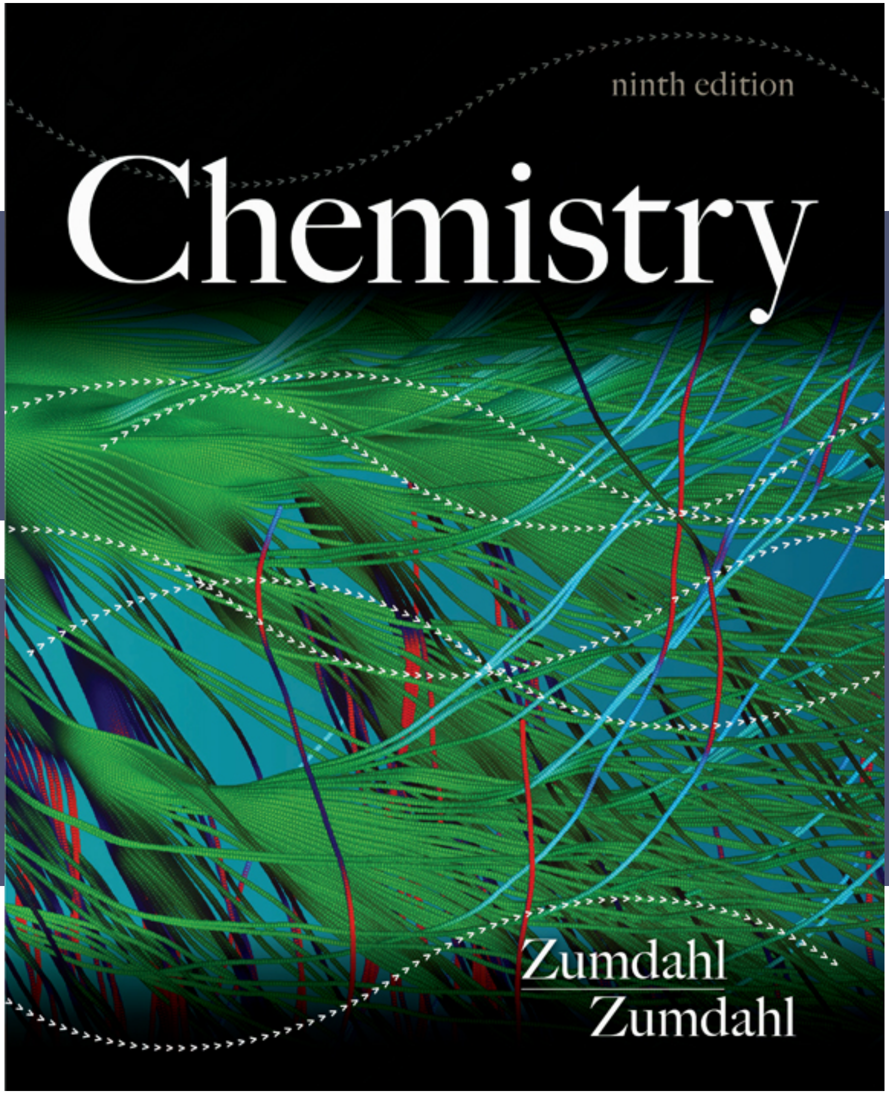


Chapter 2

Atoms, Molecules, and Ions

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.

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.

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.

Dalton’s Atomic Theory (1808)

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

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.

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.

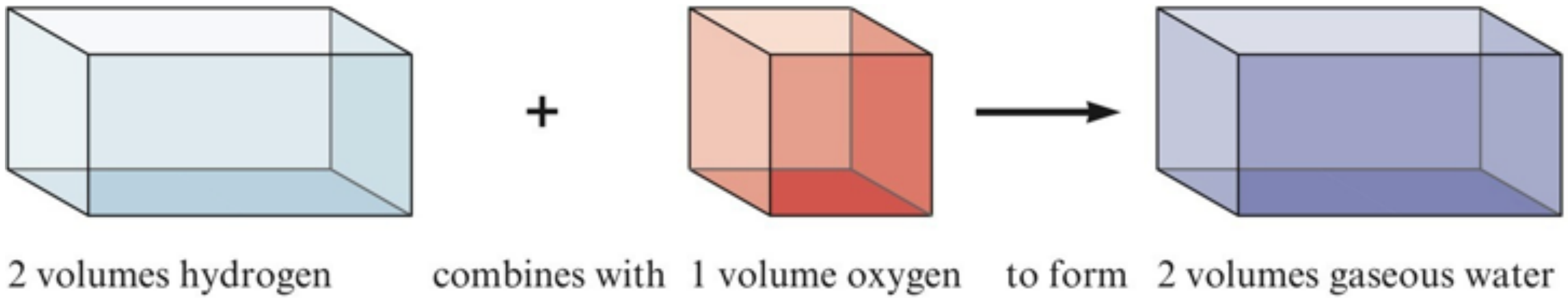
Dalton's Atomic Theory (continued)

- *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.

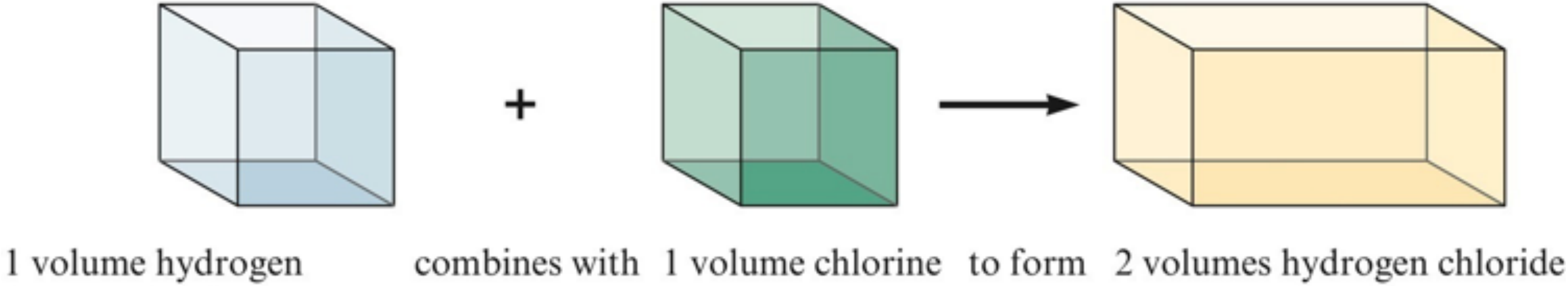
Gay-Lussac and Avogadro (1809—1811)

- 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.

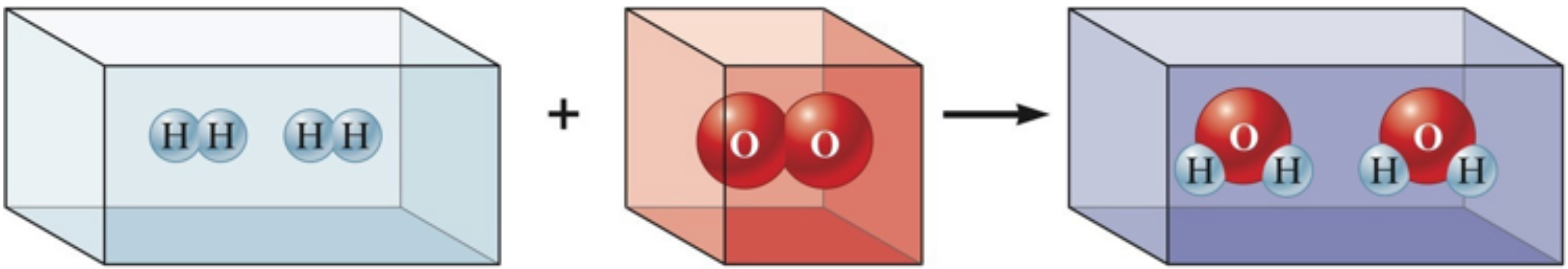
Representing Gay—Lussac's Results



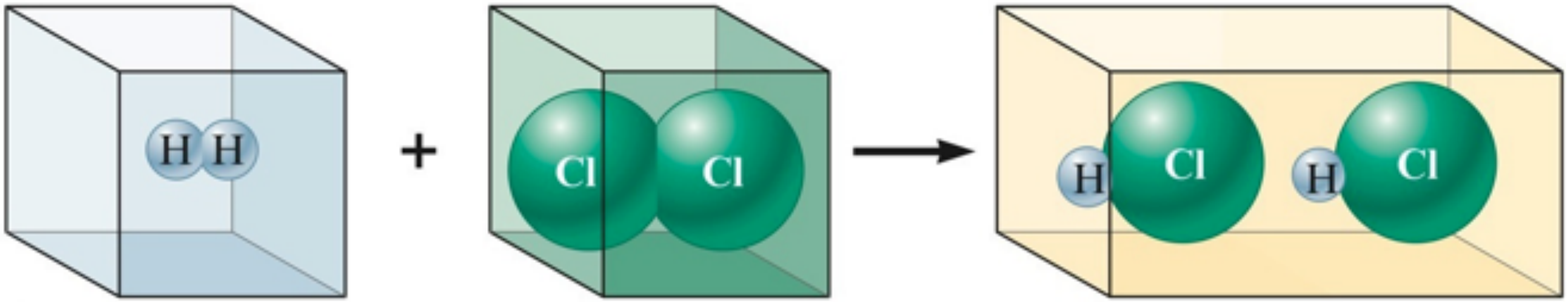
Representing Gay—Lussac' s Results



Representing Gay—Lussac' s Results



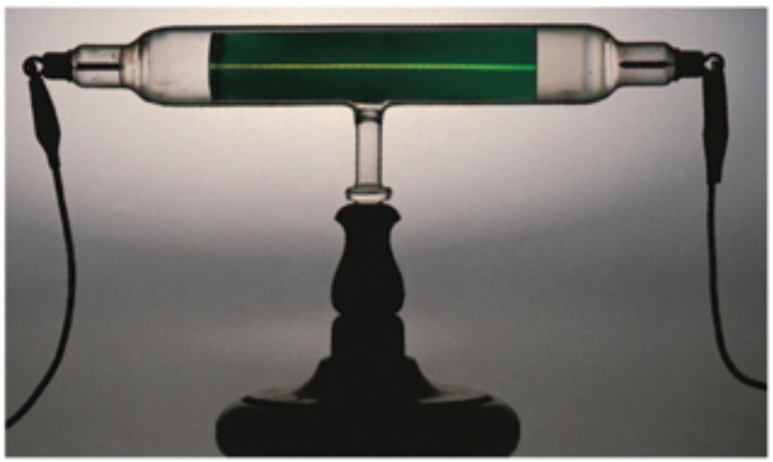
Representing Gay—Lussac’s Results



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.

Cathode-Ray Tube



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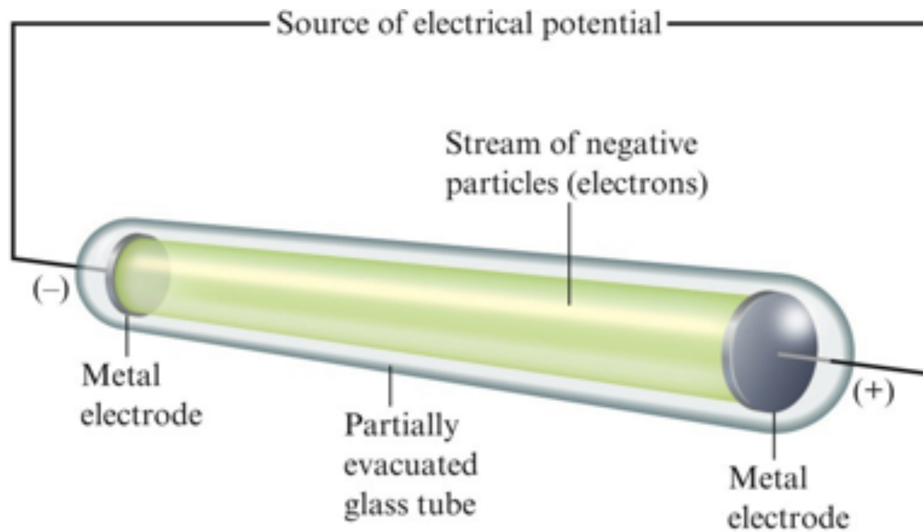
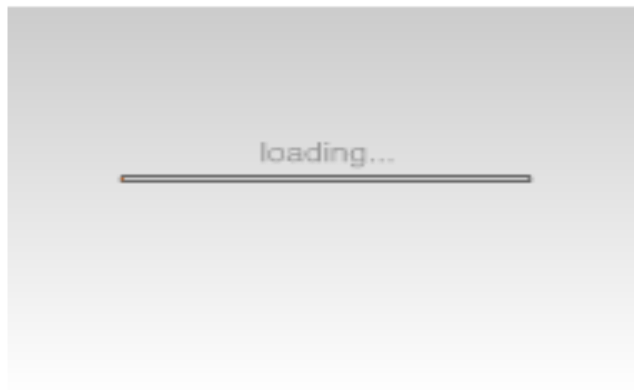


Robert Millikan (1909)

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

Millikan Oil Drop Experiment



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Henri Becquerel (1896)

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

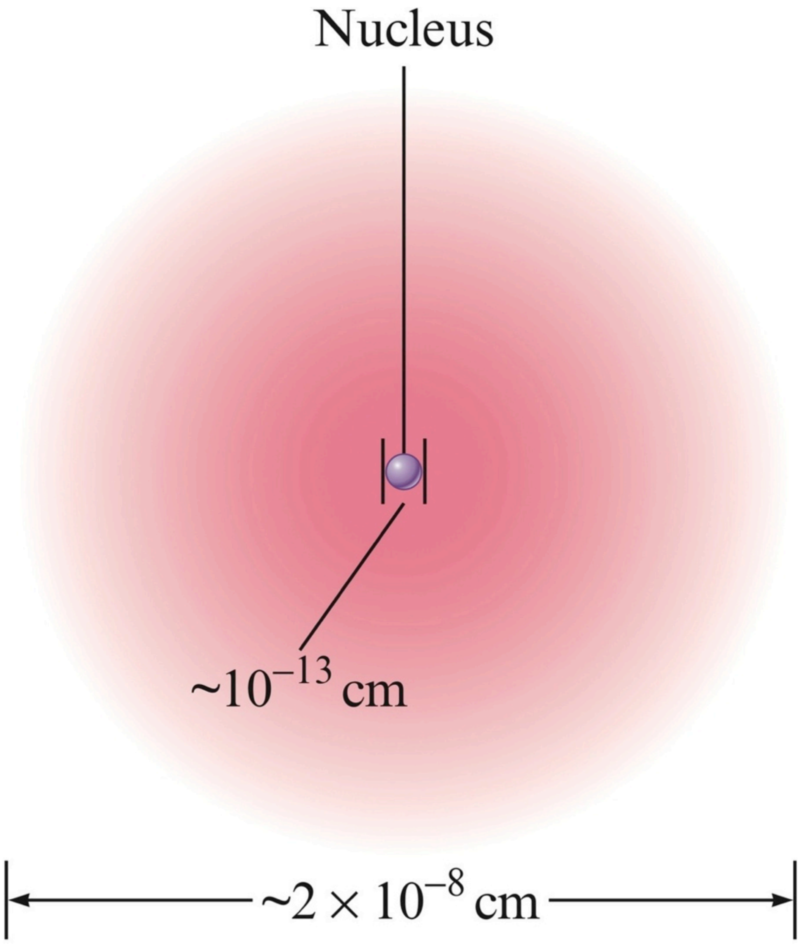
Ernest Rutherford (1911)

- 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.

- 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.

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

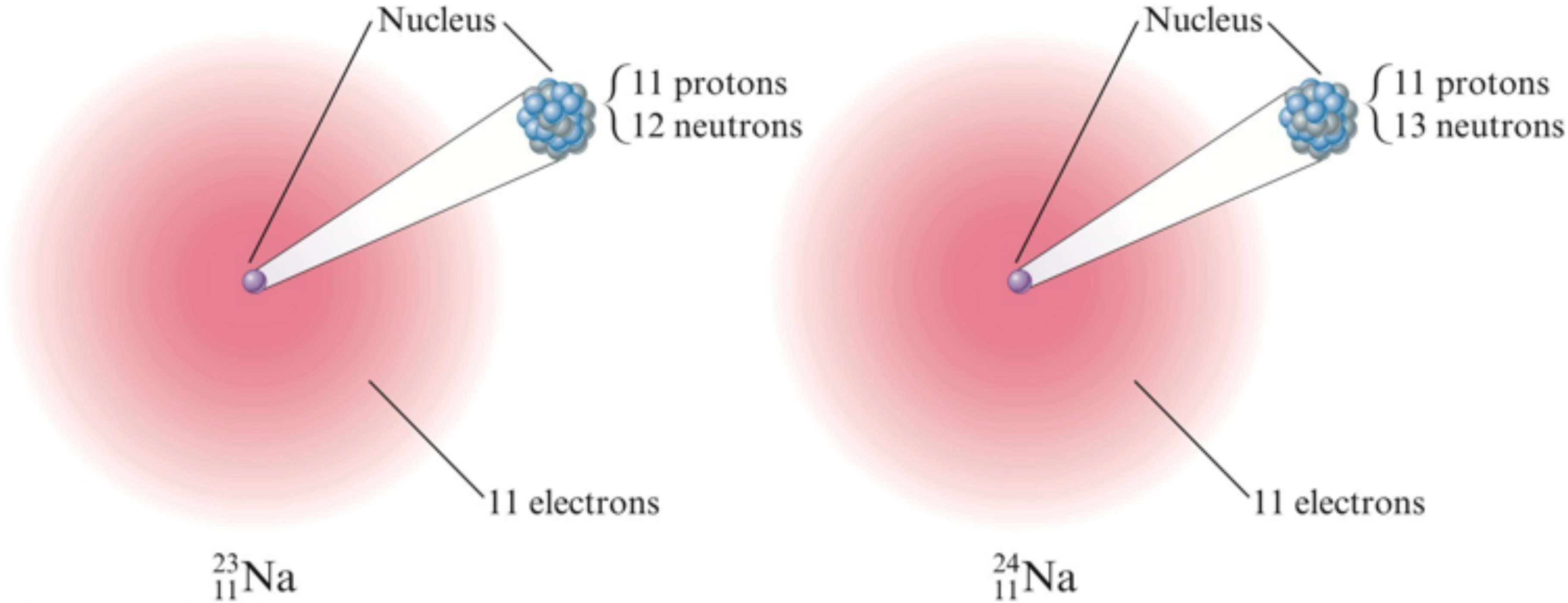
Nuclear Atom Viewed in Cross Section



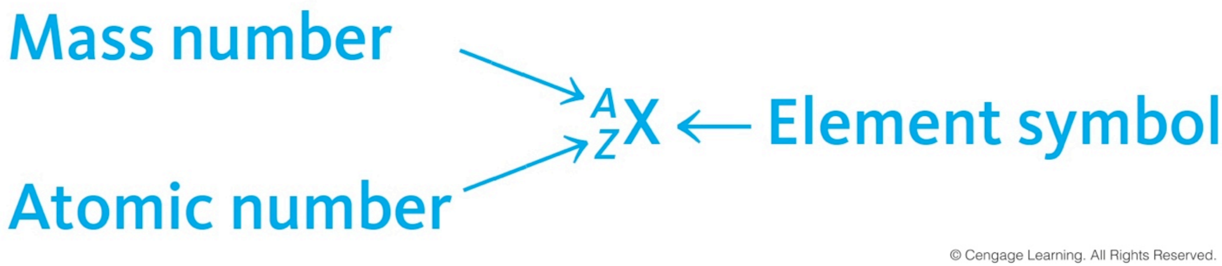
Isotopes

- 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.

Two Isotopes of Sodium



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



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

Chemical Bonds

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

Covalent Bonding

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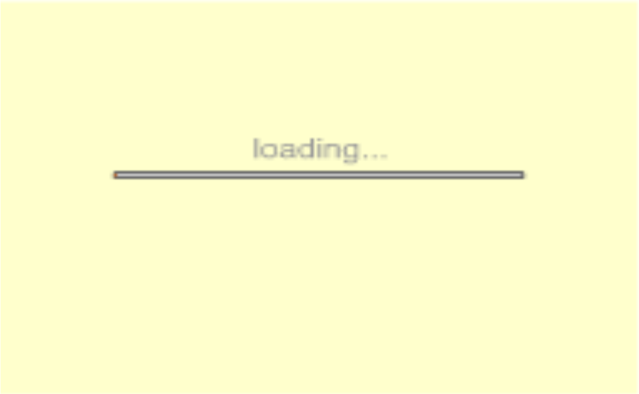
Chemical Bonds

- 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).

Molecular vs Ionic Compounds



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EXERCISE!

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

- What is the mass number of this isotope?

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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.

The Periodic Table

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

The Periodic Table

1 1A Alkali metals	2 2A Alkaline earth metals	3	4	5	6	7	8	9	10	11	12	13 3A	14 4A	15 5A	16 6A	17 7A Halogens	18 8A Noble gases
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 Fl	115 Uup	116 Lv	117 Uus	118 Uuo

Groups 3–12: Transition metals

*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

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

Naming Compounds

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

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*.

Binary Ionic Compounds (Type I)

- Examples:

KCl	Potassium chloride
$MgBr_2$	Magnesium bromide
CaO	Calcium oxide

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.

Binary Ionic Compounds (Type II)

- Examples:

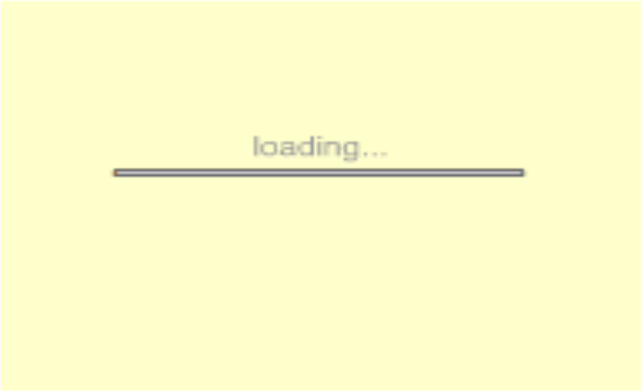
CuBr	Copper(I) bromide
FeS	Iron(II) sulfide
PbO_2	Lead(IV) oxide

Polyatomic Ions

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

$NaOH$	Sodium hydroxide
$Mg(NO_3)_2$	Magnesium nitrate
$(NH_4)_2SO_4$	Ammonium sulfate

Formation of Ionic Compounds



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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.

Prefixes Used to Indicate Number in Chemical Names

Table 2.6 Prefixes Used to Indicate Number in Chemical Names

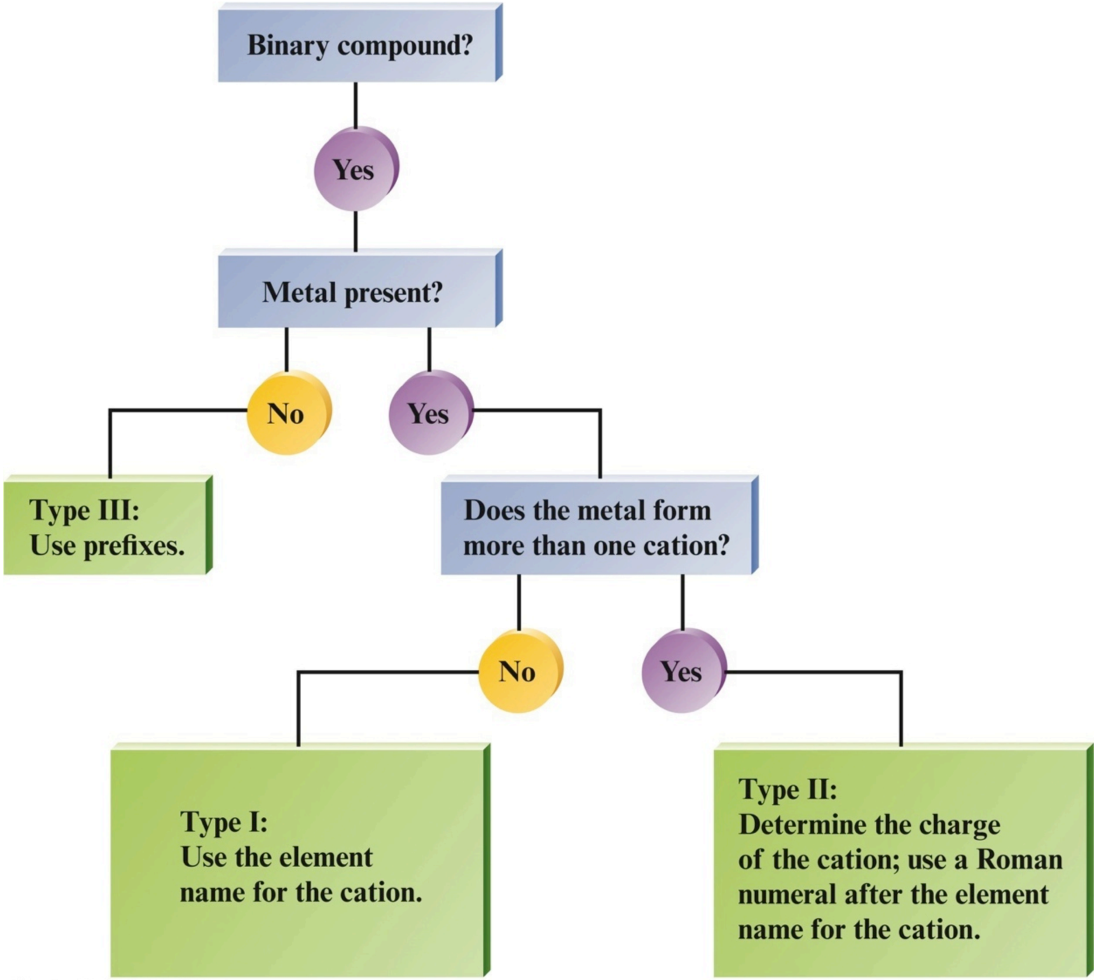
Prefix	Number Indicated
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Binary Covalent Compounds (Type III)

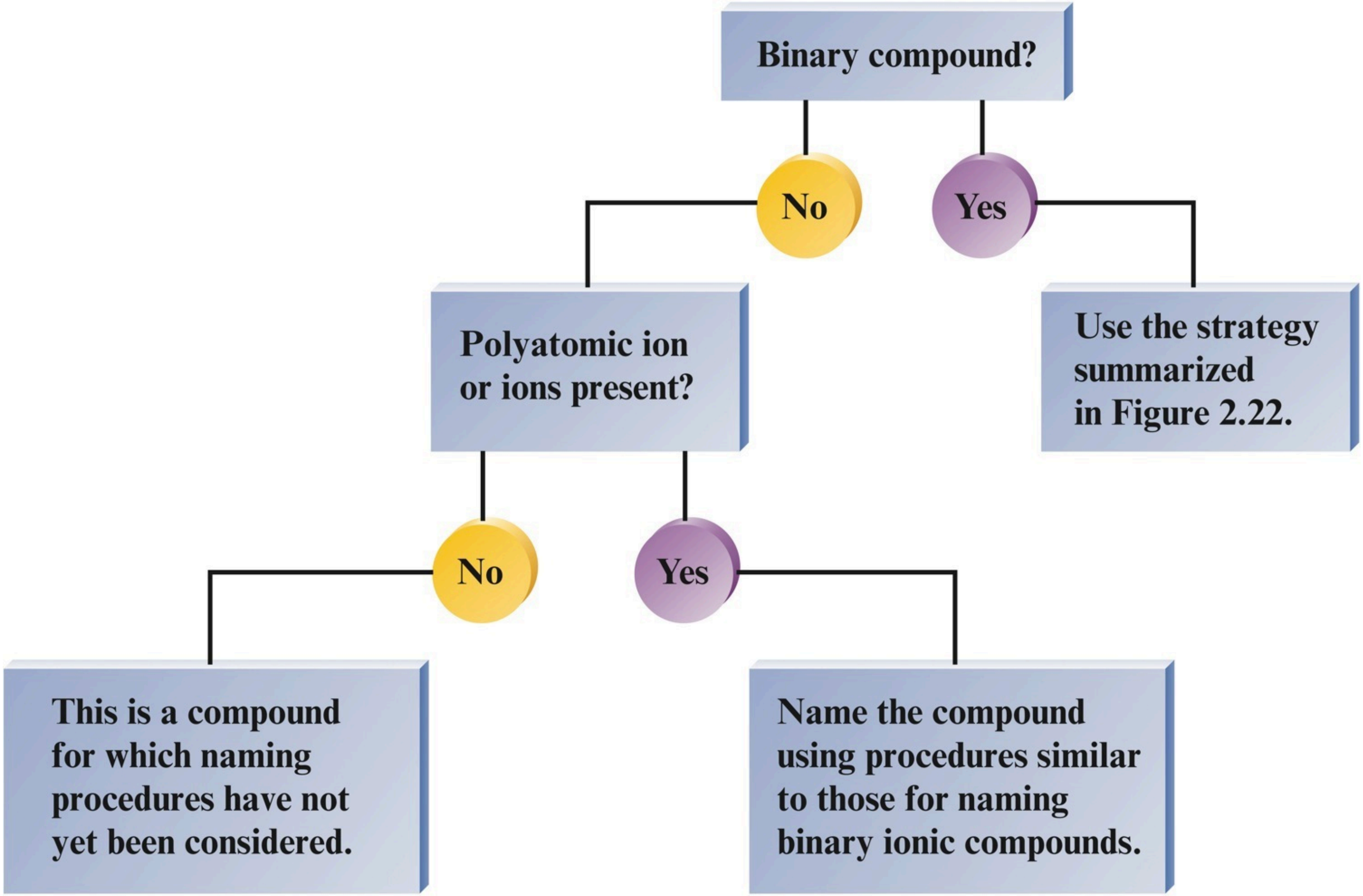
- Examples:

CO_2	Carbon dioxide
SF_6	Sulfur hexafluoride
N_2O_4	Dinitrogen tetroxide

Flowchart for Naming Binary Compounds



Overall Strategy for Naming Chemical 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.

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
H_2S	Hydrosulfuric acid

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
$HC_2H_3O_2$	Acetic acid

Acids

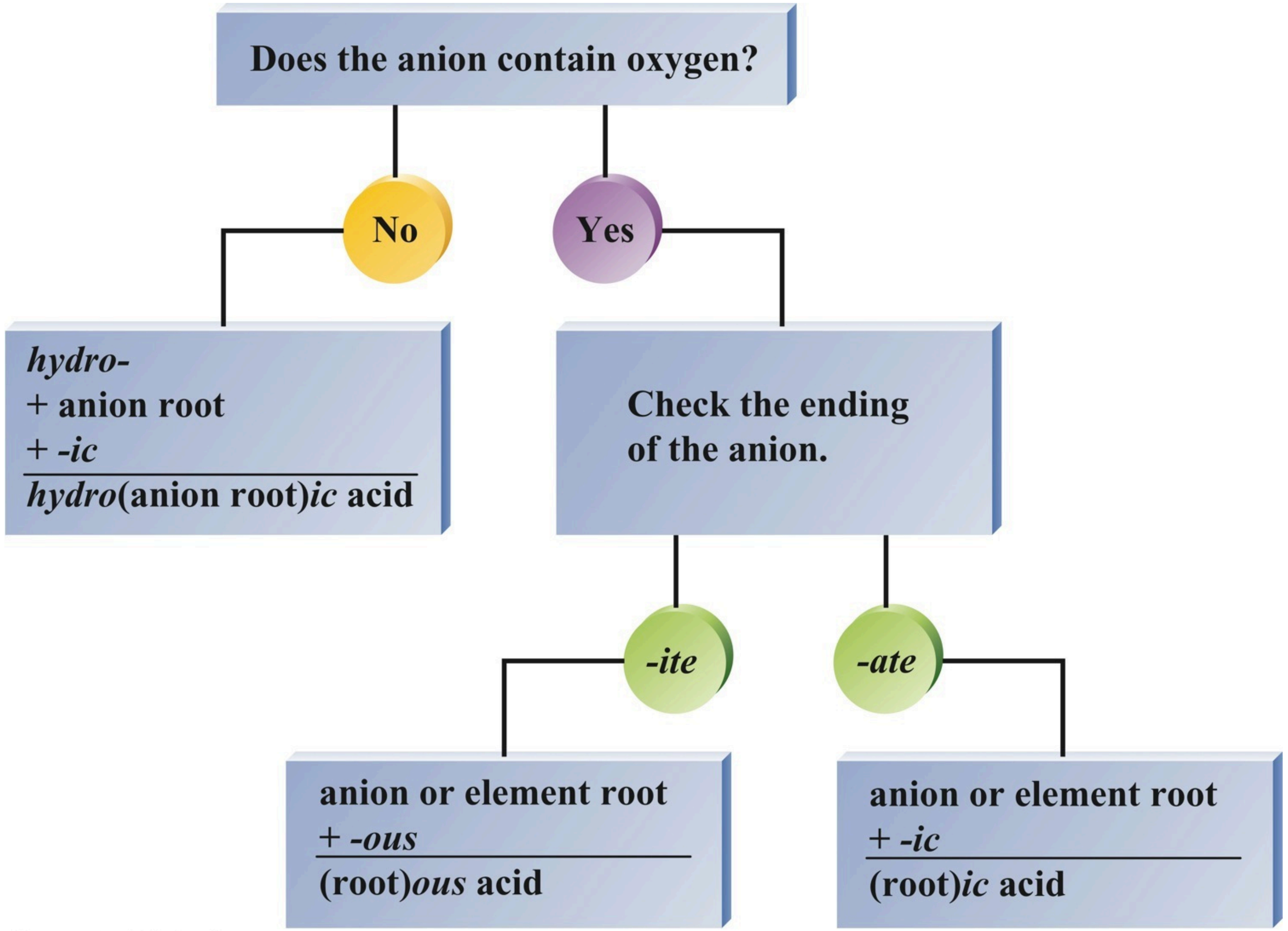
- If the anion *does* contain oxygen:
  - The suffix *–ous* is added to the root name if the anion name ends in *–ite*.
- Examples:

  HNO_2 Nitrous acid

  H_2SO_3 Sulfurous acid

  $HClO_2$ Chlorous acid

Flowchart for Naming Acids



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