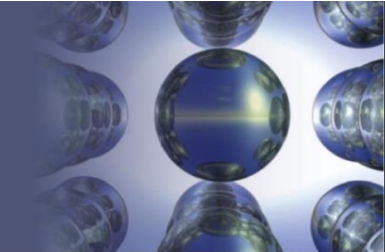
HNO_3 Nitric acid

H_2SO_4 Sulfuric acid

$\text{HC}_2\text{H}_3\text{O}_2$ Acetic acid

Section 2.8

Naming Simple Compounds



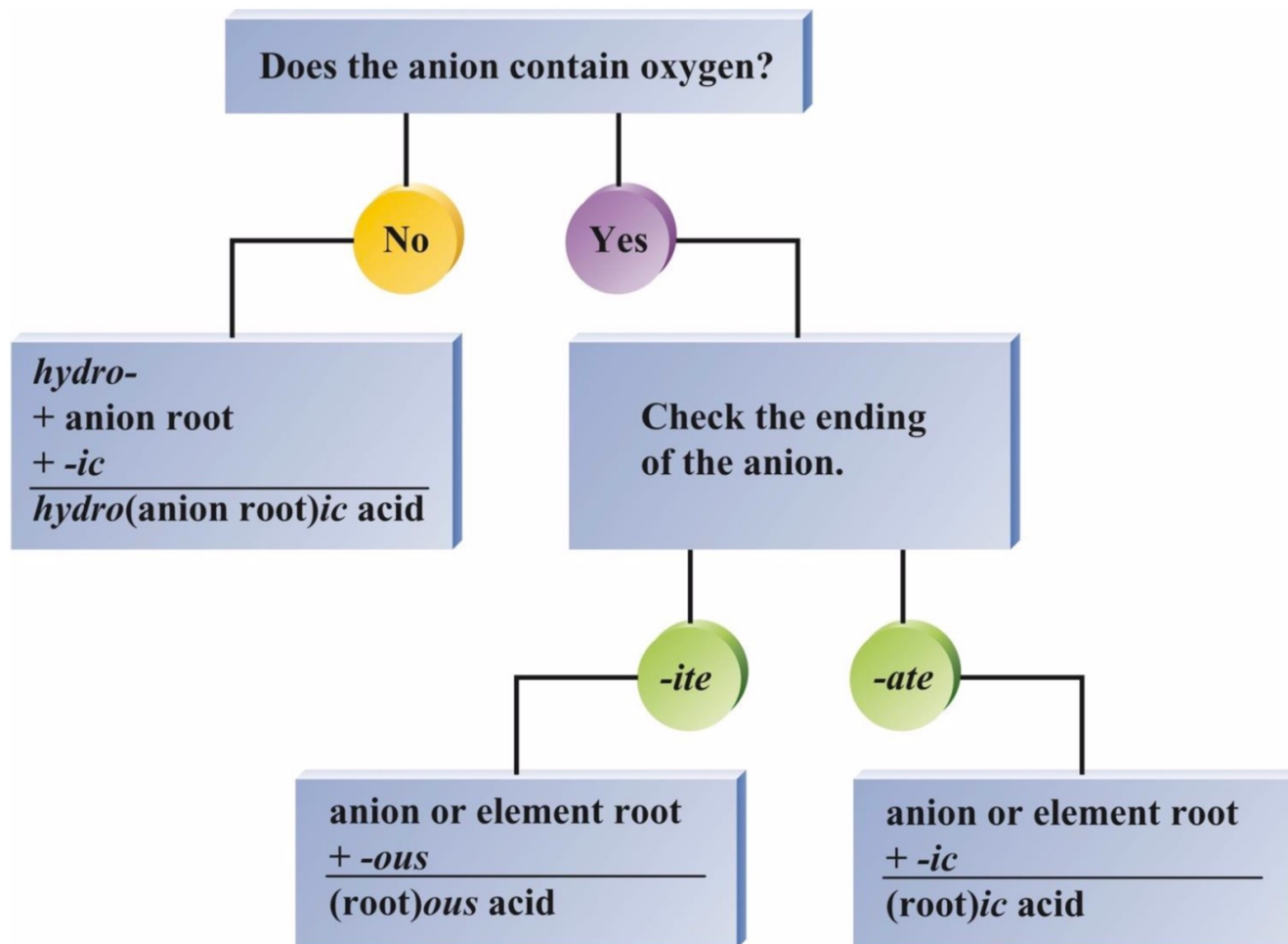
Acids

- If the anion *does* contain oxygen:
 - The suffix *-ous* is added to the root name if the anion name ends in *-ite*.
- Examples:
 - HNO_2 Nitrous acid
 - H_2SO_3 Sulfurous acid
 - HClO_2 Chlorous acid

Section 2.8

Naming Simple Compounds

Flowchart for Naming Acids



Section 2.8

Naming Simple Compounds

EXERCISE!

Which of the following compounds is named **incorrectly**?

- | | |
|----------------------|-------------------------|
| a) KNO_3 | potassium nitrate |
| b) TiO_2 | titanium(II) oxide |
| c) Sn(OH)_4 | tin(IV) hydroxide |
| d) PBr_5 | phosphorus pentabromide |
| e) CaCrO_4 | calcium chromate |