

*Chemistry, The Central Science*, 10th edition  
Theodore L. Brown; H. Eugene LeMay, Jr.;  
and Bruce E. Bursten

# Chapter 2

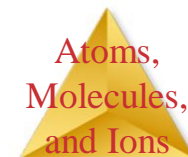
## Atoms, Molecules, and Ions

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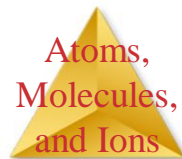
# Atomic Theory of Matter



The theory that atoms are the fundamental building blocks of matter reemerged in the early 19th century, championed by John Dalton.

# Dalton's Postulates

Each element is composed of extremely small particles called atoms.



# Dalton's Postulates

All atoms of a given element are identical to one another in mass and other properties, but the atoms of one element are different from the atoms of all other elements.



# Dalton's Postulates

Atoms of an element are not changed into atoms of a different element by chemical reactions; atoms are neither created nor destroyed in chemical reactions.



# Dalton's Postulates

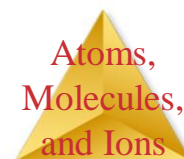
Compounds are formed when atoms of more than one element combine; a given compound always has the same relative number and kind of atoms.



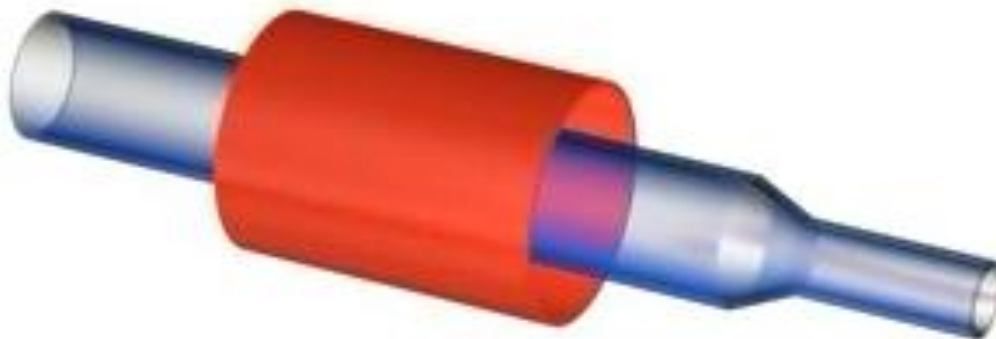
# Law of Constant Composition

*Joseph Proust (1754–1826)*

- Also known as the law of definite proportions.
- The elemental composition of a pure substance never varies.



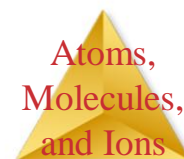
# Multiple Proportions



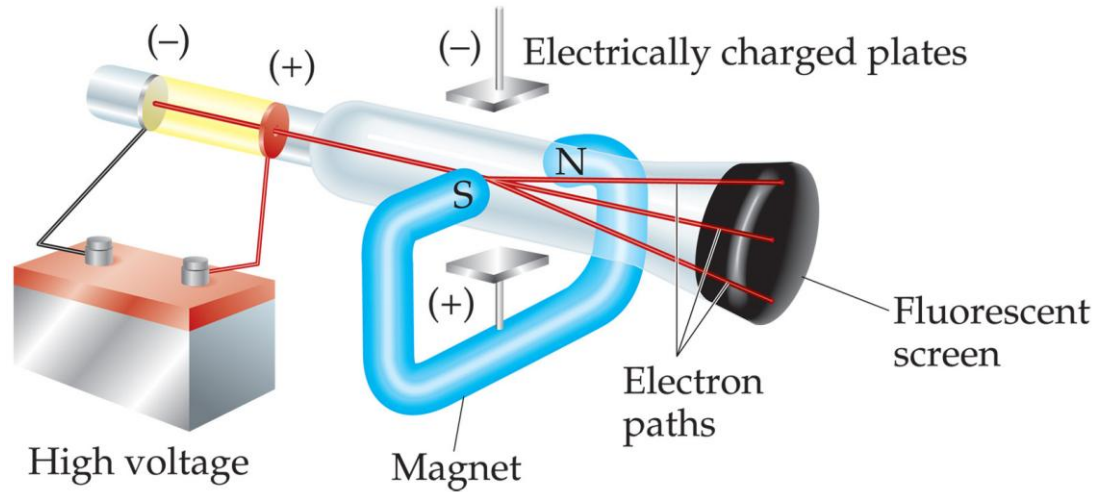


# Law of Conservation of Mass

The total mass of substances present at the end of a chemical process is the same as the mass of substances present before the process took place.

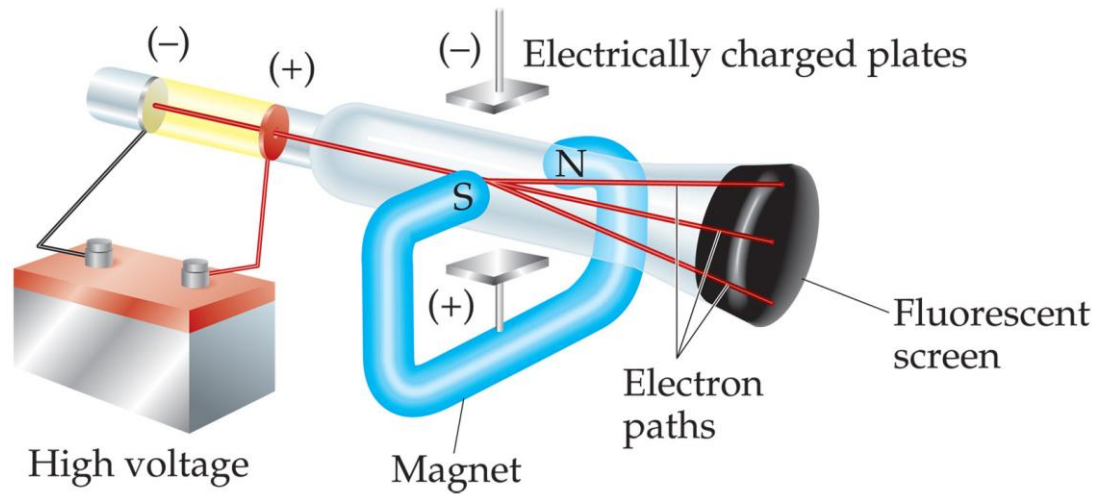


# The Electron



- Streams of negatively charged particles were found to emanate from cathode tubes.
- J. J. Thompson is credited with their discovery (1897).

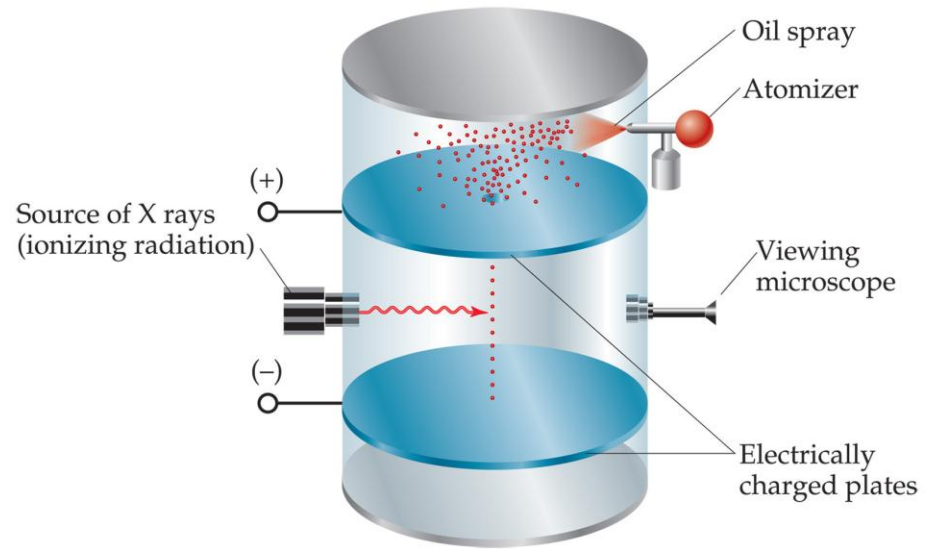
# The Electron



Thompson measured the charge/mass ratio of the electron to be  $1.76 \times 10^8$  coulombs/g.

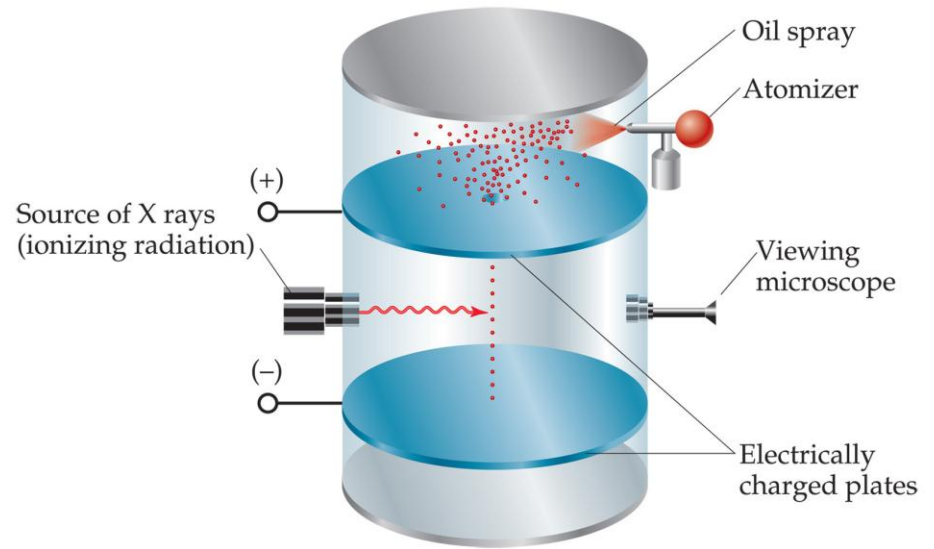
# Millikan Oil Drop Experiment

Once the charge/mass ratio of the electron was known, determination of either the charge or the mass of an electron would yield the other.



# Millikan Oil Drop Experiment

Robert Millikan  
(University of Chicago)  
determined the charge  
on the electron in  
1909.

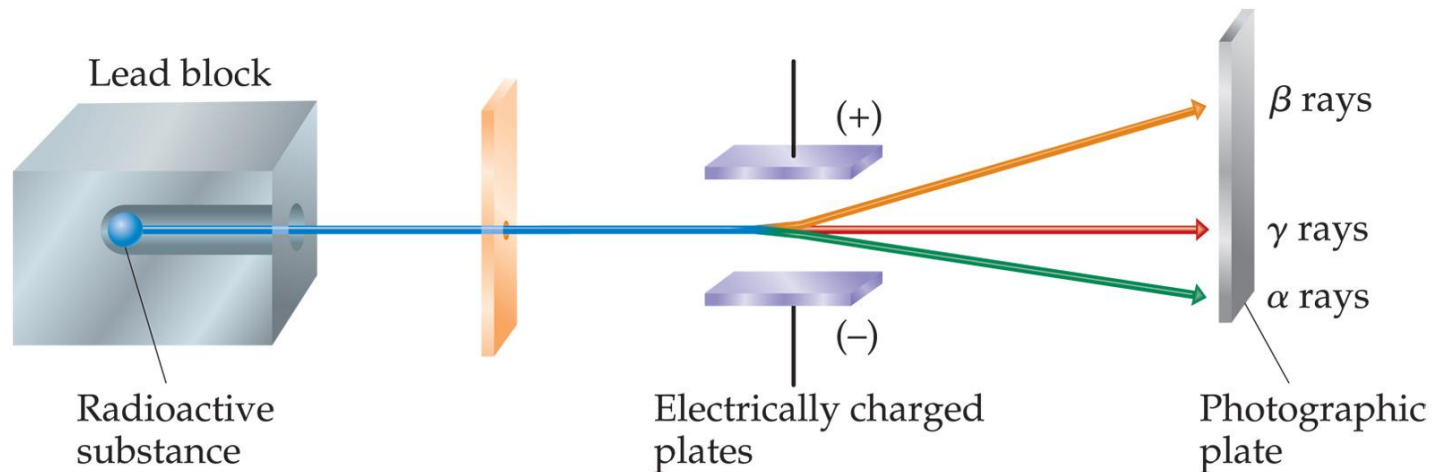


# Radioactivity:

- The spontaneous emission of radiation by an atom.
- First observed by Henri Becquerel.
- Also studied by Marie and Pierre Curie.

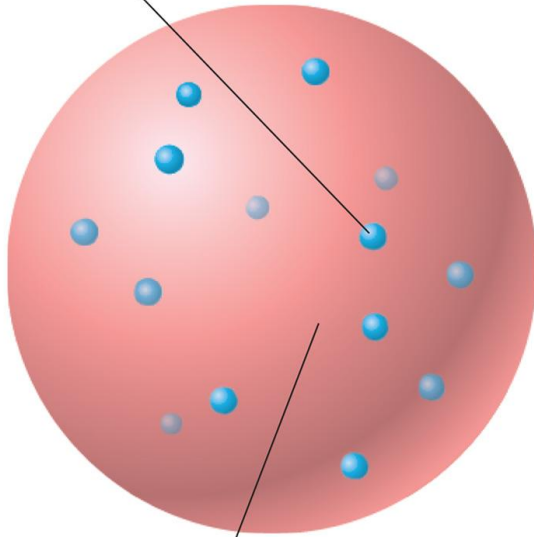
# Radioactivity

- Three types of radiation were discovered by Ernest Rutherford:
  - $\alpha$  particles
  - $\beta$  particles
  - $\gamma$  rays



# The Atom, circa 1900:

Negative  
electron

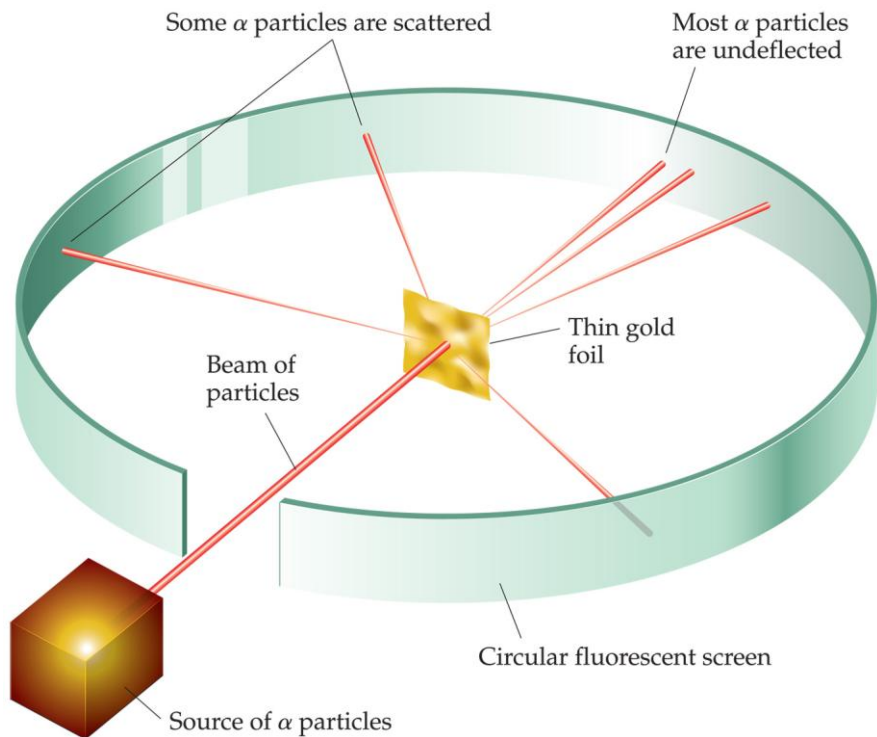


Positive charge  
spread over sphere

- “Plum pudding” model, put forward by Thompson.
- Positive sphere of matter with negative electrons imbedded in it.



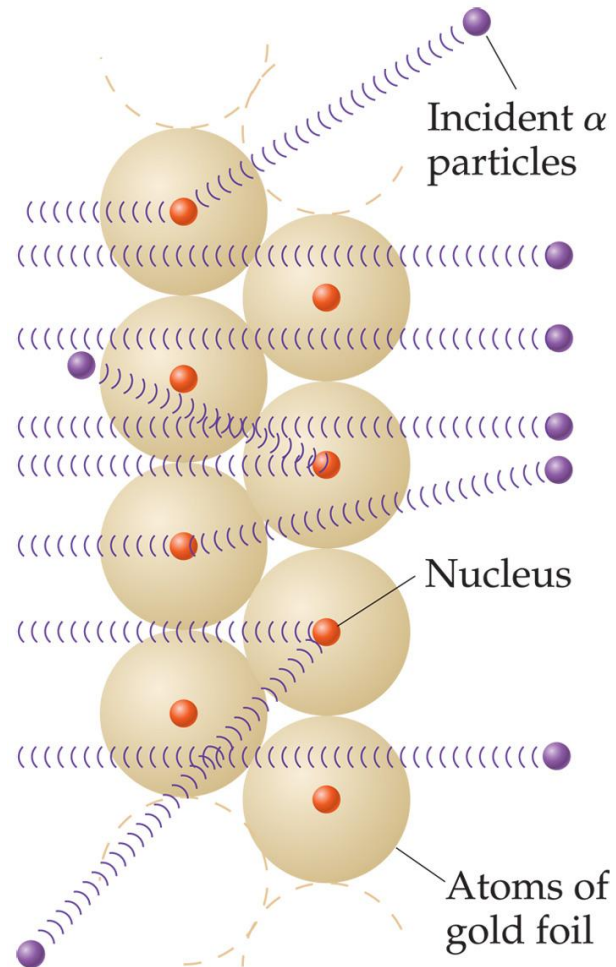
# Discovery of the Nucleus



Ernest Rutherford shot  $\alpha$  particles at a thin sheet of gold foil and observed the pattern of scatter of the particles.

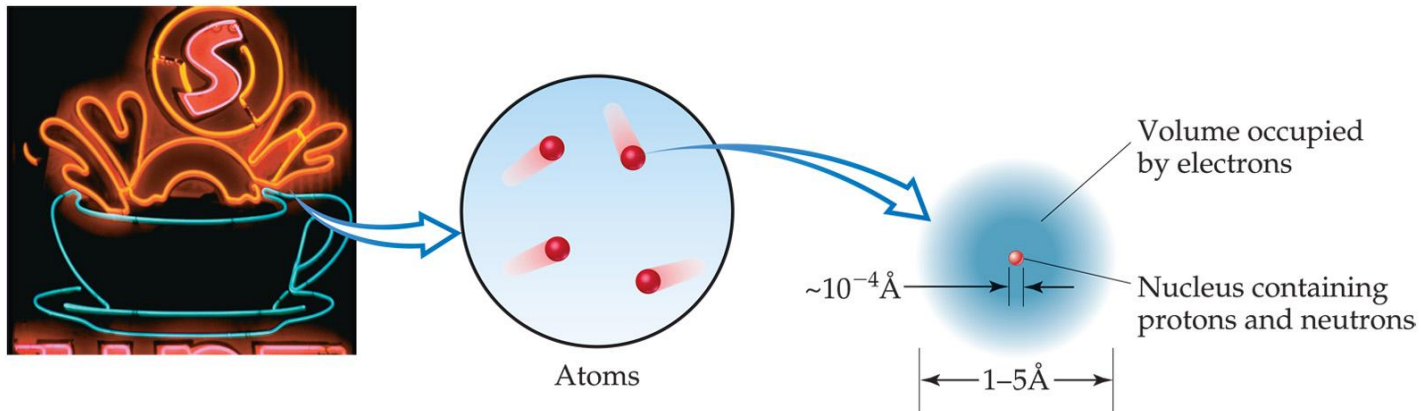
# The Nuclear Atom

Since some particles were deflected at large angles, Thompson's model could not be correct.



# The Nuclear Atom

- Rutherford postulated a very small, dense nucleus with the electrons around the outside of the atom.
- Most of the volume of the atom is empty space.



# Other Subatomic Particles

- Protons were discovered by Rutherford in 1919.
- Neutrons were discovered by James Chadwick in 1932.

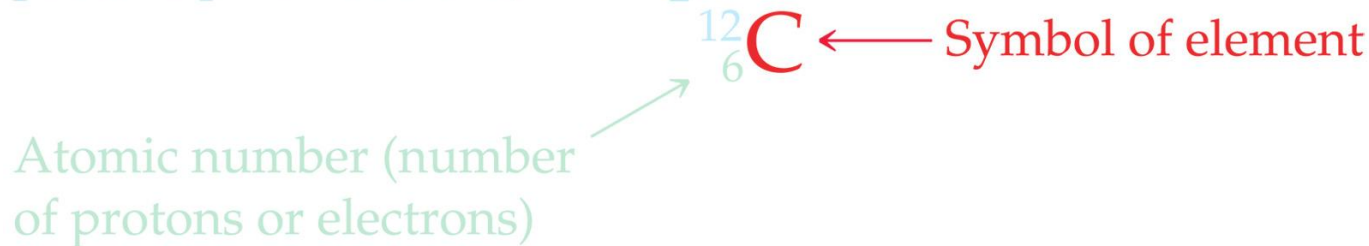
# Subatomic Particles

- Protons and electrons are the only particles that have a charge.
- Protons and neutrons have essentially the same mass.
- The mass of an electron is so small we ignore it.

Particle	Charge	Mass (amu)
Proton	Positive (1+)	1.0073
Neutron	None (neutral)	1.0087
Electron	Negative (1-)	$5.486 \times 10^{-4}$

# Symbols of Elements

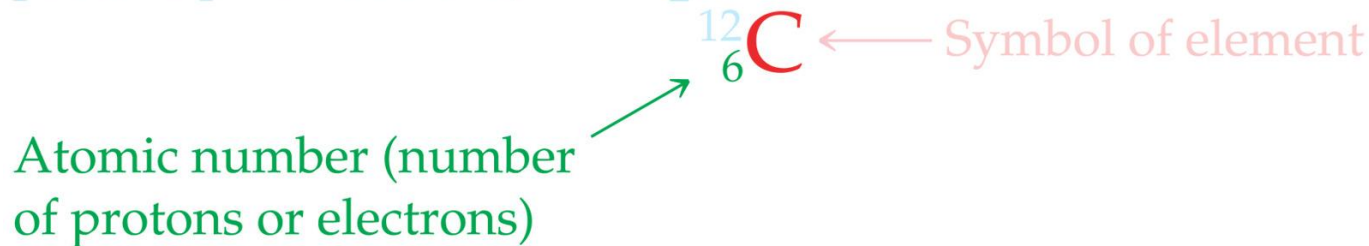
Mass number (number of  
protons plus neutrons)



Elements are symbolized by one or two letters.

# Atomic Number

Mass number (number of  
protons plus neutrons)

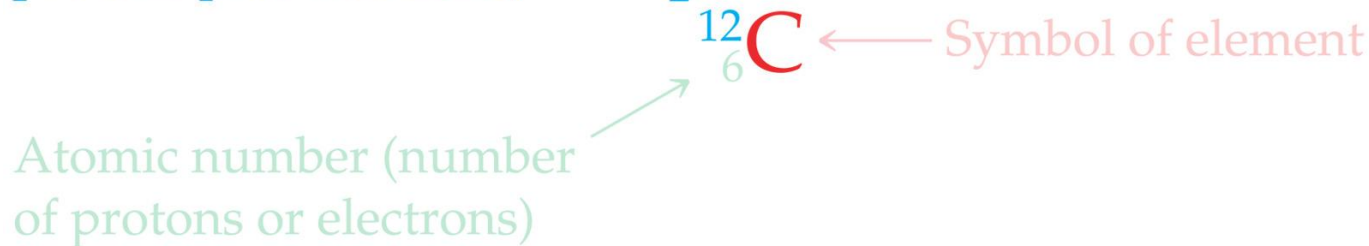


All atoms of the same element have the same  
number of protons:

The atomic number (Z)

# Atomic Mass

Mass number (number of protons plus neutrons)

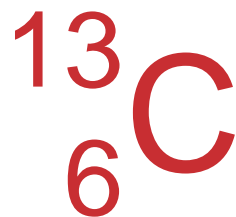


The mass of an atom in atomic mass units (amu) is the total number of protons and neutrons in the atom.

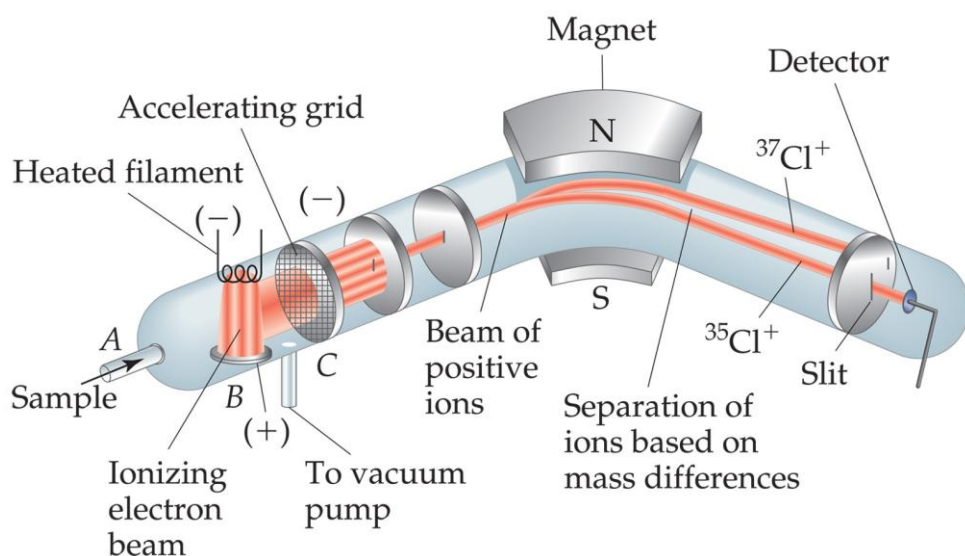


# Isotopes:

- Atoms of the same element with different masses.
- Isotopes have different numbers of neutrons.



# Atomic Mass



Atomic and molecular masses can be measured with great accuracy with a mass spectrometer.

# Average Mass

- Because in the real world we use large amounts of atoms and molecules, we use average masses in calculations.
- Average mass is calculated from the isotopes of an element weighted by their relative abundances.

# Periodic Table:

[illegible]

- A systematic catalog of elements.
- Elements are arranged in order of atomic number.

# Periodicity

Atomic number	1	2	3	4	---	9	10	11	12	---	17	18	19	20	---
Symbol	H	He	Li	Be	---	F	Ne	Na	Mg	---	Cl	Ar	K	Ca	---
		Nonreactive gas	Soft, reactive metal			Nonreactive gas	Soft, reactive metal				Nonreactive gas	Soft, reactive metal			

When one looks at the chemical properties of elements, one notices a repeating pattern of reactivities.

# Periodic Table

- The rows on the periodic chart are periods.
- Columns are groups.
- Elements in the same group have similar chemical properties.

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Atoms,  
Molecules,  
and Ions

# Groups

Group	Name	Elements
1A	Alkali metals	Li, Na, K, Rb, Cs, Fr
2A	Alkaline earth metals	Be, Mg, Ca, Sr, Ba, Ra
6A	Chalcogens	O, S, Se, Te, Po
7A	Halogens	F, Cl, Br, I, At
8A	Noble gases (or rare gases)	He, Ne, Ar, Kr, Xe, Rn

These five groups are known by their names.

# Periodic Table

[illegible]

Nonmetals are on the right side of the periodic table (with the exception of H).



# Periodic Table

[illegible]

Metalloids border the stair-step line (with the exception of Al and Po).

# Periodic Table

[illegible]

Metals are on the left side of the chart.

# Chemical Formulas



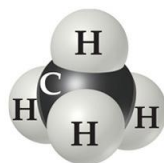
Water,  $\text{H}_2\text{O}$



Carbon dioxide,  $\text{CO}_2$



Carbon monoxide,  $\text{CO}$



Methane,  $\text{CH}_4$



Hydrogen peroxide,  $\text{H}_2\text{O}_2$



Oxygen,  $\text{O}_2$

The subscript to the right of the symbol of an element tells the number of atoms of that element in one molecule of the compound.

# Molecular Compounds



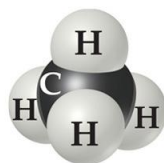
Water,  $\text{H}_2\text{O}$



Carbon dioxide,  $\text{CO}_2$



Carbon monoxide,  $\text{CO}$



Methane,  $\text{CH}_4$



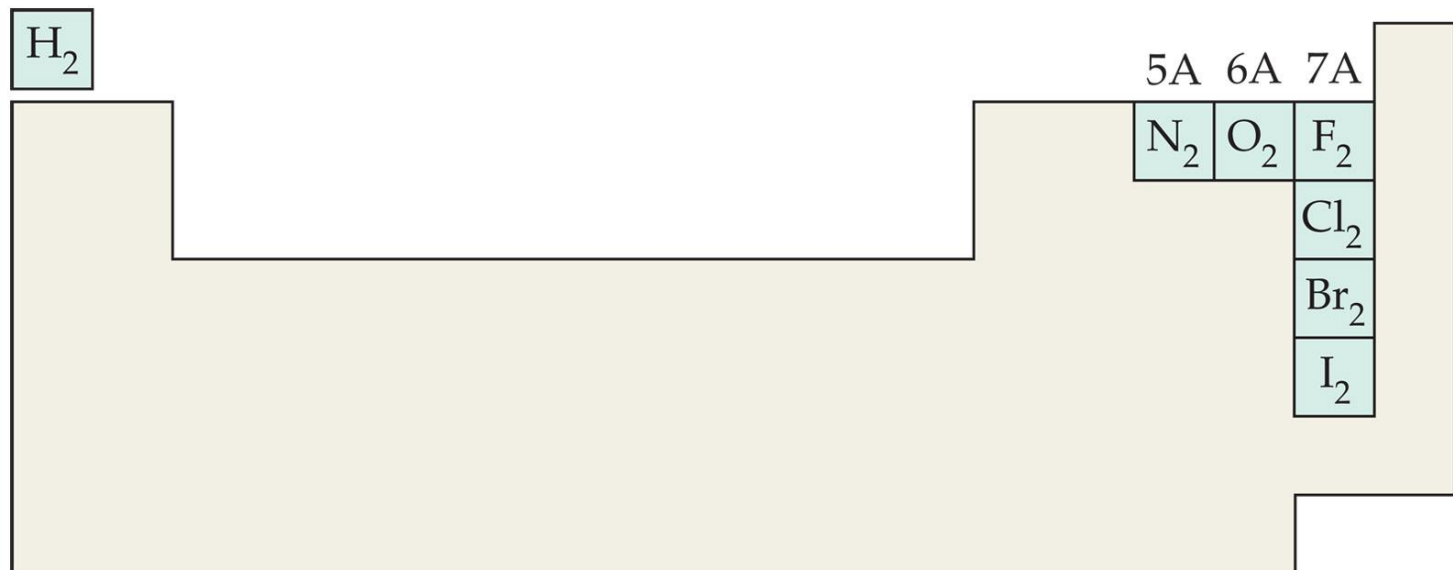
Hydrogen peroxide,  $\text{H}_2\text{O}_2$



Oxygen,  $\text{O}_2$

Molecular compounds are composed of molecules and almost always contain only nonmetals.

# Diatomic Molecules

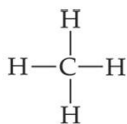


These seven elements occur naturally as molecules containing two atoms.

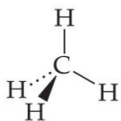
# Types of Formulas

- Empirical formulas give the lowest whole-number ratio of atoms of each element in a compound.
- Molecular formulas give the exact number of atoms of each element in a compound.

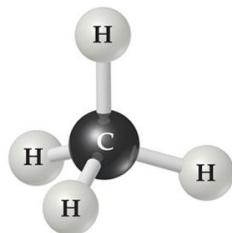
# Types of Formulas



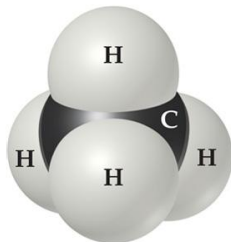
Structural formula



Perspective drawing



Ball-and-stick model



Space-filling model

- Structural formulas show the order in which atoms are bonded.
- Perspective drawings also show the three-dimensional array of atoms in a compound.

# Ions

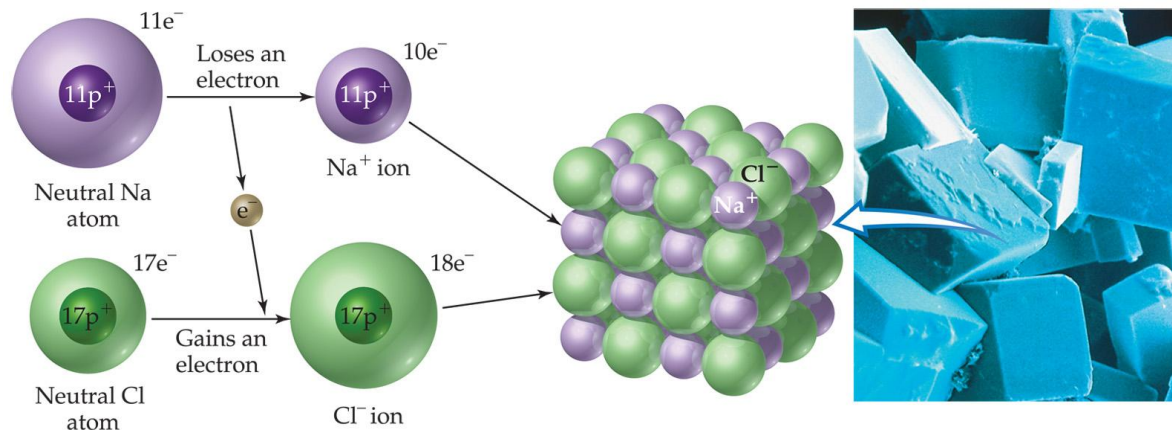
1A	2A											3A	4A	5A	6A	7A	8A
H <sup>+</sup>														N <sup>3-</sup>	O <sup>2-</sup>	H <sup>-</sup>	N O B L E  G A S E S
Li <sup>+</sup>												Al <sup>3+</sup>			S <sup>2-</sup>	Cl <sup>-</sup>	
Na <sup>+</sup>	Mg <sup>2+</sup>														Se <sup>2-</sup>	Br <sup>-</sup>	
K <sup>+</sup>	Ca <sup>2+</sup>														Te <sup>2-</sup>	I <sup>-</sup>	
Rb <sup>+</sup>	Sr <sup>2+</sup>																
Cs <sup>+</sup>	Ba <sup>2+</sup>																

- When atoms lose or gain electrons, they become ions.
  - Cations are positive and are formed by elements on the left side of the periodic chart.
  - Anions are negative and are formed by elements on the right side of the periodic chart.

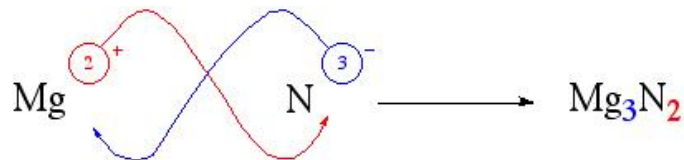


# Ionic Bonds

Ionic compounds (such as NaCl) are generally formed between metals and nonmetals.



# Writing Formulas

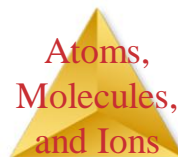


- Because compounds are electrically neutral, one can determine the formula of a compound this way:
  - The charge on the cation becomes the subscript on the anion.
  - The charge on the anion becomes the subscript on the cation.
  - If these subscripts are not in the lowest whole-number ratio, divide them by the greatest common factor.

# Common Cations

Charge	Formula	Name	Formula	Name
1+	<b>H<sup>+</sup></b>	<b>Hydrogen ion</b>	<b>NH<sub>4</sub><sup>+</sup></b>	<b>Ammonium ion</b>
	Li <sup>+</sup>	Lithium ion	Cu <sup>+</sup>	Copper(I) or cuprous ion
	<b>Na<sup>+</sup></b>	<b>Sodium ion</b>		
	<b>K<sup>+</sup></b>	<b>Potassium ion</b>		
	Cs <sup>+</sup>	Cesium ion		
	<b>Ag<sup>+</sup></b>	<b>Silver ion</b>		
2+	<b>Mg<sup>2+</sup></b>	<b>Magnesium ion</b>	Co <sup>2+</sup>	Cobalt(II) or cobaltous ion
	<b>Ca<sup>2+</sup></b>	<b>Calcium ion</b>	<b>Cu<sup>2+</sup></b>	<b>Copper(II)</b> or cupric ion
	Sr <sup>2+</sup>	Strontium ion	<b>Fe<sup>2+</sup></b>	<b>Iron(II)</b> or ferrous ion
	Ba <sup>2+</sup>	Barium ion	Mn <sup>2+</sup>	Manganese(II) or manganous ion
	<b>Zn<sup>2+</sup></b>	<b>Zinc ion</b>	Hg <sub>2</sub> <sup>2+</sup>	Mercury(I) or mercurous ion
	Cd <sup>2+</sup>	Cadmium ion	<b>Hg<sup>2+</sup></b>	<b>Mercury(II)</b> or mercuric ion
			Ni <sup>2+</sup>	Nickel(II) or nickelous ion
			<b>Pb<sup>2+</sup></b>	<b>Lead(II)</b> or plumbous ion
			Sn <sup>2+</sup>	Tin(II) or stannous ion
3+	<b>Al<sup>3+</sup></b>	<b>Aluminum ion</b>	Cr <sup>3+</sup>	Chromium(III) or chromic ion
			<b>Fe<sup>3+</sup></b>	<b>Iron(III)</b> or ferric ion

\*The most common ions are in boldface.



# Common Anions

Charge	Formula	Name	Formula	Name
1−	H <sup>−</sup>	Hydride ion	C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> <sup>−</sup>	Acetate ion
	F <sup>−</sup>	<b>Fluoride ion</b>	ClO <sub>3</sub> <sup>−</sup>	Chlorate ion
	Cl <sup>−</sup>	<b>Chloride ion</b>	ClO <sub>4</sub> <sup>−</sup>	<b>Perchlorate ion</b>
	Br <sup>−</sup>	<b>Bromide ion</b>	NO <sub>3</sub> <sup>−</sup>	<b>Nitrate ion</b>
	I <sup>−</sup>	<b>Iodide ion</b>	MnO <sub>4</sub> <sup>−</sup>	Permanganate ion
	CN <sup>−</sup>	Cyanide ion		
	OH <sup>−</sup>	<b>Hydroxide ion</b>		
2−	O <sup>2−</sup>	<b>Oxide ion</b>	CO <sub>3</sub> <sup>2−</sup>	<b>Carbonate ion</b>
	O <sub>2</sub> <sup>2−</sup>	Peroxide ion	CrO <sub>4</sub> <sup>2−</sup>	Chromate ion
	S <sup>2−</sup>	<b>Sulfide ion</b>	Cr <sub>2</sub> O <sub>7</sub> <sup>2−</sup>	Dichromate ion
			SO <sub>4</sub> <sup>2−</sup>	<b>Sulfate ion</b>
3−	N <sup>3−</sup>	Nitride ion	PO <sub>4</sub> <sup>3−</sup>	<b>Phosphate ion</b>

\*The most common ions are in boldface.

# Inorganic Nomenclature

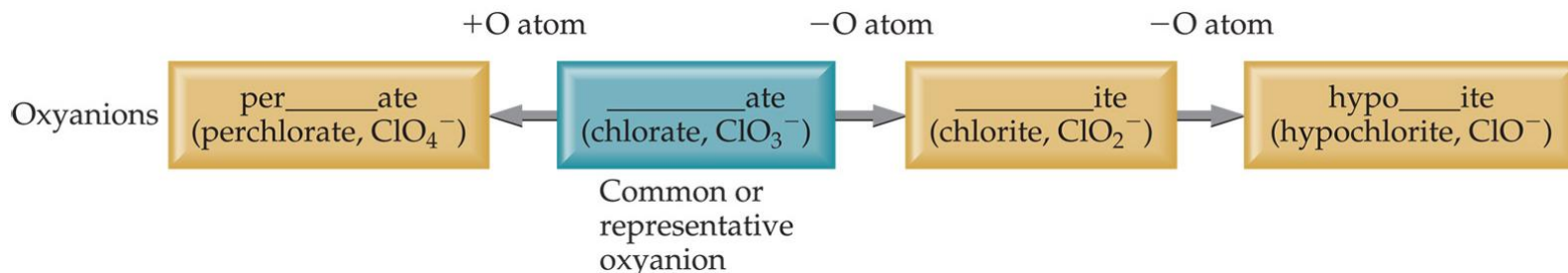
- Write the name of the cation.
- If the anion is an element, change its ending to *-ide*; if the anion is a polyatomic ion, simply write the name of the polyatomic ion.
- If the cation can have more than one possible charge, write the charge as a Roman numeral in parentheses.

# Patterns in Oxyanion Nomenclature

- When there are two oxyanions involving the same element:
  - The one with fewer oxygens ends in *-ite*
    - $\text{NO}_2^-$  : nitrite;  $\text{SO}_3^{2-}$  : sulfite
  - The one with more oxygens ends in *-ate*
    - $\text{NO}_3^-$  : nitrate;  $\text{SO}_4^{2-}$  : sulfate

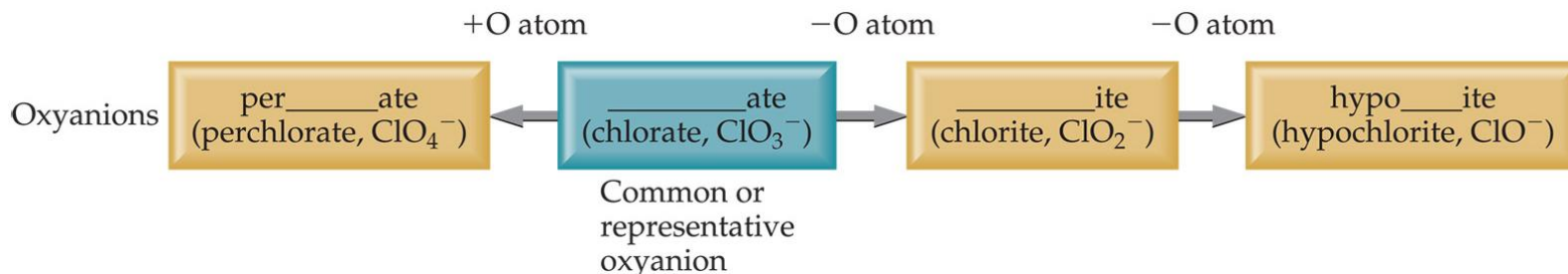
# Patterns in Oxyanion Nomenclature

- The one with the fewest oxygens has the prefix *hypo-* and ends in *-ite*
  - $\text{ClO}^-$  : hypochlorite
- The one with the most oxygens has the prefix *per-* and ends in *-ate*
  - $\text{ClO}_4^-$  : perchlorate



# Patterns in Oxyanion Nomenclature

- The one with the second fewest oxygens ends in *-ite*
  - $\text{ClO}_2^-$  : chlorite
- The one with the second most oxygens ends in *-ate*
  - $\text{ClO}_3^-$  : chlorate





# Acid Nomenclature

Anion		Acid
____ide (chloride, $\text{Cl}^-$ )	add $\text{H}^+$ ions	hydro____ic acid (hydrochloric acid, $\text{HCl}$ )
____ate (chlorate, $\text{ClO}_3^-$ ) (perchlorate, $\text{ClO}_4^-$ )	add $\text{H}^+$ ions	____ic acid (chloric acid, $\text{HClO}_3$ ) (perchloric acid, $\text{HClO}_4$ )
____ite (chlorite, $\text{ClO}_2^-$ ) (hypochlorite, $\text{ClO}^-$ )	add $\text{H}^+$ ions	____ous acid (chlorous acid, $\text{HClO}_2$ ) (hypochlorous acid, $\text{HClO}$ )

- If the anion in the acid ends in *-ide*, change the ending to *-ic acid* and add the prefix *hydro-* :
  - $\text{HCl}$ : hydrochloric acid
  - $\text{HBr}$ : hydrobromic acid
  - $\text{HI}$ : hydroiodic acid

# Acid Nomenclature

Anion		Acid
____ide (chloride, $\text{Cl}^-$ )	add $\text{H}^+$ ions	hydro____ic acid (hydrochloric acid, $\text{HCl}$ )
____ate (chlorate, $\text{ClO}_3^-$ ) (perchlorate, $\text{ClO}_4^-$ )	add $\text{H}^+$ ions	____ic acid (chloric acid, $\text{HClO}_3$ ) (perchloric acid, $\text{HClO}_4$ )
____ite (chlorite, $\text{ClO}_2^-$ ) (hypochlorite, $\text{ClO}^-$ )	add $\text{H}^+$ ions	____ous acid (chlorous acid, $\text{HClO}_2$ ) (hypochlorous acid, $\text{HClO}$ )

- If the anion in the acid ends in *-ite*, change the ending to *-ous acid*:
  - $\text{HClO}$ : hypochlorous acid
  - $\text{HClO}_2$ : chlorous acid

# Acid Nomenclature

Anion		Acid
____ide (chloride, $\text{Cl}^-$ )	add $\text{H}^+$ ions	hydro____ic acid (hydrochloric acid, $\text{HCl}$ )
____ate (chlorate, $\text{ClO}_3^-$ ) (perchlorate, $\text{ClO}_4^-$ )	add $\text{H}^+$ ions	____ic acid (chloric acid, $\text{HClO}_3$ ) (perchloric acid, $\text{HClO}_4$ )
____ite (chlorite, $\text{ClO}_2^-$ ) (hypochlorite, $\text{ClO}^-$ )	add $\text{H}^+$ ions	____ous acid (chlorous acid, $\text{HClO}_2$ ) (hypochlorous acid, $\text{HClO}$ )

- If the anion in the acid ends in *-ate*, change the ending to *-ic acid*:
  - $\text{HClO}_3$ : chloric acid
  - $\text{HClO}_4$ : perchloric acid

# Nomenclature of Binary Compounds

<i>Prefix</i>	<i>Meaning</i>
<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

- The less electronegative atom is usually listed first.
- A prefix is used to denote the number of atoms of each element in the compound (*mono-* is not used on the first element listed, however.)

# Nomenclature of Binary Compounds

<i>Prefix</i>	<i>Meaning</i>
<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

- The ending on the more electronegative element is changed to *-ide*.

- $\text{CO}_2$ : carbon dioxide
- $\text{CCl}_4$ : carbon tetrachloride

# Nomenclature of Binary Compounds

<i>Prefix</i>	<i>Meaning</i>
<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

If the prefix ends with *a* or *o* and the name of the element begins with a vowel, the two successive vowels are often elided into one:

$\text{N}_2\text{O}_5$ : dinitrogen pentoxide