Chapter 2:
Atoms, Molecules, and Ions

Learning outcomes:

- Learn the basic postulates of Dalton's atomic theory.
- Describe the key experiments that led to the discovery of electrons and the nuclear model of the atom.
- Describe the structure of the atom in terms of protons, neutrons, and electrons and express the relative electrical charges and masses of these subatomic particles.
- Use chemical symbols together with atomic number and mass number to express the subatomic composition of isotopes.
- Calculate the atomic weight of an element from the masses of individual atoms and a knowledge of natural abundances.
- Describe how elements are organized in the periodic table by atomic number and by similarities in chemical behavior, giving rise to periods and groups.
- Identify the locations of metals and nonmetals in the periodic table.
- Distinguish between molecular substances and ionic substances in terms of their composition.
- Distinguish between empirical formulas and molecular formulas.
- Describe how molecular formulas and structural formulas are used to represent the compositions of molecules.
- Explain how ions are formed by the gain or loss of electrons and use the periodic table to predict the charges of common ions.
- Write the empirical formulas of ionic compounds, given the charges of their component ions.
- Write the name of an ionic compound given its chemical formula or write the chemical formula given its name.
- Name or write chemical formulas for binary inorganic compounds and for acids.

Atomic Theory of Matter

•Democritus (Greek philosopher) believed that there was a smallest particle— "atomos" (uncuttable, indivisible)—that made up all of nature.

•Experiments in the eighteenth and nineteenth centuries led to an organized atomic theory by John Dalton in the early 1800s, which explained several laws known at that time:

- The law of constant composition
- The law of conservation of mass
- The law of multiple proportions

Joseph Proust
(1754–1826)

Law of Constant Composition

- Also known as the *law of definite proportions*.
- The elemental composition of a pure substance never varies.
- In a given compound, the relative numbers and kinds of atoms are constant.
- Basis of Dalton's Postulate #4

Antoine Laurent Lavoisier
(1743-1794)

Law of Conservation of Mass (Matter)

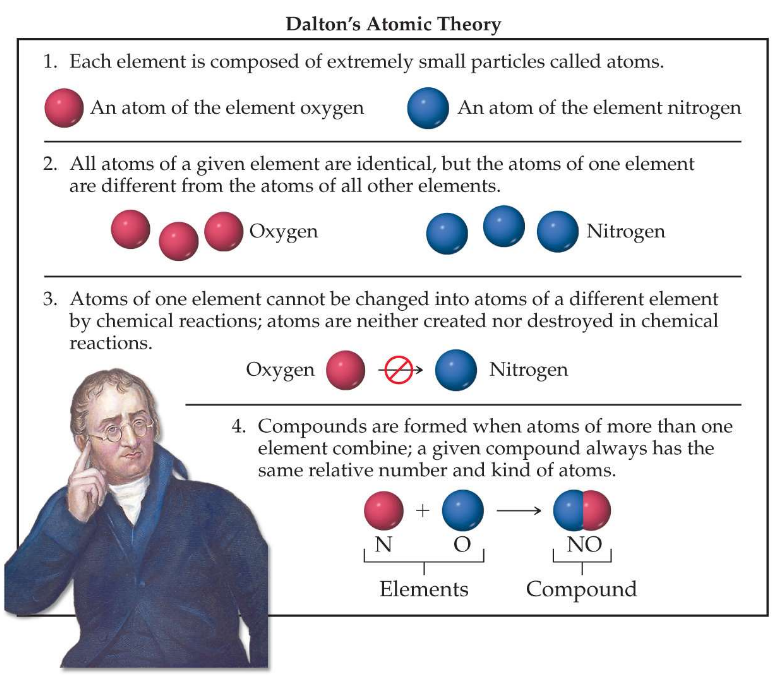
- The total mass of substances present at the end of a chemical process is the same as the mass of substances present before the process took place.
- Basis of Dalton's Postulate #3
- Can't create matter in a chemical reaction!

John Dalton
(1766-1844)

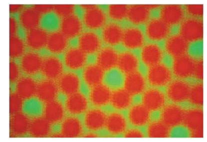
Law of Multiple Proportions

- If two elements A and B combine to form more than one compound, the masses of B that can combine with a given mass of A are in the ratio of small whole numbers.

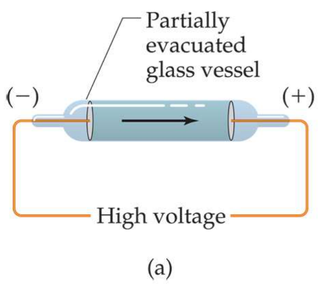
four postulates

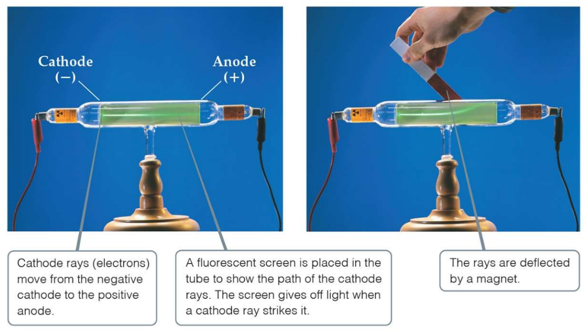


Discovery of Atomic Structure



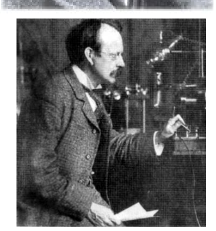
Mid 1800's – scientists studied electrical discharge through partially evacuated tubes.

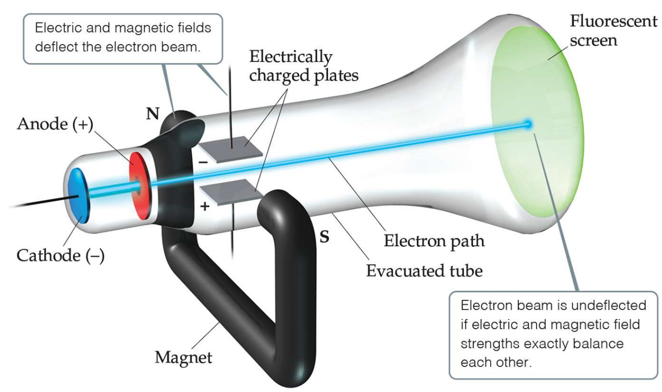




Path of cathode rays deflected by presence of a magnet.

J.J. Thomson (1856-1940)

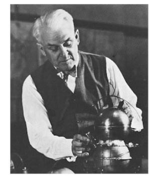


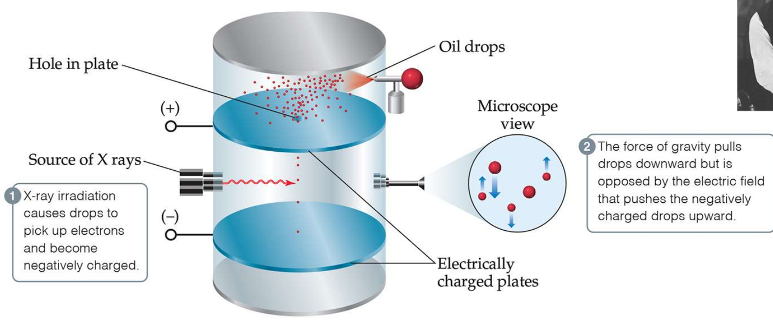


Proposed in 1897 that Cathode Rays were actually particles (negatively charged) that we now know are electrons.

Thomson measured the charge/mass ratio of the electron to be 1.76×10^8 coulombs/gram (C/g).

Millikan Oil Drop Experiment



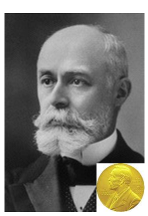


- Once the charge/mass ratio of the electron was known, determination of either the charge or the mass of an electron would yield the other.
- Robert Millikan (1909) determined the charge on the electron, equal to 1.60×10^{-19} C.
- Mass of an e^- could be calculated as 9.11×10^{-28} g.

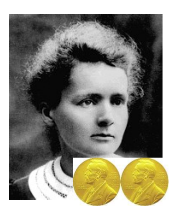
Radioactivity:

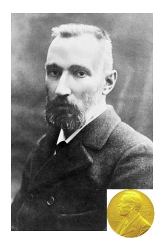
The spontaneous emission of radiation by an atom.

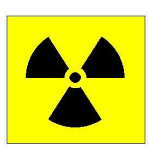
First observed by Henri Becquerel. **(1852-1908)**



Also studied by Marie (1867-1934) and Pierre Curie (1857-1906). Discovered Po and Ra. Suggested that atoms of certain substances can disintegrate.







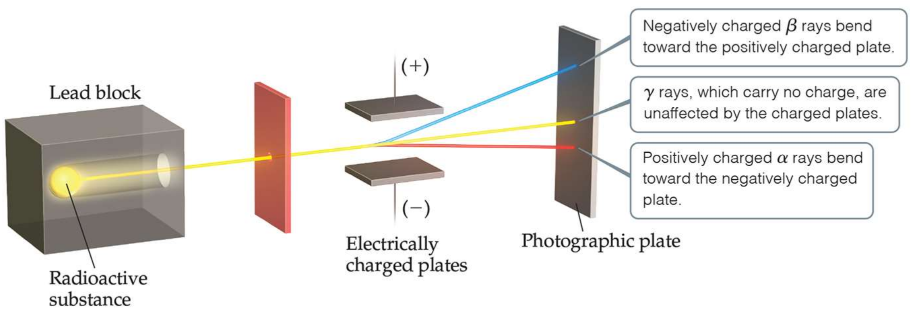
Radioactivity

Three types of radiation were discovered by Ernest Rutherford:

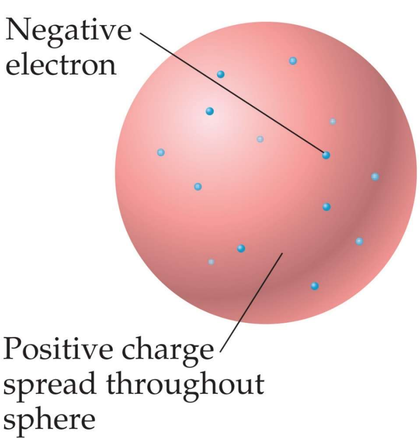
α particles positively charged (+2), large mass

β particles negatively charged (-1), small mass

γ rays no charge, no mass

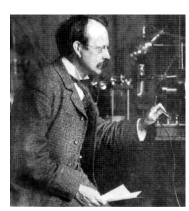


The Atom, circa 1900

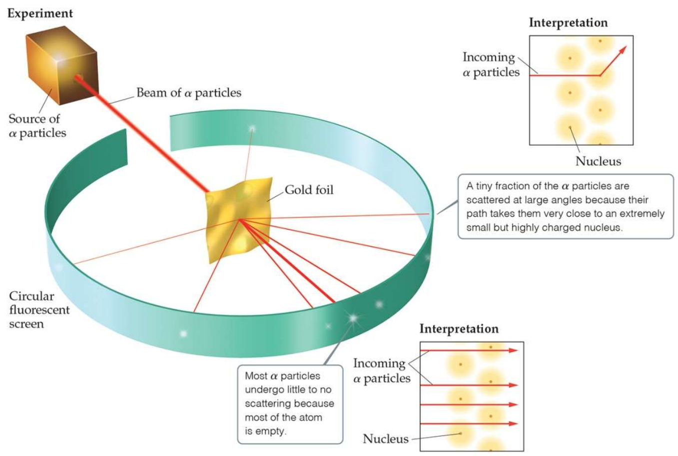


- "Plum pudding" model, put forward by Thomson.
- Positive sphere of matter with negative electrons imbedded in it.

Sir Joseph John Thomson (1856-1940)



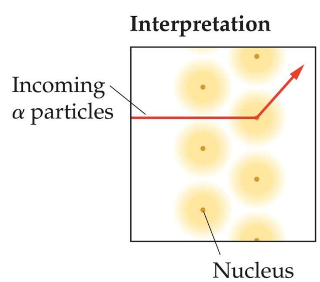
Discovery of the Nucleus



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

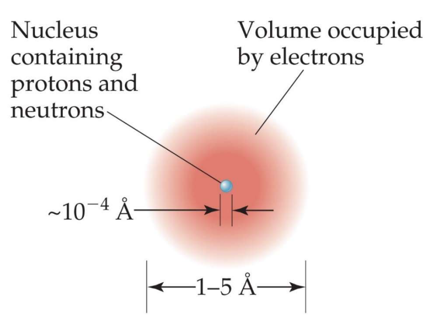
Ernest Rutherford
(1871-1937)

The Nuclear Atom

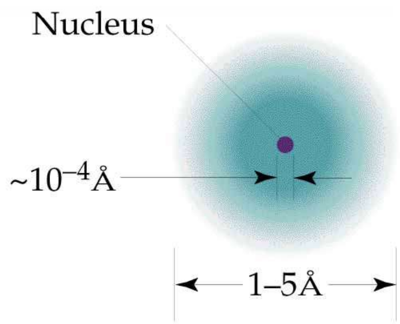


- Since some particles were deflected at large angles. Thomson's model could not be correct.

- Rutherford's nuclear model of the atom: all of the positive charge and most of the mass is concentrated at the center – the nucleus. Electrons occupy the rest of the space (volume) of the atom.



Subatomic Particles



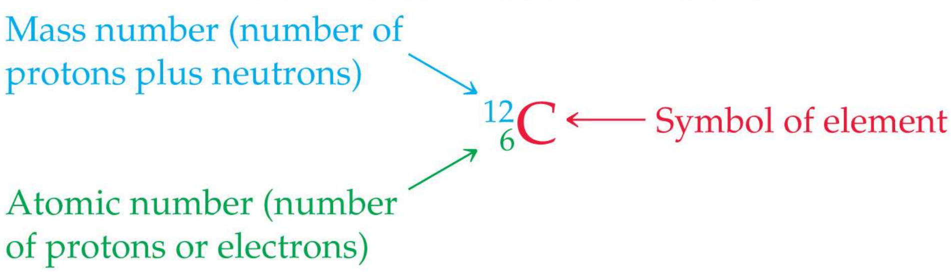
Neutrons were discovered by James Chadwick in 1932.

TABLE 2.1 Comparison of the Proton, Neutron, and Electron

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

1 amu is defined as 1/12 the mass of an unbound carbon atom carbon-12 at its rest state.

Atomic and Mass Numbers

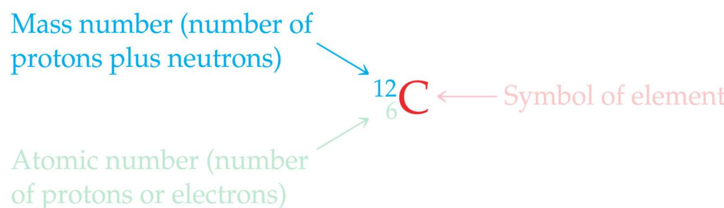


Elements are represented by a one or two letter **symbol**, for which the first letter is always capitalized. C is the symbol for carbon.

Atomic number: equal to the number of protons in the nucleus. All atoms of the same element have the same number of protons. Denoted by “Z”.

Mass number: equal to the sum of the number of protons and neutrons for an atom.

Atomic Mass



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

A carbon (C) atom with 6 protons and 6 neutrons is assigned a mass of exactly 12 amu.

Atomic mass unit (amu) is one-twelfth of the mass of an atom of carbon with 6 protons and 6 neutrons.

$1 \text{ amu} = 1.66054 \times 10^{-24} \text{ g}$

$1 \text{ g} = 6.02214 \times 10^{23} \text{ amu}$

Isotopes

- Isotopes are atoms of the same element with different masses.
- Isotopes have different numbers of neutrons, thus different mass numbers.

$^{11}_{6}C \quad ^{12}_{6}C \quad ^{13}_{6}C \quad ^{14}_{6}C$

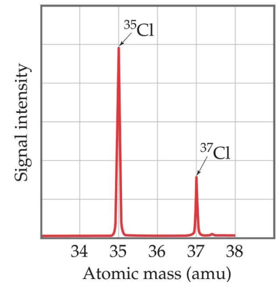
TABLE 2.2 Some Isotopes of Carbon[a]

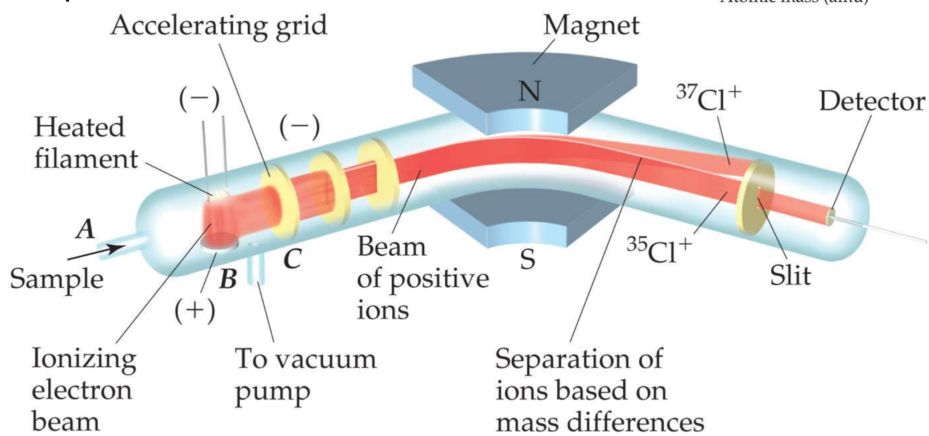
Symbol	Number of Protons	Number of Electrons	Number of Neutrons
^{11}C	6	6	5
^{12}C	6	6	6
^{13}C	6	6	7
^{14}C	6	6	8

[a] Almost 99% of the carbon found in nature is ^{12}C.

Atomic Mass

Atomic and molecular masses (actually the mass to charge ratio) can be measured with great accuracy with a mass spectrometer.





Atomic Masses of the Elements

- **Isotopic mass** is the mass in amu (u), of a particular isotope of an element.
- Different isotopes of an element all react essentially the same, so a *weighted average* of isotopic masses can be used in calculations.
- The **atomic weight** is the weighted average mass, of the naturally occurring element. It is calculated from the isotopes of an element weighted by their relative abundances.

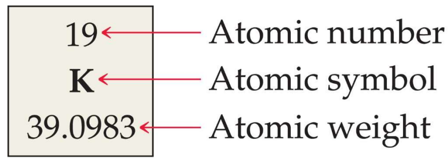
Atomic weight = $\text{fraction}_A m_A + \text{fraction}_B m_B + \cdots$

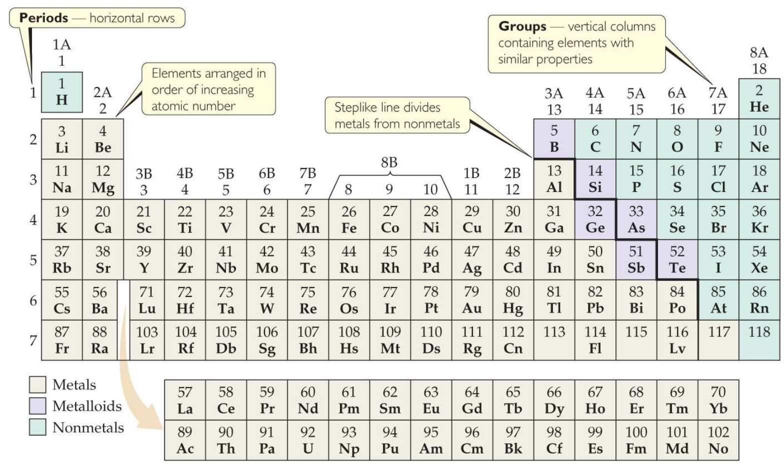
Atomic weight = $\sum$[(isotope mass)×(fractional isotope abundance)]

Average atomic mass is known as the atomic weight.

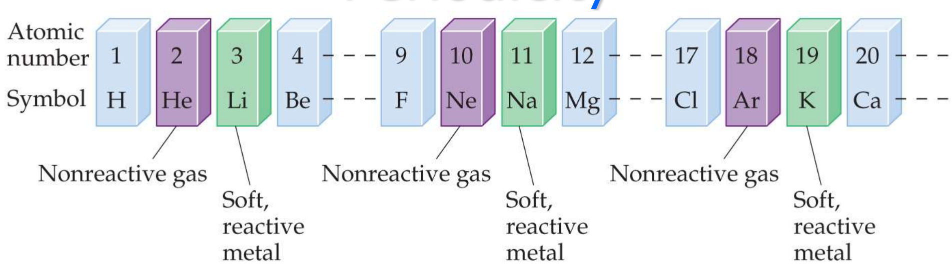
Periodic Table:

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





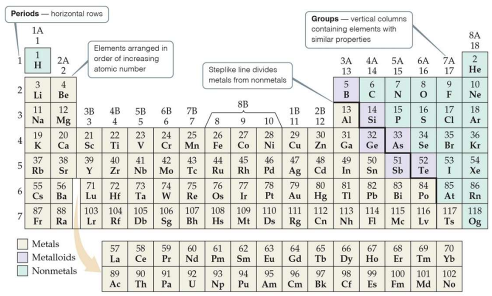
Periodicity



A repeating pattern of chemical and physical properties is observed.

- **Law of chemical periodicity**: the properties of the elements are periodic functions of atomic number.
- Elements in the same *group* have similar chemical and physical properties.

The Periodic Table



Families/Groups: vertical columns. These have similar chemical/physical properties.

A groups: the main group elements

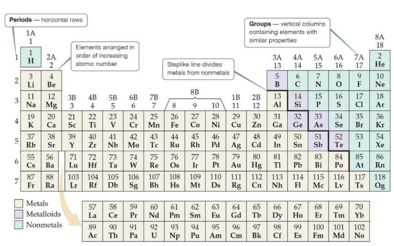
B groups: the transition elements

TABLE 2.3 Names of Some Groups in the Periodic Table

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	He, Ne, Ar, Kr, Xe, Rn

The Periodic Table: Metals

Less metallic →



Metals are on the left side of the periodic table.

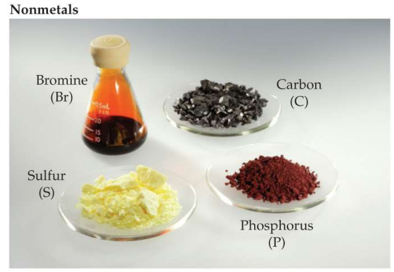
- Shiny luster, ductile, malleable
- Conduct heat and electricity
- Solids (except mercury)

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

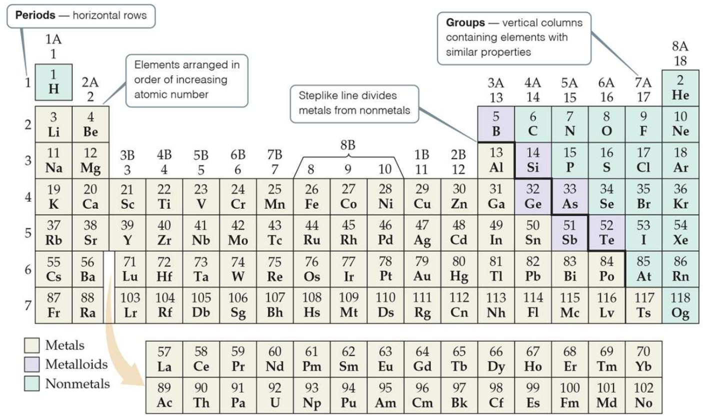
- Have a wide variety of properties (solids, liquids and gases) and do not conduct electricity well (except C as graphite).

Metals

Nonmetals



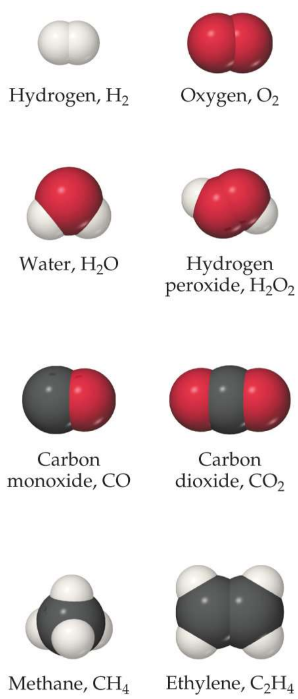
The Periodic Table: Metalloids



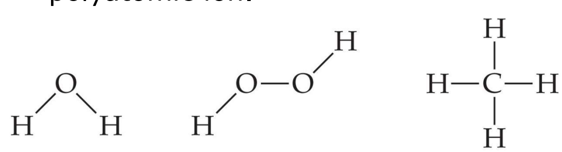
Metalloids or semimetals: some ***physical*** characteristics of metals and ***chemical*** characteristics of nonmetal.

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

Chemical Formulas



- 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 often contain only nonmetals.
- The attraction between molecules are often relatively weak, explaining why gases and liquids are common among molecular substances.
- Carbon is typically listed first in the formula, unless C is part of a polyatomic ion.



Water Hydrogen peroxide Methane

Diatomic Molecules

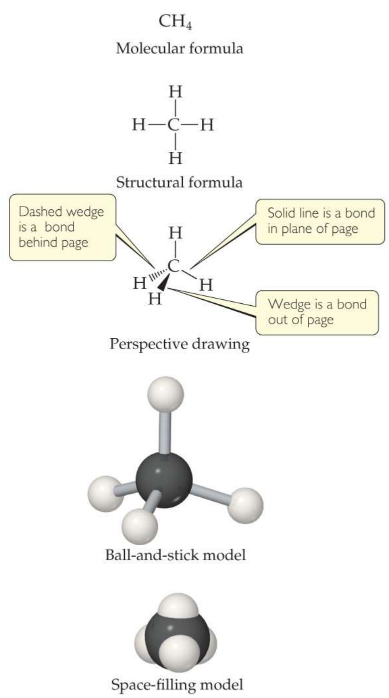
Seven elements occur naturally as molecules containing two atoms:

- Hydrogen
- Nitrogen
- Oxygen
- Fluorine
- Chlorine
- Bromine
- Iodine

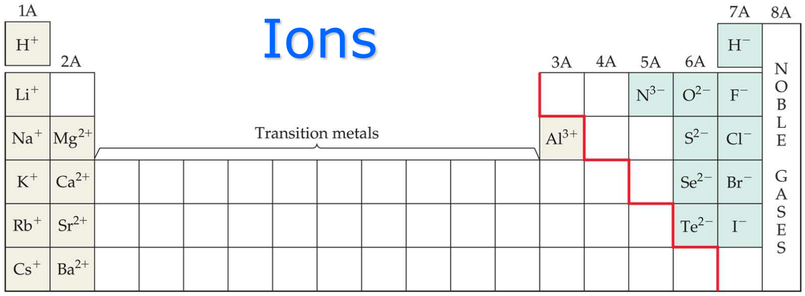
Oxygen, O_2

Types of Formulas

- Molecular formulas give the exact number of atoms of each element in a compound.
- Empirical formulas give the lowest whole-number ratio of atoms of each element in a compound.
- Structural formulas show the order in which atoms are bonded.
- Perspective drawings also show the three-dimensional shape of atoms in a compound.

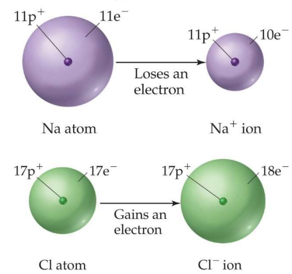


Ions



When atoms lose or gain electrons, they become ions.

- Cations are positive and are formed by elements on the left side of the periodic table.
- Anions are negative and are formed by elements on the right side of the periodic table.

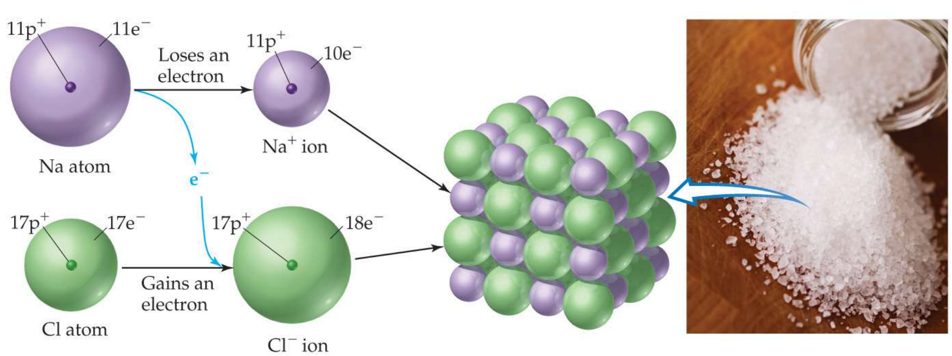


The element's symbol is followed by a superscript number and a sign that shows the charge on the ion in electron charge units.

If the ionic charge is one unit, the number is often omitted, e.g. Na^+ is the symbol for a sodium ion.

Ionic Bonds

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



Writing Formulas

$$Mg^{2+} \quad N^{3-} \longrightarrow Mg_3N_2$$

- Because compounds are electrically neutral, one can determine the formula of a binary 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.

Naming Monoatomic Ions

<u>Cations</u>

- metal + "ion" or "cation"

  Al^{3+} aluminum ion or aluminum cation

- For metals that have more than one oxidation state, the charge is indicated by Roman numerals.

  Fe^{2+} iron (II) ion, Fe^{3+} iron (III) ion

  Bi^{3+} bismuth (III) ion, Bi^{5+} bismuth (V) ion

<u>Anions</u>

- **-ide** ending

  O^{2-} oxide ion, I^{-} iodide ion, H^{-} hydride ion

TABLE 2.4 Common Cations[a]

Charge	Formula	Name	Formula	Name
1+	**H^+**	**hydrogen ion**	**NH_4^+**	**ammonium ion**
	Li^+	lithium ion	Cu^+	copper(I) or cuprous ion
	Na^+	**sodium ion**		
	K^+	**potassium ion**		
	Cs^+	cesium ion		
	Ag^+	**silver ion**		
2+	**Mg^{2+}**	**magnesium ion**	Co^{2+}	cobalt(II) or cobaltous ion
	Ca^{2+}	**calcium ion**	**Cu^{2+}**	**copper(II)** or cupric ion
	Sr^{2+}	strontium ion	**Fe^{2+}**	**iron(II)** or ferrous ion
	Ba^{2+}	barium ion	Mn^{2+}	manganese(II) or manganous ion
	Zn^{2+}	**zinc ion**	Hg_2^{2+}	mercury(I) or mercurous ion
	Cd^{2+}	cadmium ion	Hg^{2+}	mercury(II) or mercuric ion
			Ni^{2+}	nickel(II) or nickelous ion
			Pb^{2+}	**lead(II)** or plumbous ion
			Sn^{2+}	tin(II) or stannous ion
3+	**Al^{3+}**	**aluminum ion**	Cr^{3+}	chromium(III) or chromic ion
			Fe^{3+}	**iron(III)** or ferric ion

[a]The ions we use most often in this course are in boldface. Learn them first.

Fe^{2+}	iron(II) ion	Cu^+	copper(I) ion
Fe^{3+}	iron(III) ion	Cu^{2+}	copper(II) ion

Fe^{2+}	ferrous ion	Cu^+	cuprous ion
Fe^{3+}	ferric ion	Cu^{2+}	cupric ion

TABLE 2.5 Common Anions[a]

Charge	Formula	Name	Formula	Name
1−	H^-	hydride ion	**CH_3COO^-** (or $C_2H_3O_2^-$)	**acetate ion**
	F^-	**fluoride ion**	ClO_3^-	chlorate ion
	Cl^-	**chloride ion**	**ClO_4^-**	**perchlorate ion**
	Br^-	**bromide ion**	**NO_3^-**	**nitrate ion**
	I^-	**iodide ion**	MnO_4^-	permanganate ion
	CN^-	cyanide ion		
	OH^-	**hydroxide ion**		
2−	**O^{2-}**	**oxide ion**	**CO_3^{2-}**	**carbonate ion**
	O_2^{2-}	peroxide ion	CrO_4^{2-}	chromate ion
	S^{2-}	**sulfide ion**	$Cr_2O_7^{2-}$	dichromate ion
			SO_4^{2-}	**sulfate ion**
3−	N^{3-}	nitride ion	**PO_4^{3-}**	**phosphate ion**

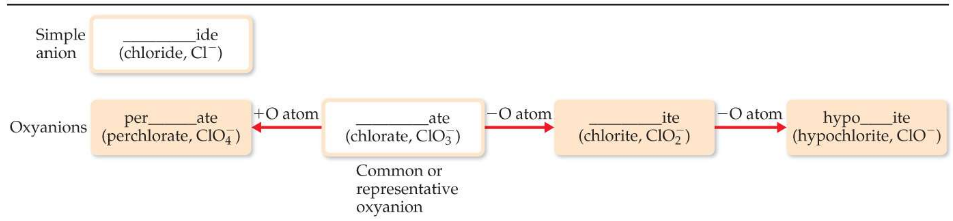
[a]The ions we use most often are in boldface. Learn them first.

H^-	hydride ion	O^{2-}	oxide ion	N^{3-}	nitride ion
OH^-	hydroxide ion	CN^-	cyanide ion	O_2^{2-}	peroxide ion

NO_3^-	nitrate ion	SO_4^{2-}	sulfate ion
NO_2^-	nitrite ion	SO_3^{2-}	sulfite ion

CO_3^{2-}	carbonate ion	PO_4^{3-}	phosphate ion
HCO_3^-	hydrogen carbonate ion	$H_2PO_4^-$	dihydrogen phosphate ion

ClO_4^-	perchlorate ion (one more O atom than chlorate)
ClO_3^-	chlorate ion
ClO_2^-	chlorite ion (one O atom fewer than chlorate)
ClO^-	hypochlorite ion (one O atom fewer than chlorite)

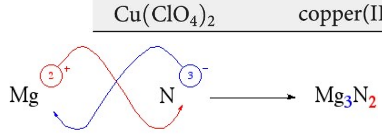


Nomenclature: Ionic Compounds

- 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 (e.g. iron), write the charge as a Roman numeral in parentheses.

Names of ionic compounds consist of the cation name followed by the anion name:

$CaCl_2$	calcium chloride
$Al(NO_3)_3$	aluminum nitrate
$Cu(ClO_4)_2$	copper(II) perchlorate (or cupric perchlorate)



First, write the name of the cation.

magnesium

Second, if the anion is an element, change its ending to –*ide.*

nitrogen → nitride

magnesium nitride

Chromium(III) oxide, used as a green paint pigment, is composed of Cr^{3+} and O^{2-} ions.

What is the formula of chromium(III) oxide?

Strontium oxide, is composed of Sr^{2+} and O^{2-} ions.

What is the formula for this compound?

Potassium chromate, is a brightly colored compound of chromium. Chromium comes from the Greek word *chroma*, meaning color. The chromate anion is CrO_4^{2-}.

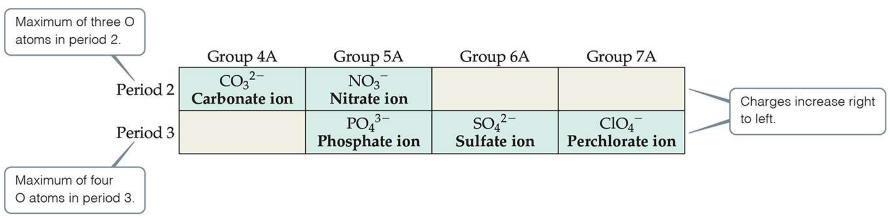
What is the ionic formula for this compound?

Iron(II) phosphate is found in the hydrated form in the mineral Vivianite.

What is the formula for this compound?

Patterns in Oxyanion Nomenclature

- When there are ***two*** oxyanions involving the same element:
  - The one with fewer oxygens ends in -*ite*
    - NO_2^- : nitrite; SO_3^{2-} : sulfite
  - The one with more oxygens ends in -*ate*
    - NO_3^- : nitrate; SO_4^{2-} : sulfate

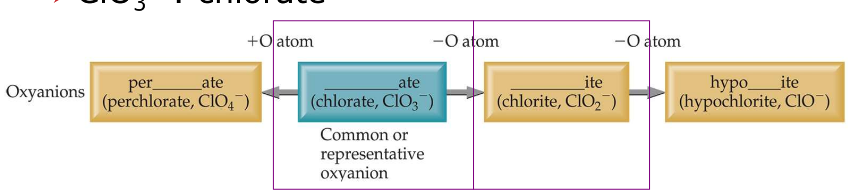


Most common form of the oxyanion.

Patterns in Oxyanion Nomenclature

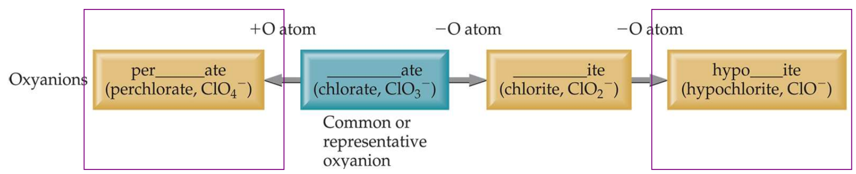
- When there are ***more than two*** oxyanions involving the same element:

- The one with the second *fewest* oxygens ends in -*ite*
  - ClO_2^- : chlorite
- The one with the second *most* oxygens ends in -*ate*
  - ClO_3^- : chlorate

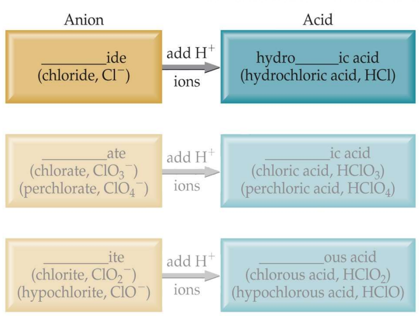


Patterns in Oxyanion Nomenclature

- The one with the fewest oxygens has the prefix *hypo*- and ends in -*ite*
  - ➢ClO^- : hypochlorite
- The one with the most oxygens has the prefix *per*- and ends in -*ate*
  - ➢ClO_4^- : perchlorate



Acid Nomenclature

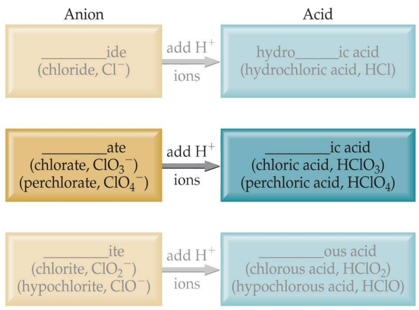


- If the anion in the acid ends in -*ide*, change the ending to -*ic acid* and add the prefix *hydro*- :
  - HCl: hydrochloric acid
  - HBr: hydrobromic acid
  - HI: hydroiodic acid

Name H_2S

Anion	Corresponding Acid
Cl^- (chloride)	HCl (hydrochloric acid)
S^{2-} (sulfide)	H_2S (hydrosulfuric acid)

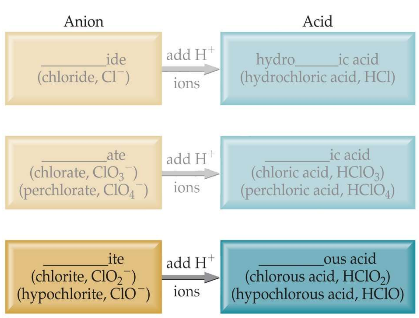
Acid Nomenclature



- If the anion in the acid ends in -*ate*, change the ending to -*ic acid:*
  - $HClO_3$: chloric acid
  - $HClO_4$: perchloric acid

Name the corresponding acid to the SO_4^{2-} anion.

Acid Nomenclature



- If the anion in the acid ends in -*ite*, change the ending to -*ous acid*:
  - HClO: hypochlorous acid
  - $HClO_2$: chlorous acid

Name the corresponding acid to the SO_3^{2-} anion.

Anion		Corresponding Acid	
ClO_4^-	(perchlorate)	$HClO_4$	(perchloric acid)
ClO_3^-	(chlorate)	$HClO_3$	(chloric acid)
ClO_2^-	(chlorite)	$HClO_2$	(chlorous acid)
ClO^-	(hypochlorite)	$HClO$	(hypochlorous acid)

TABLE 2.6 Prefixes Used in Naming Binary Compounds Formed between Nonmetals

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

Nomenclature of Binary Compounds (molecular) Formed Between *Nonmetals*.

- The less electronegative atom (farthest to left) is usually listed first.
- A Greek 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.)
- The ending on the more electronegative element is changed to -*ide*.
  - CCl_4: carbon tetrachloride
  - CO_2: carbon dioxide
  - CO: carbon monoxide
  - N_2O_4: dinitrogen tetroxide
- In nonmetal compounds containing carbon, the C atom is listed first in the chemical formula, followed by H.

TABLE 2.6 Prefixes Used in Naming Binary Compounds Formed between Nonmetals

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

N_2O_5: dinitrogen pentoxide

Cl_2O	dichlorine monoxide	NF_3	nitrogen trifluoride
N_2O_4	dinitrogen tetroxide	P_4S_{10}	tetraphosphorus decasulfide

Summary of Nomenclature

Cations

1. Cations formed from metal atoms have the same name as the metal
2. If a metal can form different cations, the positive charge is indicated by a Roman numeral in parenthesis following the name of the metal.
    - Sometimes the ending *–ous* and *–ic* will be used to distinguish lower and higher charged ions, respectively.
3. Cations formed from nonmetal atoms have names the end in *–ium*.

Anions

1. Names of monatomic anions formed by replacing the ending of the element with *–ide*.
2. Polyatomic ions (oxyanions) containing oxygen have names that end in *–ate* or *–ite*.
3. Anions derived by adding H^+ to an oxyanion are named by adding as a prefix the word hydrogen or dihydrogen, as appropriate.