

Chapter 2

Atoms, Molecules, and Ions

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

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

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.

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).

Thompson measured the charge/mass ratio of the electron to be 1.76×10^8 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.

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.
- Three types of radiation were discovered by Ernest Rutherford:
 - α particles
 - β particles
 - γ rays

The Atom, circa 1900:

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

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.

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

- 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-)	5.486×10^{-4}

Symbols of Elements

Mass number (number of protons plus neutrons)

Atomic number (number of protons or electrons)



← Symbol of element

Elements are symbolized by one or two letters.

Atomic Number

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

The atomic number (Z)

Atomic Mass

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.

Isotopes of Hydrogen

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.

1A 1																	8A 18
1 1 H	2A 2											3A 13	4A 14	5A 15	6A 16	7A 17	2 He
2 3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3 11 Na	12 Mg	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8 9 10			1B 11	2B 12	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4 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
5 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
6 55 Cs	56 Ba	71 Lu	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
7 87 Fr	88 Ra	103 Lr	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112	113	114	115	116		

Metals	57 La	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
Metalloids	89 Ac	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
Nonmetals														

- 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	Nonreactive gas	Soft, reactive metal	Nonreactive gas	Soft, reactive metal	Nonreactive gas	Nonreactive gas	Soft, reactive metal	Nonreactive gas	Soft, reactive metal	Nonreactive gas	Nonreactive gas	Soft, reactive metal	Nonreactive gas

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

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

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.

Nonmetals are on the right side of the periodic table (with the exception of Al and Po).
 Metalloids border the stair-step line (with the exception of Al and Po).
 Metals are on the left side of the chart.

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 are composed of molecules and almost always contain only nonmetals.

These seven elements occur naturally as molecules containing two atoms.

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

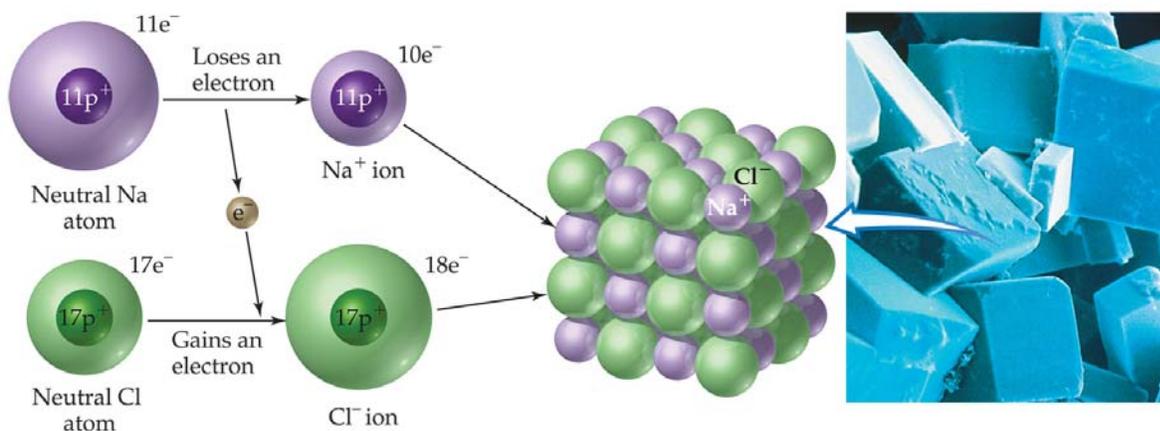
Ions:

When atoms lose or gain electrons, they become ions.

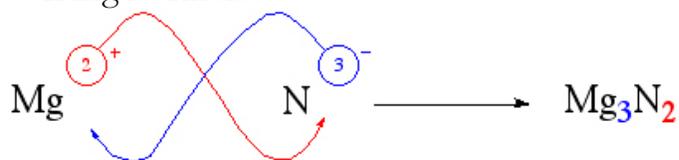
- Cations are positive and are formed by elements on the left side of the periodic chart.
- Anions are negative and are formed by elements on the right side of the periodic chart.

1A	2A	Transition metals										3A	4A	5A	6A	7A	8A	
H ⁺															N ³⁻	O ²⁻	H ⁻	
Li ⁺													Al ³⁺			S ²⁻	Cl ⁻	
Na ⁺	Mg ²⁺															Se ²⁻	Br ⁻	
K ⁺	Ca ²⁺															Te ²⁻	I ⁻	
Rb ⁺	Sr ²⁺																	
Cs ⁺	Ba ²⁺																	

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.

Charge	Formula	Name	Formula	Name
1+	H⁺	Hydrogen ion	NH₄⁺	Ammonium ion
	Li ⁺	Lithium ion	Cu ⁺	Copper(I) or cuprous ion
	Na ⁺	Sodium ion		
	K ⁺	Potassium ion		
	Cs ⁺	Cesium ion		
	Ag ⁺	Silver ion		
2+	Mg²⁺	Magnesium ion	Co ²⁺	Cobalt(II) or cobaltous ion
	Ca²⁺	Calcium ion	Cu²⁺	Copper(II) or cupric ion
	Sr ²⁺	Strontium ion	Fe²⁺	Iron(II) or ferrous ion
	Ba ²⁺	Barium ion	Mn ²⁺	Manganese(II) or manganous ion
	Zn²⁺	Zinc ion	Hg ₂ ²⁺	Mercury(I) or mercurous ion
	Cd ²⁺	Cadmium ion	Hg²⁺	Mercury(II) or mercuric ion
			Ni ²⁺	Nickel(II) or nickelous ion
			Pb²⁺	Lead(II) or plumbous ion
			Sn ²⁺	Tin(II) or stannous ion
3+	Al³⁺	Aluminum ion	Cr ³⁺	Chromium(III) or chromic ion
			Fe³⁺	Iron(III) or ferric ion

*The most common ions are in boldface.

Charge	Formula	Name	Formula	Name
1-	H ⁻	Hydride ion	C ₂ H ₃ O 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 ₂ O ₇ ²⁻	Dichromate ion
			SO ₄ ²⁻	Sulfate ion
3-	N ³⁻	Nitride ion	PO ₄ ³⁻	Phosphate ion

*The most common ions are in boldface.

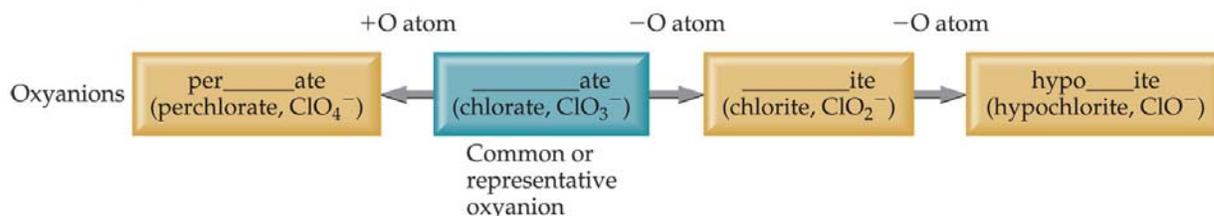
Inorganic Nomenclature

- 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.

<p>Cations:</p> <ul style="list-style-type: none"> <input type="radio"/> NH₄⁺ <input type="radio"/> Li⁺ <input type="radio"/> Na⁺ <input type="radio"/> Ca²⁺ <input type="radio"/> Ba²⁺ <input type="radio"/> Ag⁺ <input checked="" type="radio"/> Fe²⁺ <input type="radio"/> Fe³⁺ <input type="radio"/> Al³⁺ <input type="radio"/> Pb²⁺ 	<div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 5px auto;">iron(II) phosphate</div> <div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 5px auto;">Fe₃(PO₄)₂</div> <div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 5px auto;">insoluble in water</div> <p>Solubility defined as greater than 0.1 g solute / 100 g H₂O</p>	<p>Anions:</p> <ul style="list-style-type: none"> <input type="radio"/> F⁻ <input type="radio"/> Cl⁻ <input type="radio"/> Br⁻ <input type="radio"/> OH⁻ <input type="radio"/> S²⁻ <input type="radio"/> CO₃²⁻ <input type="radio"/> SO₄²⁻ <input checked="" type="radio"/> PO₄³⁻ <input type="radio"/> NO₃⁻ <input type="radio"/> ClO₄⁻
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Patterns in Oxyanion Nomenclature

- When there are two oxyanions involving the same element:
 - The one with fewer oxygens ends in *-ite*
 - NO₂⁻ : nitrite; SO₃²⁻ : sulfite
 - The one with more oxygens ends in *-ate*
 - The one with the second fewest oxygens ends in *-ite*
 - ClO₂⁻ : chlorite
 - The one with the second most oxygens ends in *-ate*
 - ClO₃⁻ : chlorate
 - NO₃⁻ : nitrate; SO₄²⁻ : sulfate

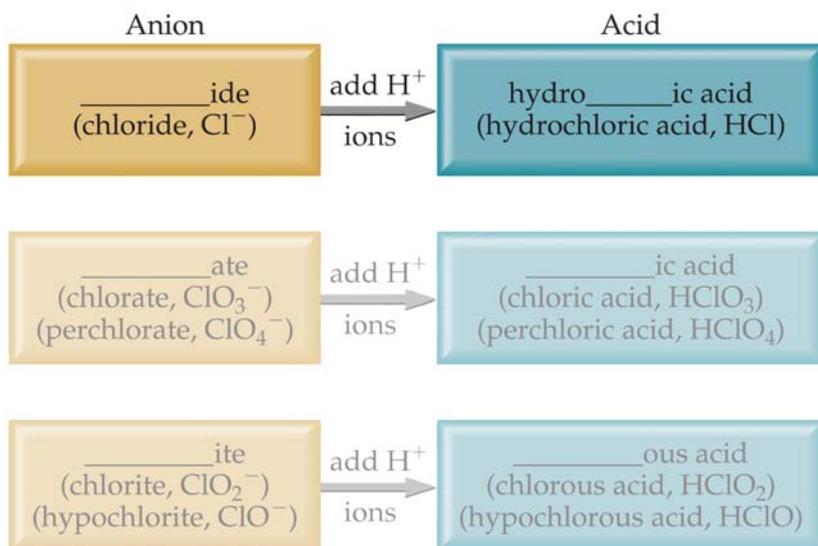


- The one with the fewest oxygens has the prefix *hypo-* and ends in *-ite*
 - ClO⁻ : hypochlorite
- The one with the most oxygens has the prefix *per-* and ends in *-ate*

➤ 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
- If the anion in the acid ends in *-ite*, change the ending to *-ous acid*:
 - HClO: hypochlorous acid
 - HClO₂: chlorous acid
- If the anion in the acid ends in *-ate*, change the ending to *-ic acid*:
 - HClO₃: chloric acid
 - HClO₄: 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.)
- The ending on the more electronegative element is changed to *-ide*.

➤ CO₂: carbon dioxide

➤ CCl₄: carbon tetrachloride

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:

N₂O₅: dinitrogen pentoxide