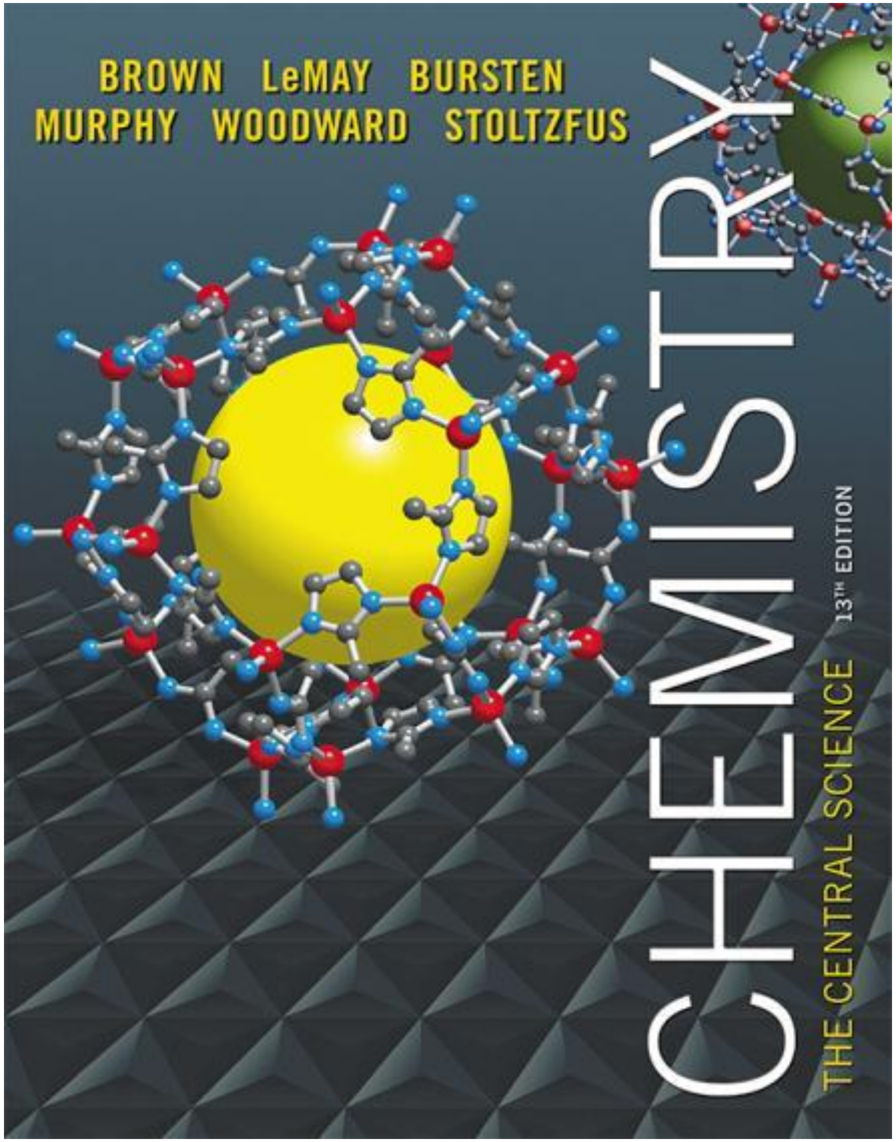

Lecture Presentation

Chapter 2

Atoms, Molecules, and Ions

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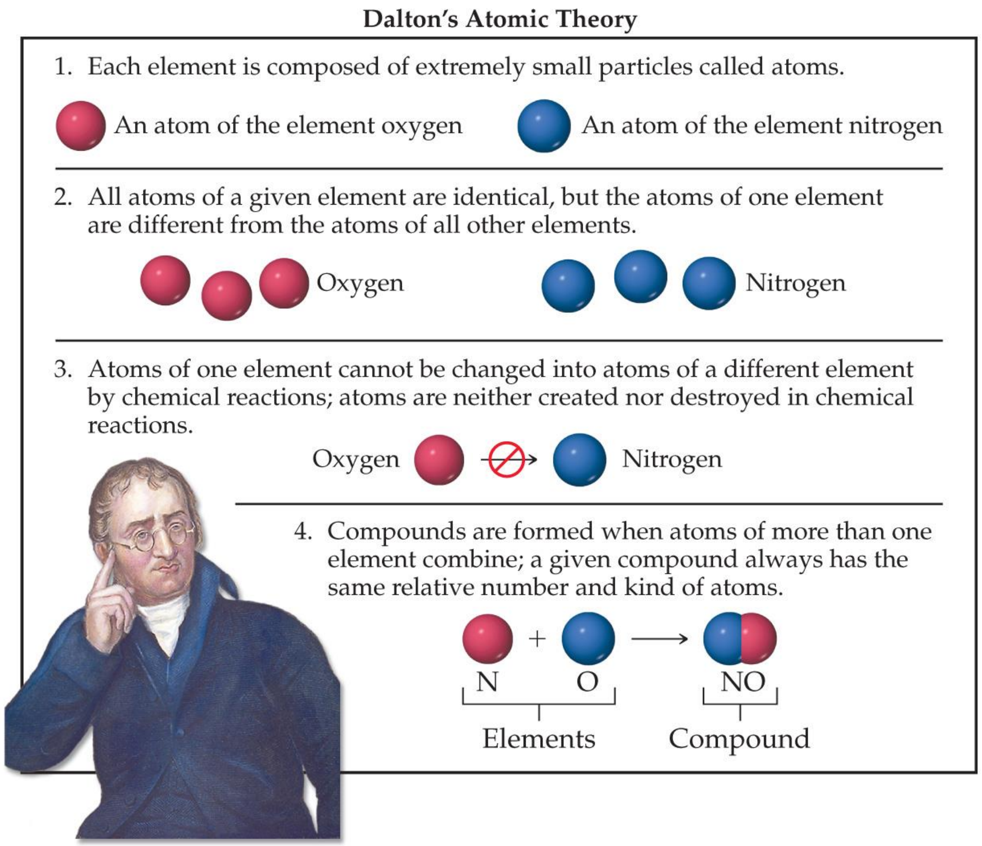
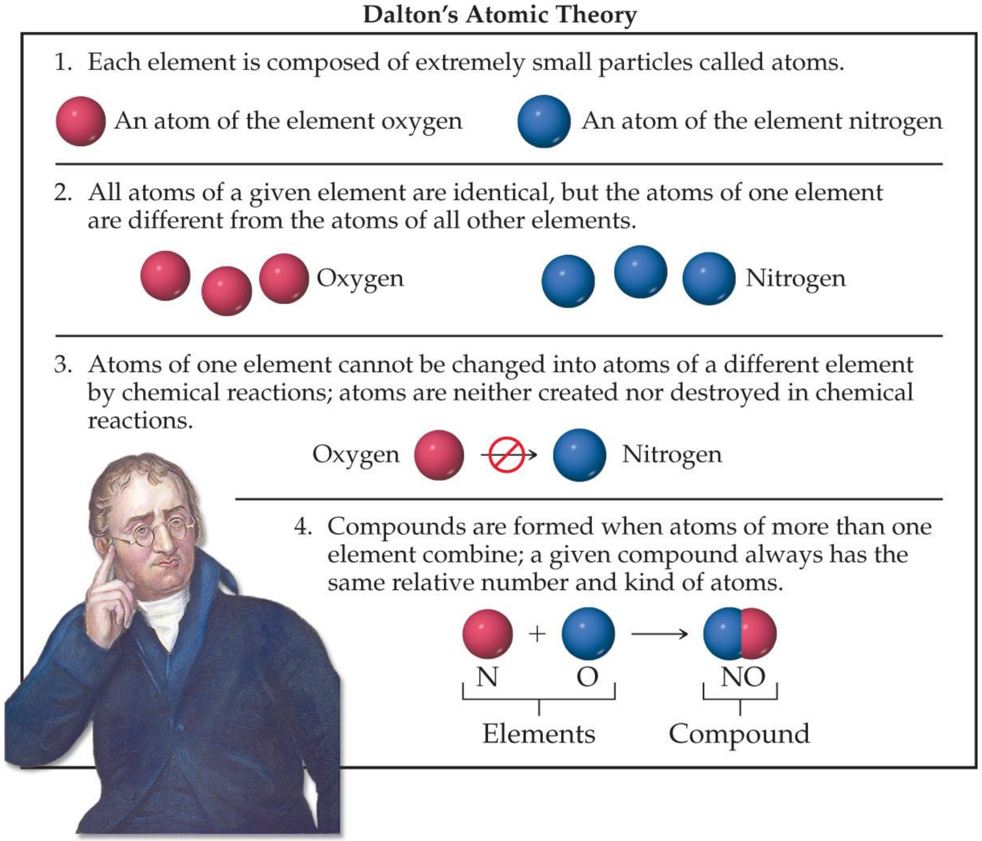
Atomic Theory of Matter



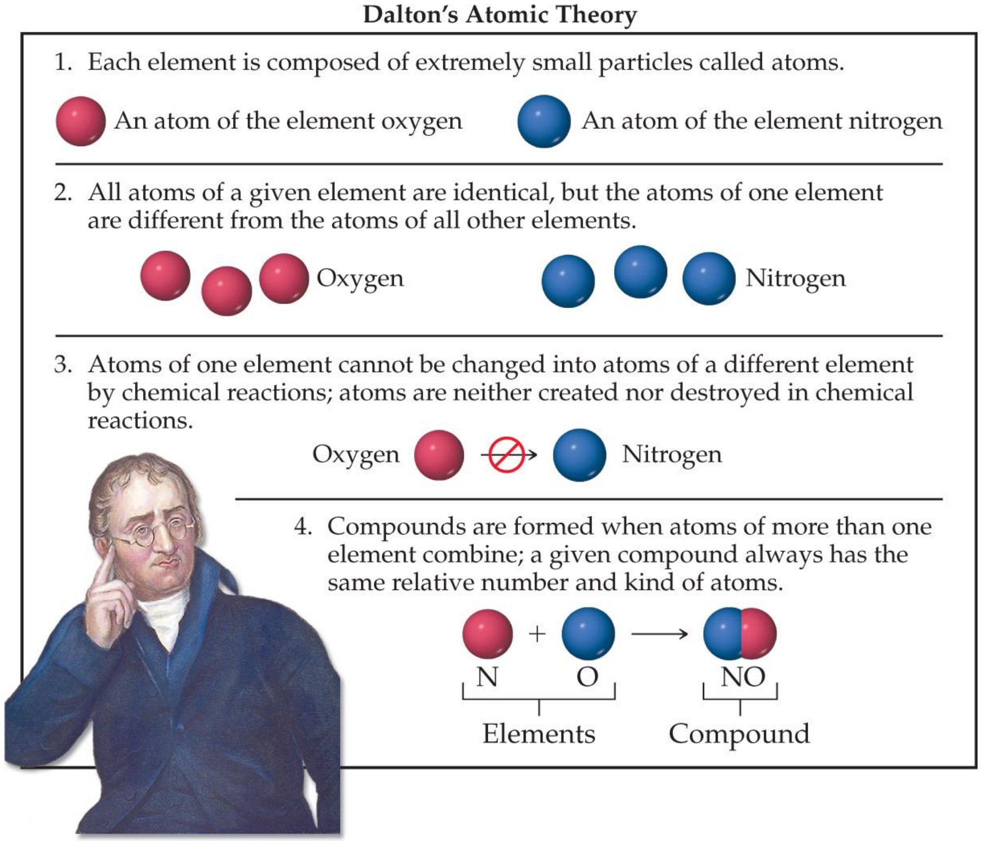
The theory that atoms are the fundamental building blocks of matter reemerged in the early nineteenth century, championed by John Dalton.

Dalton's Postulates



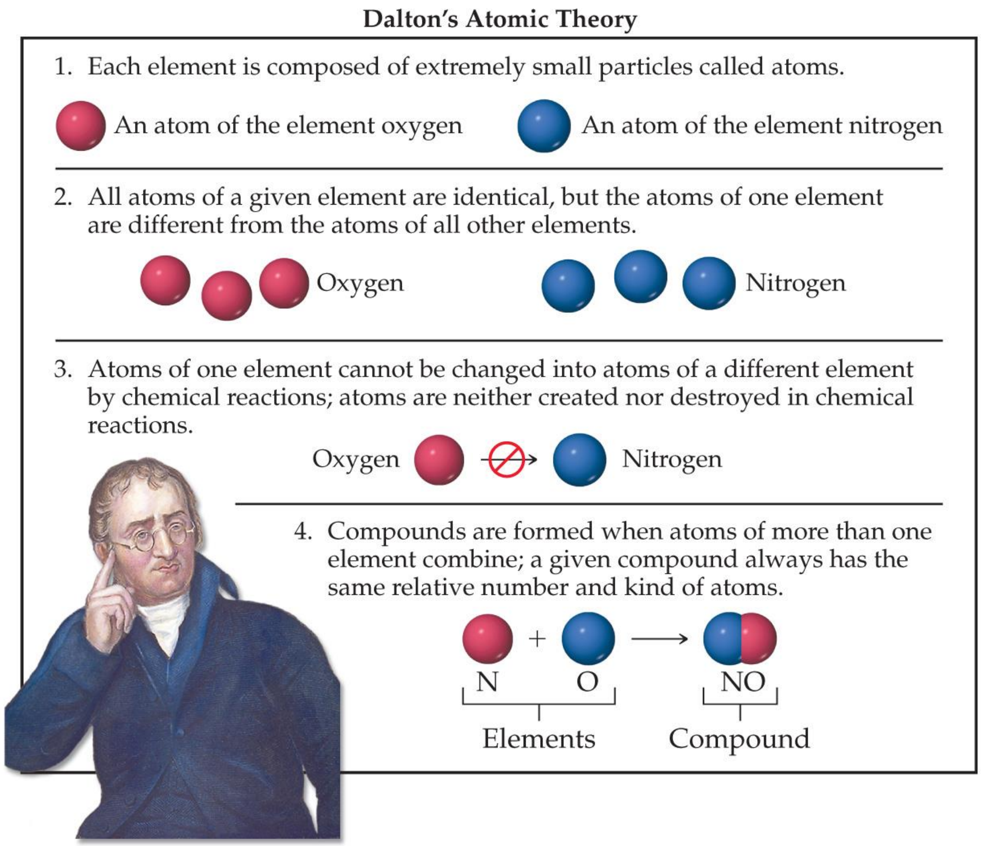
1) Each element is composed of extremely small particles called **atoms**.

Dalton’s Postulates



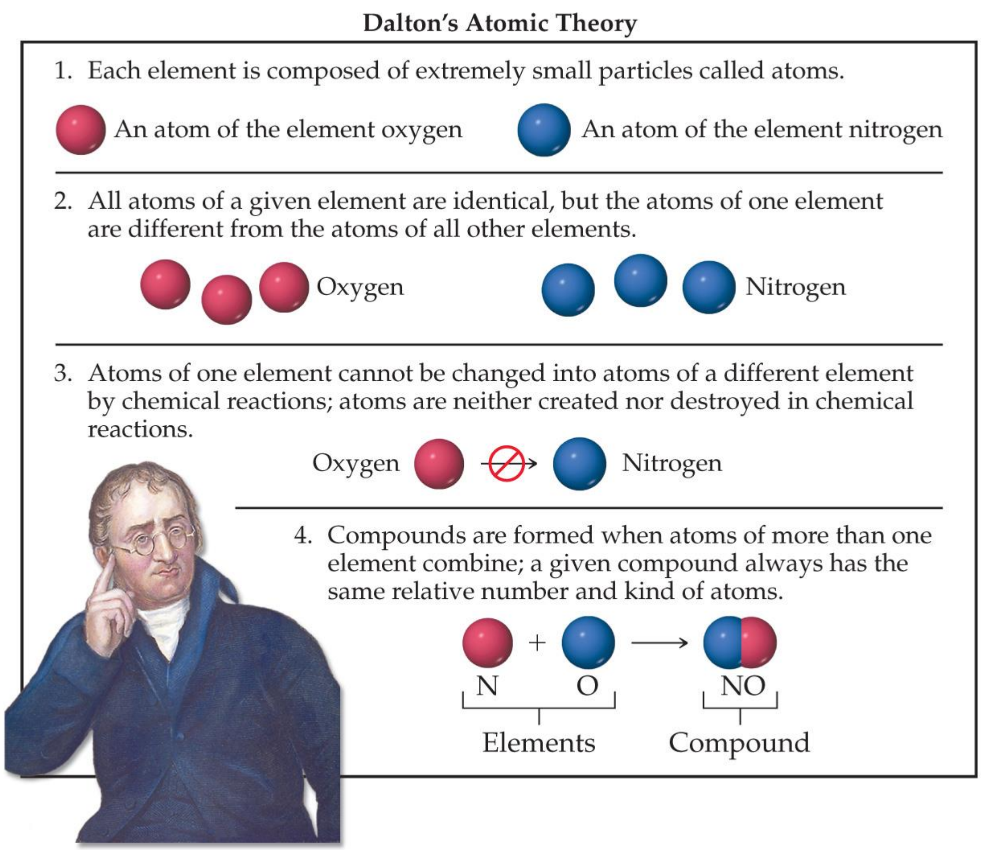
2) All atoms of a given element are identical to one another in mass and other properties, but the atoms of one element are different from the atoms of all other elements.

Dalton's Postulates



3) 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



4) Atoms of more than one element combine to form **compounds**; a given compound always has the same relative number and kind of atoms.

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.
- This law was one of the laws on which Dalton's atomic theory was based.

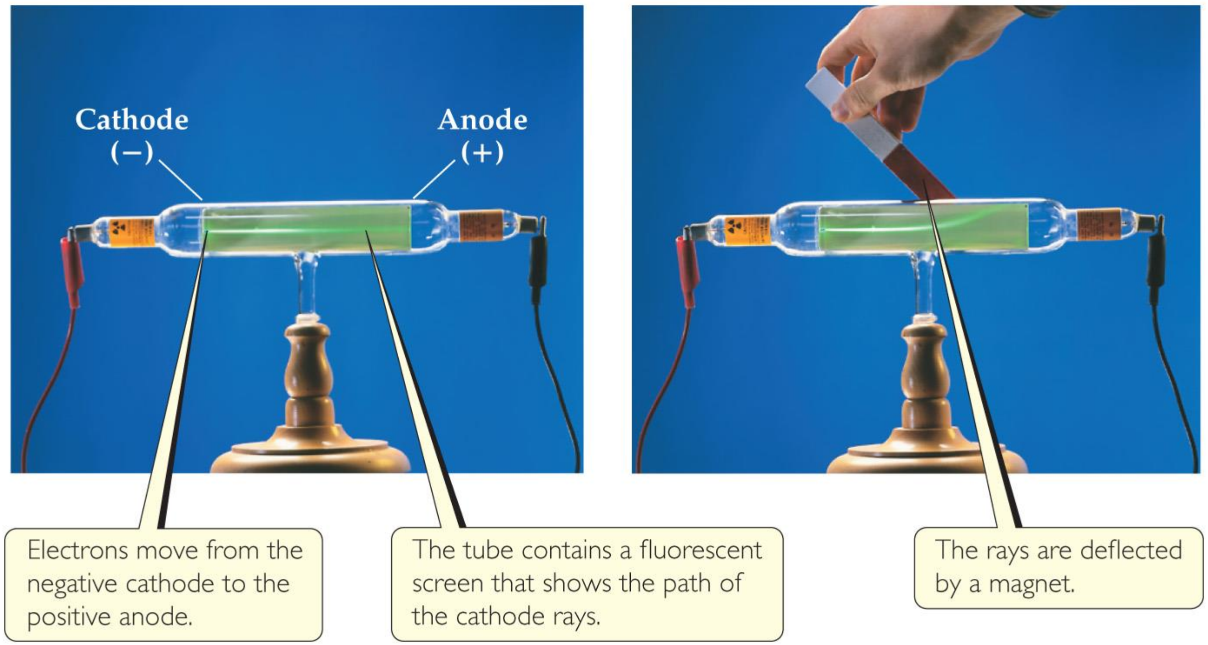
Law of Multiple Proportions

- If two elements, A and B, form more than one compound, the masses of B that combine with a given mass of A are in the ratio of small whole numbers.
- Dalton predicted this law and observed it while developing his atomic theory.
- When two or more compounds exist from the same elements, they can *not* have the same relative number of atoms.

Discovery of Subatomic Particles

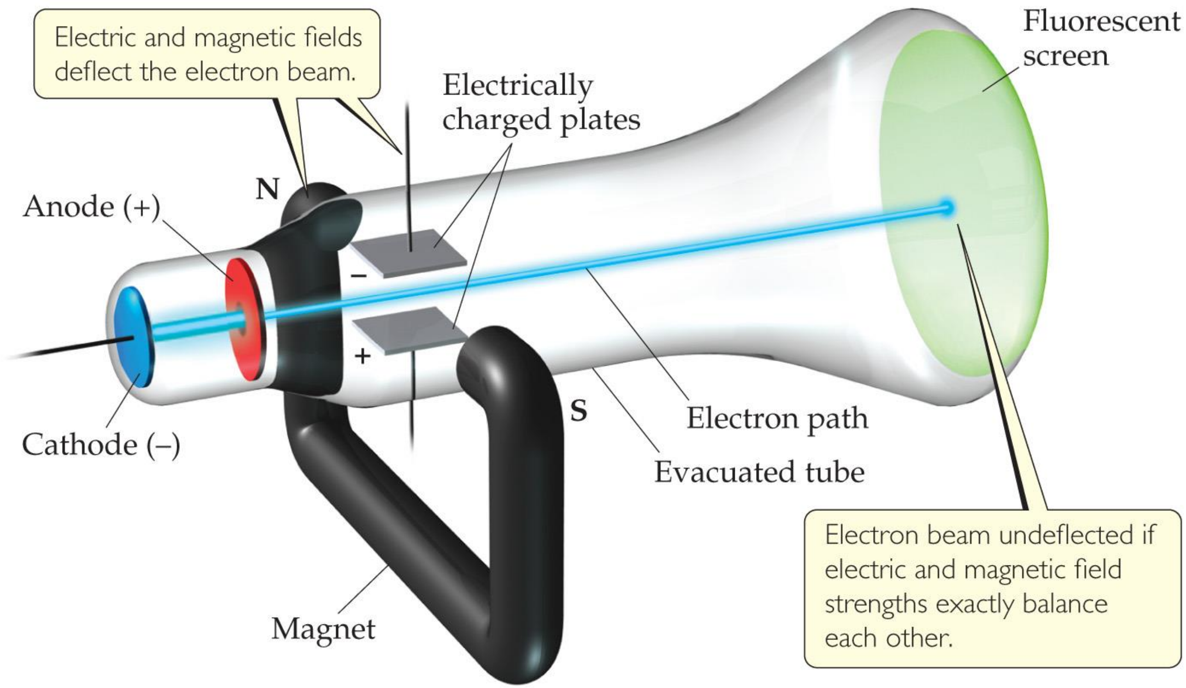
- In Dalton’s view, the atom was the smallest particle possible. Many discoveries led to the fact that the atom itself was made up of smaller particles.
  - Electrons and cathode rays
  - Radioactivity
  - Nucleus, protons, and neutrons

The Electron (Cathode Rays)



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

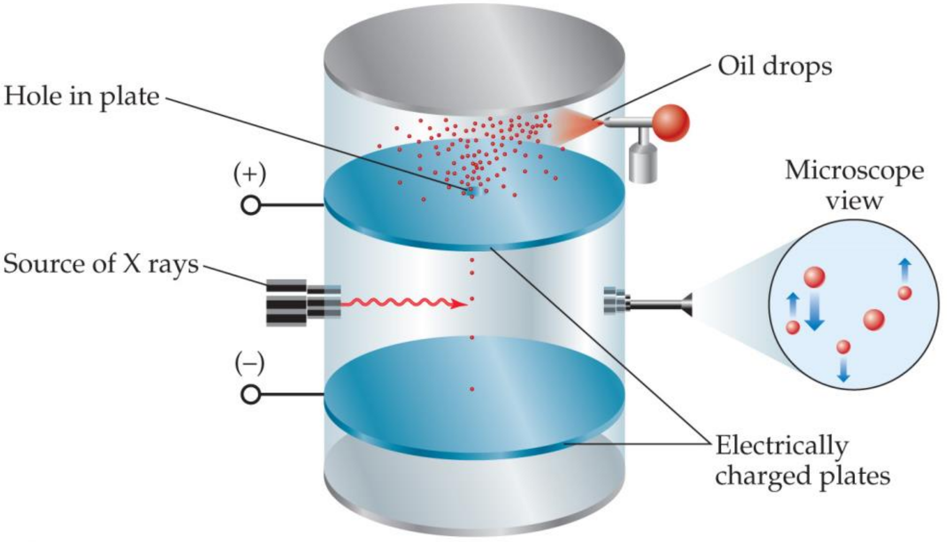
The Electron



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

Millikan Oil-Drop Experiment (Electrons)

- 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 determined the charge on the electron in 1909.

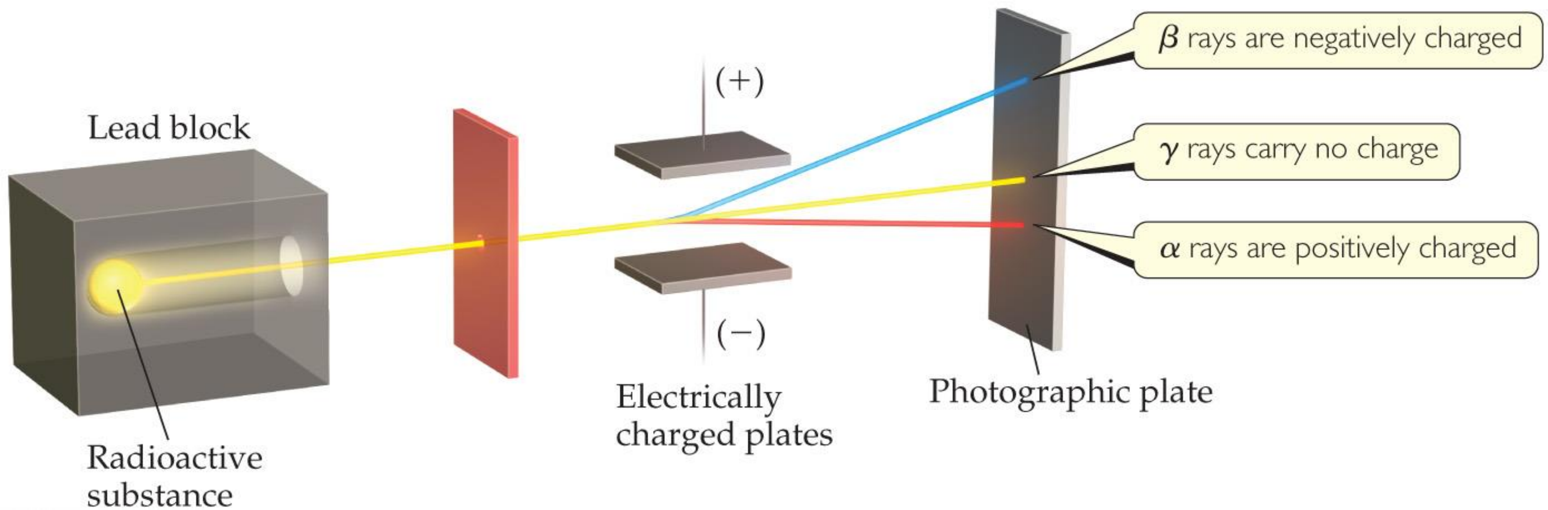


Radioactivity

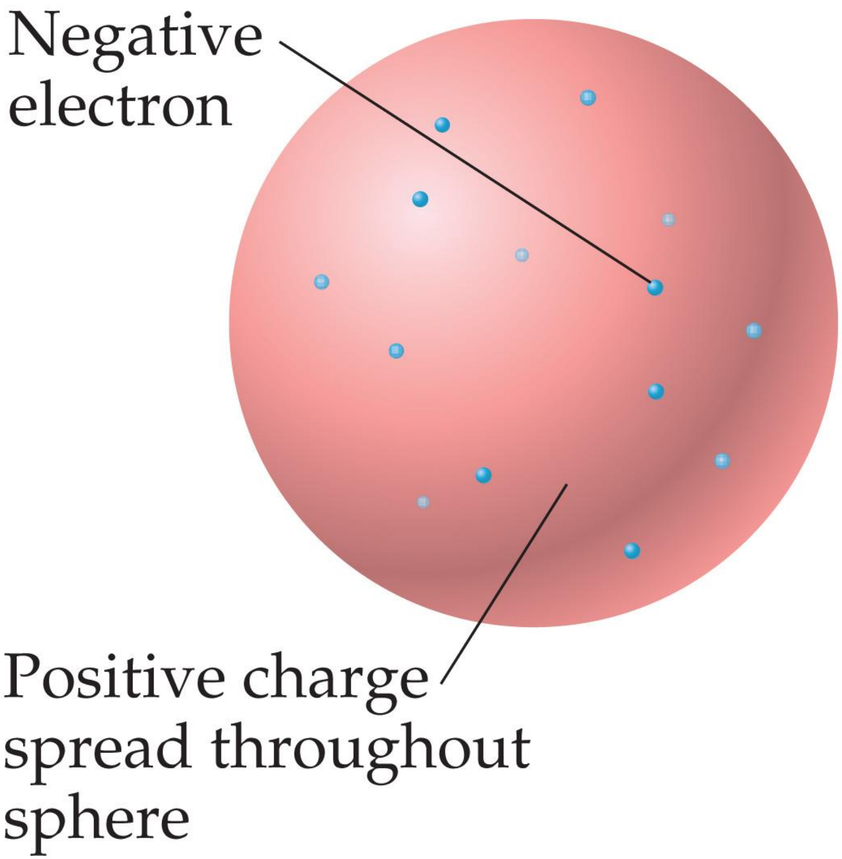
- **Radioactivity** is the spontaneous emission of high-energy radiation by an atom.
- It was first observed by Henri Becquerel.
- Marie and Pierre Curie also studied it.
- Its discovery showed that the atom had more subatomic particles and energy associated with it.

Radioactivity

- Three types of radiation were discovered by Ernest Rutherford:
  - α particles (positively charged)
  - β particles (negatively charged, like electrons)
  - γ rays (uncharged)

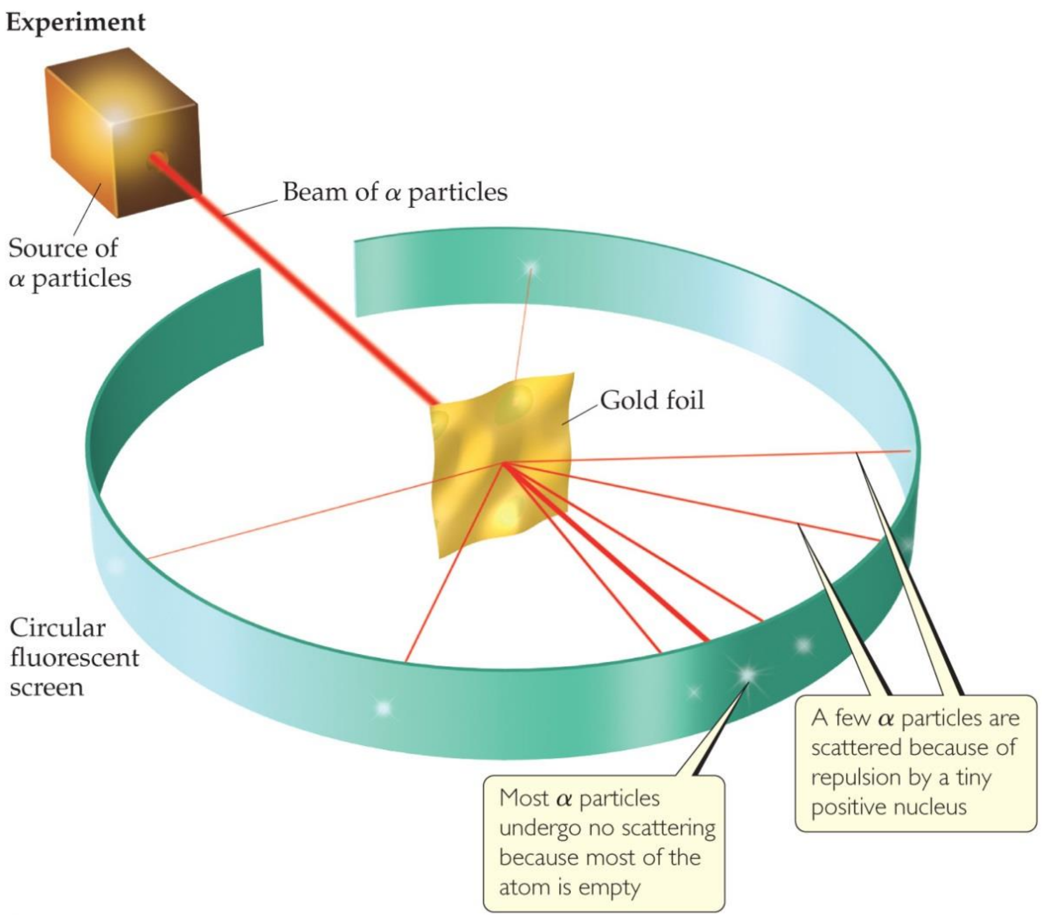


The Atom, circa 1900



- The prevailing theory was that of the “plum pudding” model, put forward by Thomson.
- It featured a positive sphere of matter with negative electrons embedded in it.

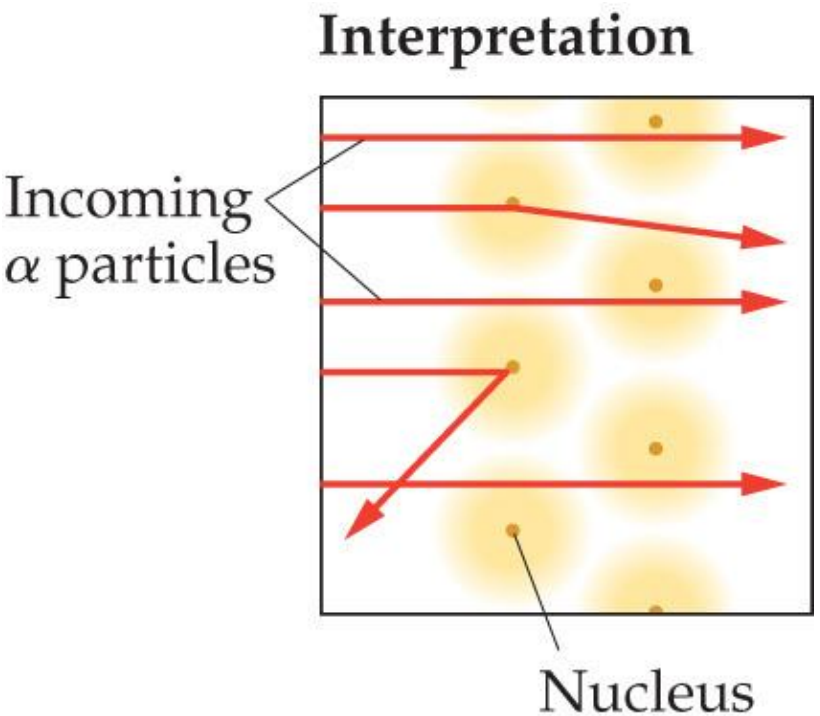
Discovery of the Nucleus



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

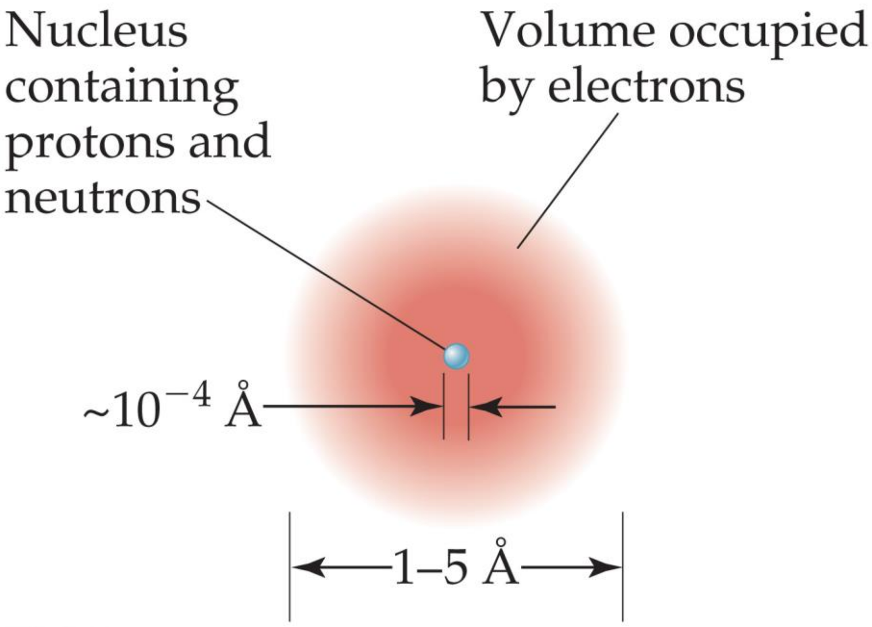
The Nuclear Atom

Since some particles were deflected at large angles, Thomson'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 is empty space.
- Atoms are very small; 1 – 5 Å or 100 – 500 pm.
- Other subatomic particles (protons and neutrons) were discovered.



Subatomic Particles

- Protons (+1) and electrons (–1) have a charge; neutrons are neutral.
- Protons and neutrons have essentially the same mass (relative mass 1). The mass of an electron is so small we ignore it (relative mass 0).
- Protons and neutrons are found in the nucleus; electrons travel around the nucleus.

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}

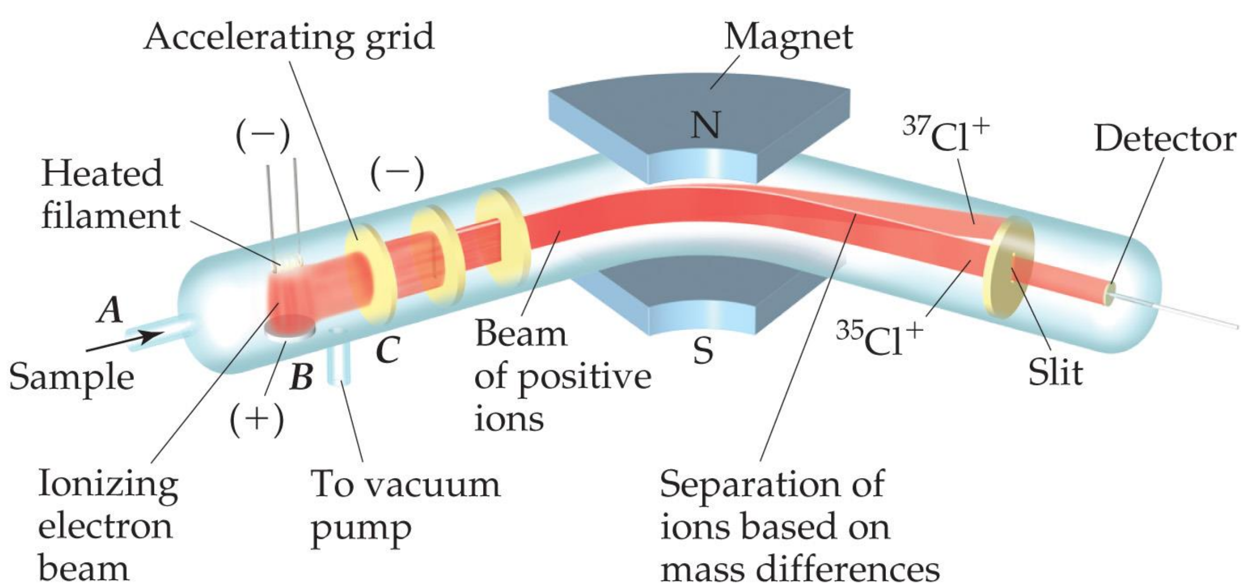
Atomic Mass

- Atoms have extremely small masses.
- The heaviest known atoms have a mass of approximately 4×10^{-22} g.
- A mass scale on the atomic level is used, where an atomic mass unit (amu) is the base unit.

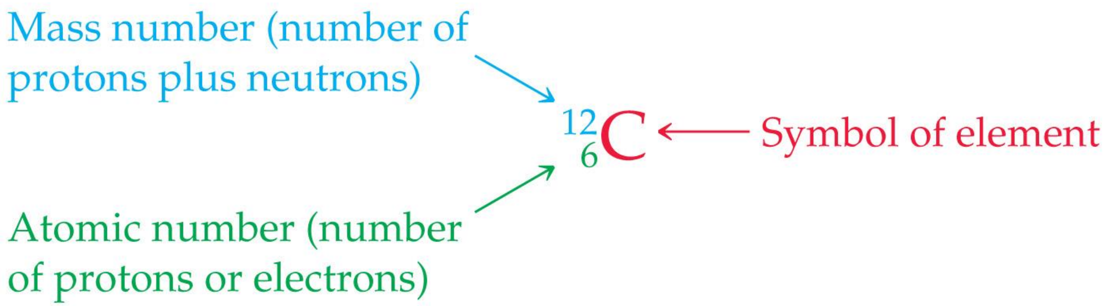
➢ $1 \text{ amu} = 1.66054 \times 10^{-24}$ g

Atomic Weight Measurement

- Atomic and molecular weight can be measured with great accuracy using a mass spectrometer.
- Masses of atoms are compared to the carbon atom with 6 protons and 6 neutrons (C-12).



Symbols of Elements



- Elements are represented by a one or two letter **symbol**. This is the symbol for carbon.
- All atoms of the same element have the same number of protons, which is called the **atomic number**, *Z*. It is written as a subscript BEFORE the symbol.
- The **mass number** is the total number of protons and neutrons in the nucleus of an atom. It is written as a superscript BEFORE the symbol.

Isotopes

- **Isotopes** are atoms of the same element with different masses.
- Isotopes have different numbers of neutrons, but the same number of protons.

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

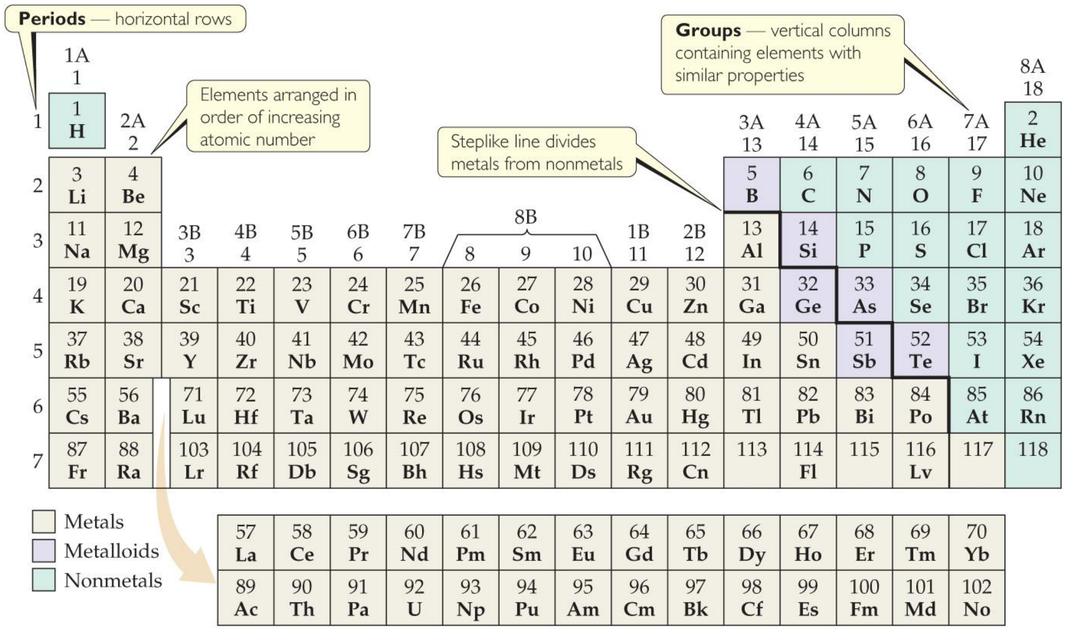
[a]Almost 99% of the carbon found in nature is ^{12}C.

Atomic Weight

- Because in the real world we use large amounts of atoms and molecules, we use average masses in calculations.
- An average mass is found using all isotopes of an element weighted by their relative abundances. This is the element's **atomic weight**.
- That is, Atomic Weight = Σ [(isotope mass) $\times$ (fractional natural abundance)]. Note: the sum is for ALL isotopes of an element.

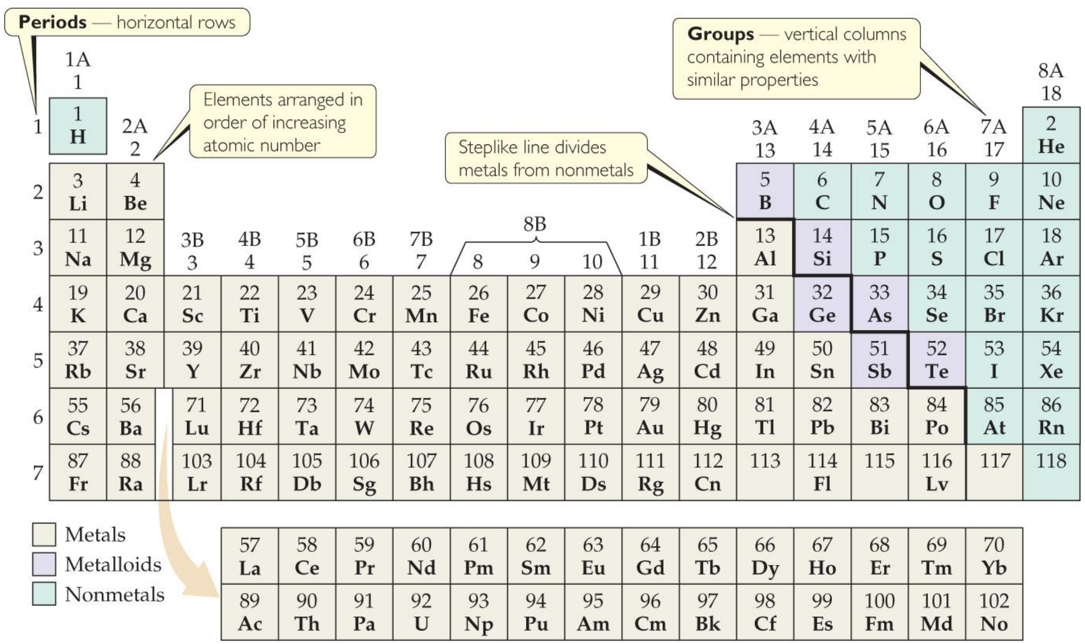
Periodic Table

- The **periodic table** is a systematic organization of the elements.
- Elements are arranged in order of atomic number.



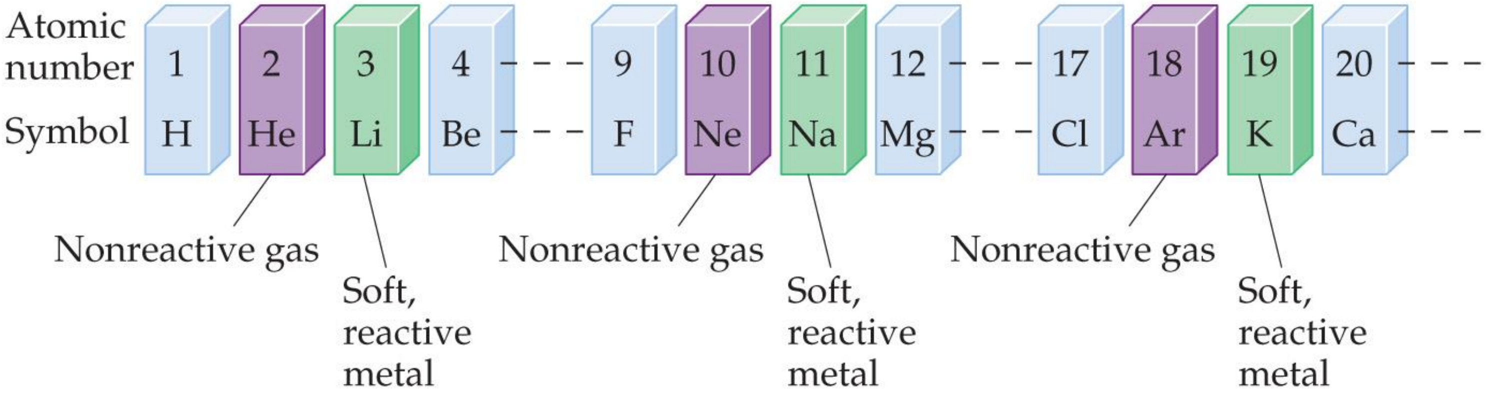
- Unlike the way we write isotopes, the atomic number is at the TOP of a box in the periodic table.
- The atomic weight of an element appears at the BOTTOM of the box. (They are not shown on this version of the Periodic Table.)

Periodic Table



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

Periodicity



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

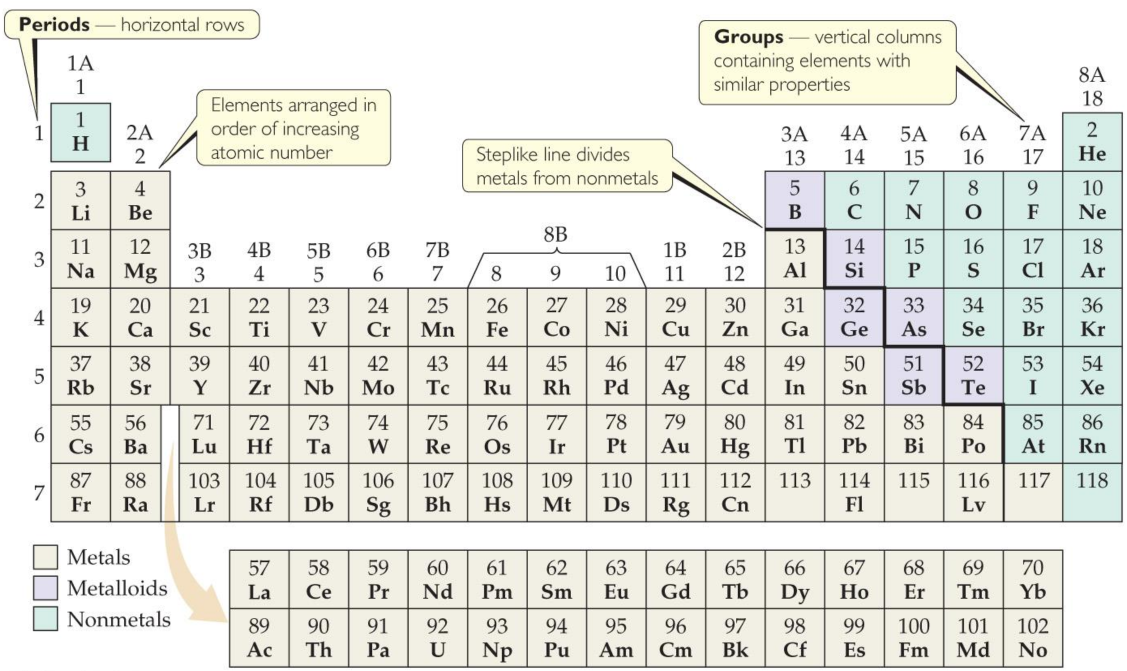
Groups

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 (or rare gases)	He, Ne, Ar, Kr, Xe, Rn

These five groups are known by their names.

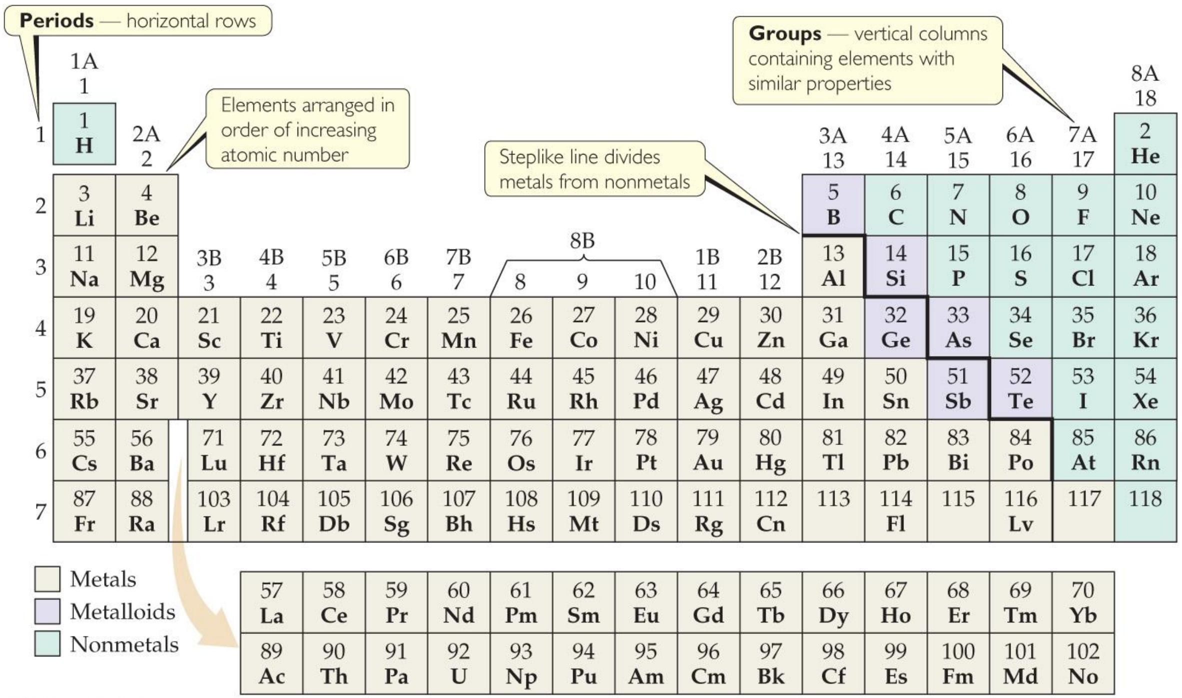
Periodic Table



- Metals are on the left side of the periodic table.
- Some properties of metals include
  - shiny luster.
  - conducting heat and electricity.
  - solidity (except mercury).

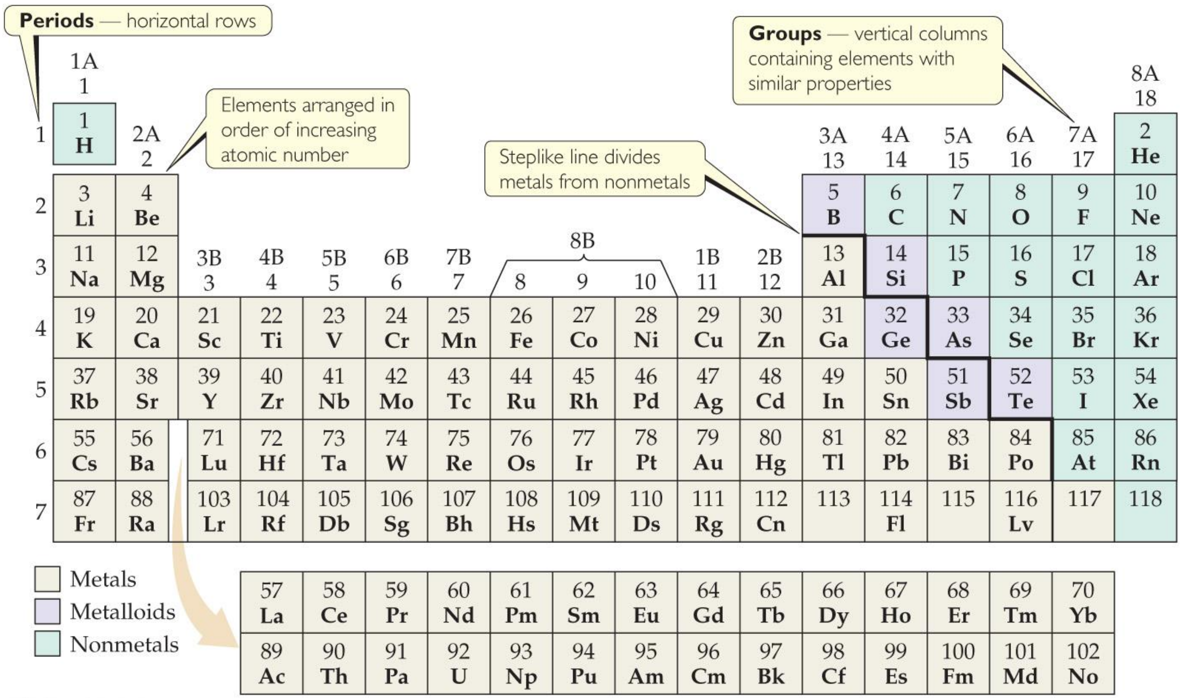
Atoms, Molecules, and Ions

Periodic Table



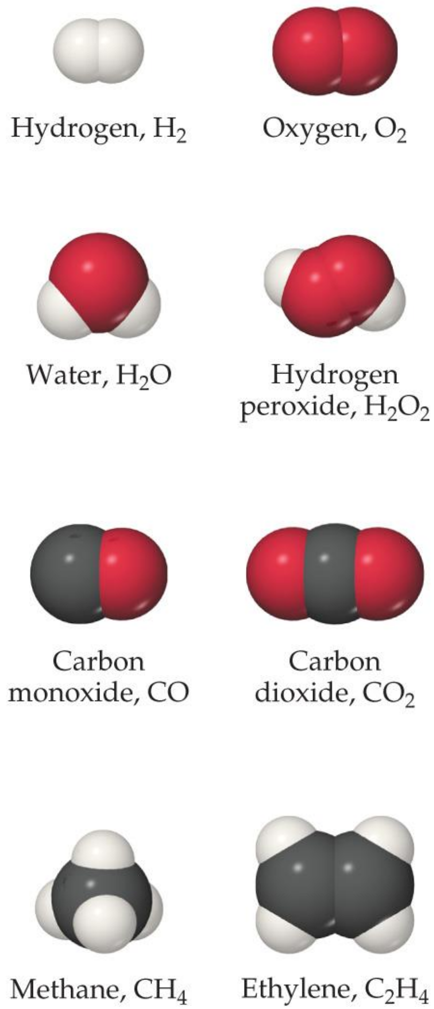
- Nonmetals are on the right side of the periodic table (with the exception of H).
- They can be solid (like carbon), liquid (like bromine), or gas (like neon) at room temperature.

Periodic Table



- Elements on the steplike line are metalloids (except Al, Po, and At).
- Their properties are sometimes like metals and sometimes like nonmetals.

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** are composed of molecules and almost always contain only nonmetals.

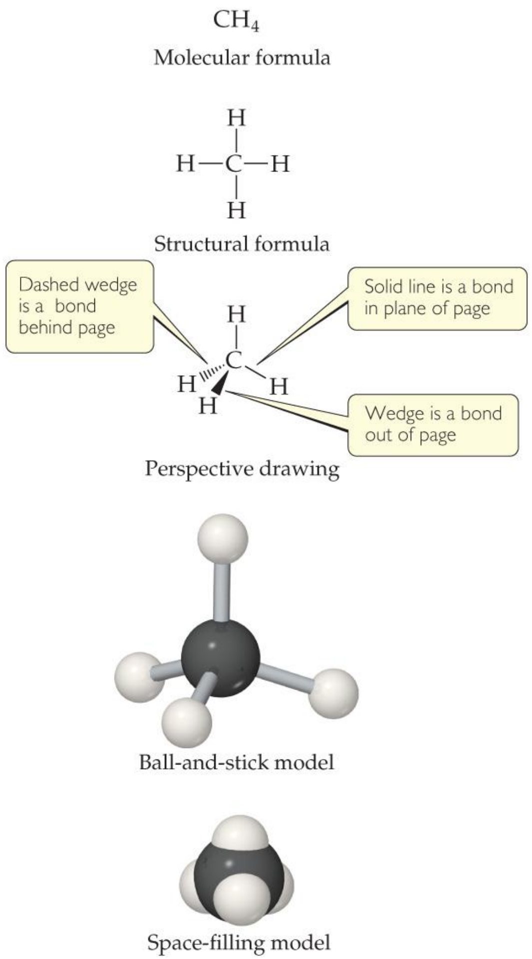
Diatomic Molecules

- These seven elements occur naturally as molecules containing two atoms:
  - Hydrogen
  - Nitrogen
  - Oxygen
  - Fluorine
  - Chlorine
  - Bromine
  - Iodine

Types of Formulas

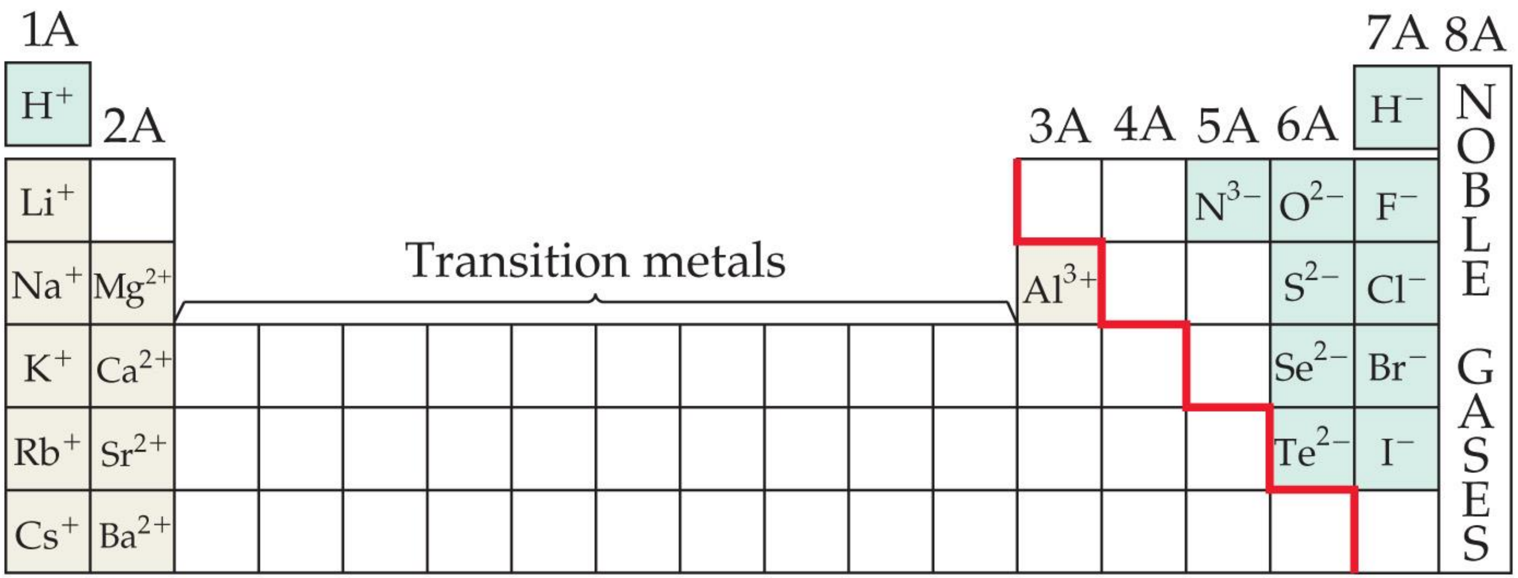
- **Empirical formulas** give the lowest whole-number ratio of atoms of each element in a compound.
- **Molecular formulas** give the exact number of atoms of each element in a compound.
- If we know the molecular formula of a compound, we can determine its empirical formula. The converse is not true!

Types of Formulas



- **Structural formulas** show the order in which atoms are attached. They do NOT depict the three-dimensional shape of molecules.
- **Perspective drawings** also show the three-dimensional order of the atoms in a compound. These are also demonstrated using **models**.

Ions



- When an atom of a group of atoms loses or gains electrons, it becomes an **ion**.
- **Cations** are formed when at least one electron is lost. Monatomic cations are formed by metals.
- **Anions** are formed when at least one electron is gained. Monatomic anions are formed by nonmetals.

Common Cations

Table 2.4 Common Cations[a]

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

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

Common Anions

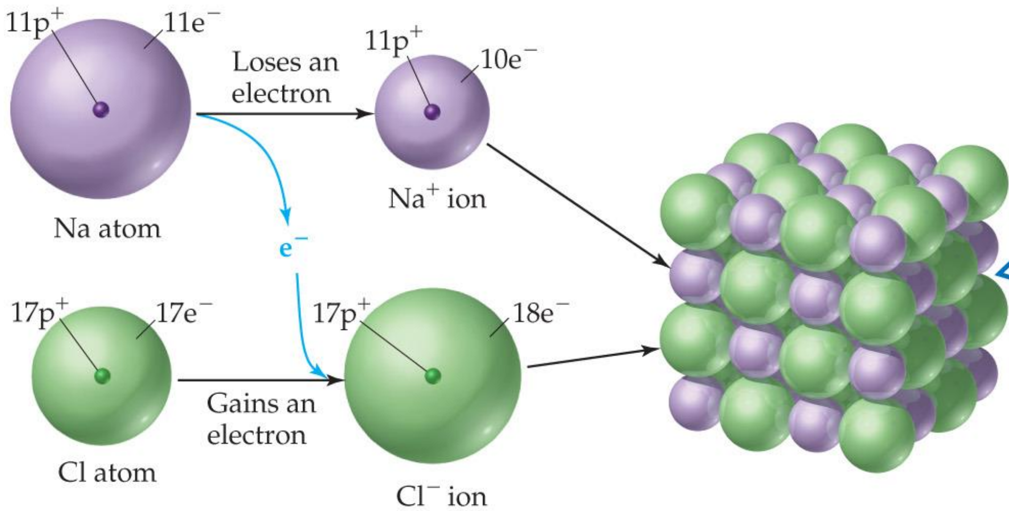
Table 2.5 Common Anions[a]

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**

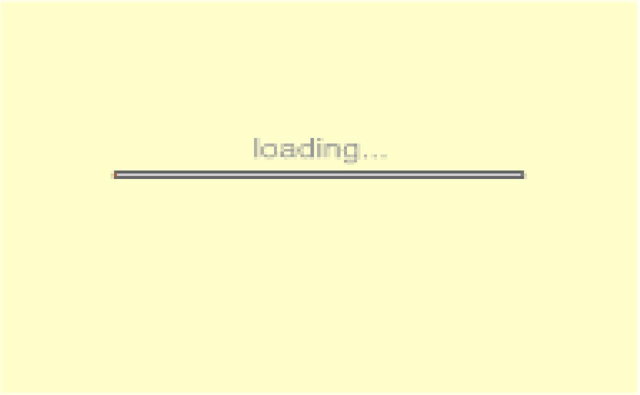
[a]The ions we use most often are in boldface. Learn them first.

Ionic Compounds

- **Ionic compounds** (such as NaCl) are generally formed between metals and nonmetals.
- Electrons are transferred from the metal to the nonmetal. The oppositely charged ions attract each other. Only empirical formulas are written.



Molecular vs Ionic Compounds



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Atoms,
Molecules,
and Ions

Writing Formulas

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

- 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.

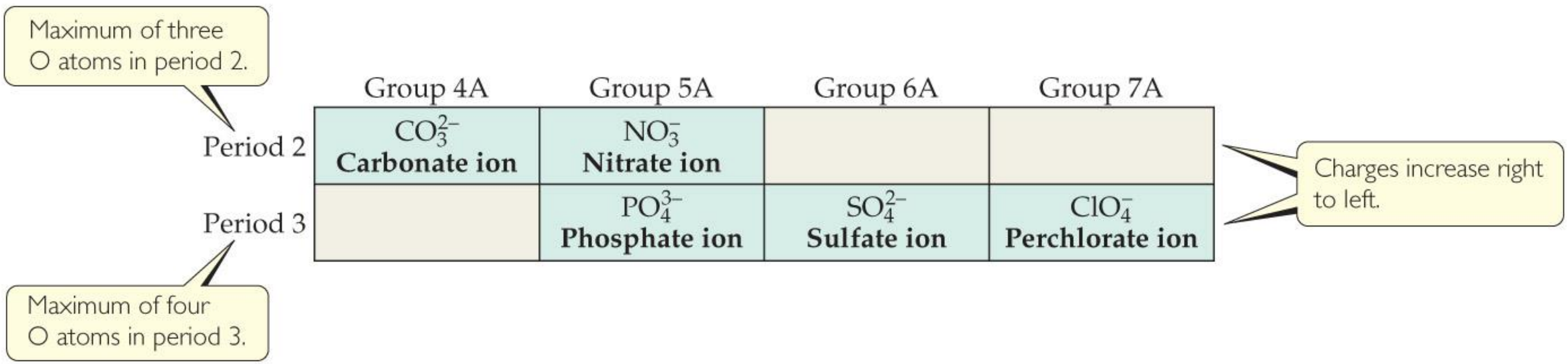
Inorganic Nomenclature

- Write the name of the cation. If the cation can have more than one possible charge, write the charge as a Roman numeral in parentheses.
- 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.

Patterns in Oxyanion Nomenclature

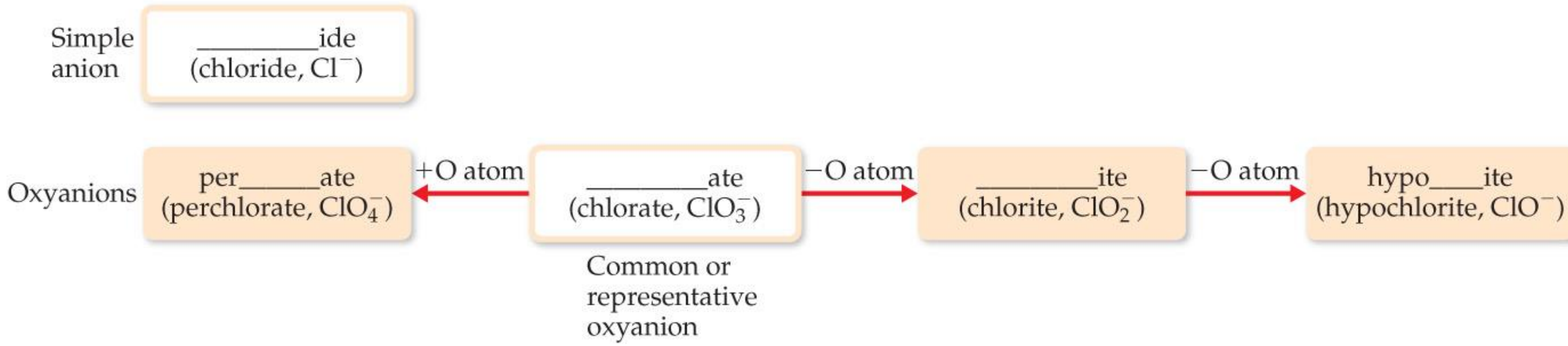
- When there are two oxyanions involving the same element
  - the one with fewer oxygens ends in ***-ite.***
  - the one with more oxygens ends in ***-ate.***
    - NO_2^-: nitrite; NO_3^-: nitrate
    - SO_3^{2-}: sulfite; SO_4^{2-}: sulfate

Patterns in Oxyanion Nomenclature



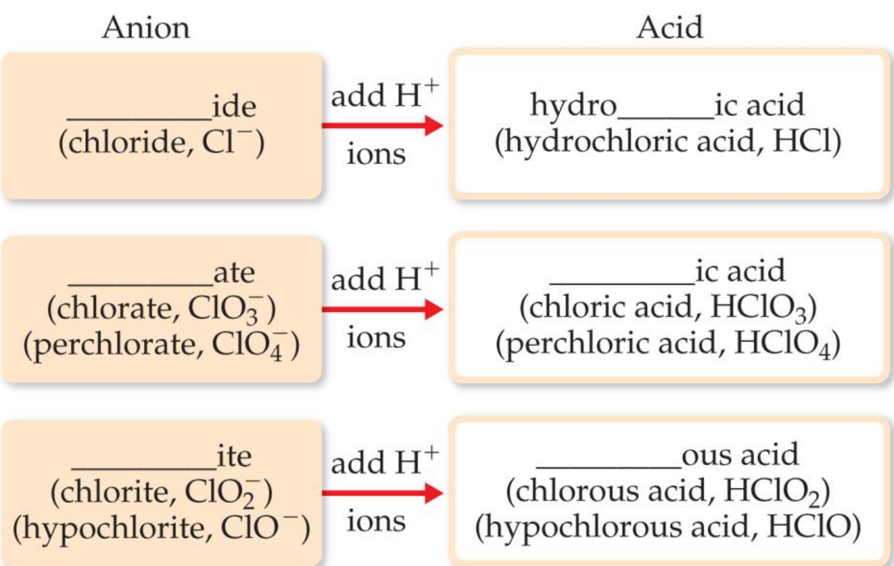
- Central atoms on the second row have a bond to, at most, three oxygens; those on the third row take up to four.
- Charges increase as you go from *right* to *left*.

Patterns in Oxyanion Nomenclature



- The one with the second fewest oxygens ends in ***-ite***: ClO_2^- is chlorite.
- The one with the second most oxygens ends in ***-ate***: ClO_3^- is chlorate.
- The one with the fewest oxygens has the prefix ***hypo-*** and ends in ***-ite***: ClO^- is hypochlorite.
- The one with the most oxygens has the prefix ***per-*** and ends in ***-ate***: ClO_4^- is 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 ends in ***-ite***, change the ending to ***-ous acid***.
  - HClO: hypochlorous acid
  - $HClO_2$: chlorous acid
- If the anion ends in ***-ate***, change the ending to ***-ic acid***.
  - $HClO_3$: chloric acid
  - $HClO_4$: perchloric acid

Nomenclature of Binary Molecular Compounds

Table 2.6 Prefixes Used in Naming Binary Compounds Formed between Nonmetals

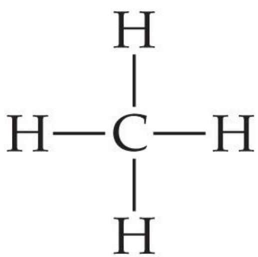
Prefix	Meaning
Mono-	*1*
Di-	*2*
Tri-	*3*
Tetra-	*4*
Penta-	*5*
Hexa-	*6*
Hepta-	*7*
Octa-	*8*
Nona-	*9*
Deca-	*10*

- The name of the element farther to the left in the periodic table (closer to the metals) or lower in the same group is usually written 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).

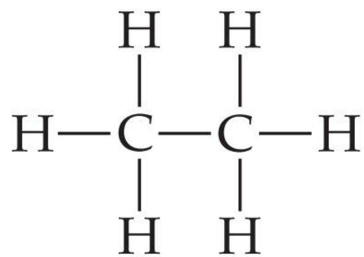
Nomenclature of Binary Compounds

- The ending on the second element is changed to ***-ide***.
  - CO_2: carbon dioxide
  - CCl_4: 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.
  - N_2O_5: dinitrogen pentoxide

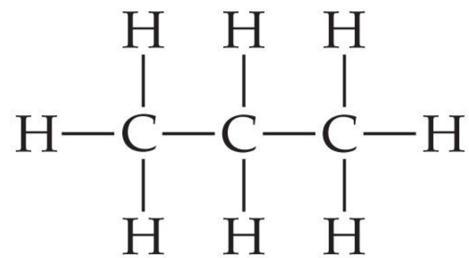
Nomenclature of Organic Compounds



Methane



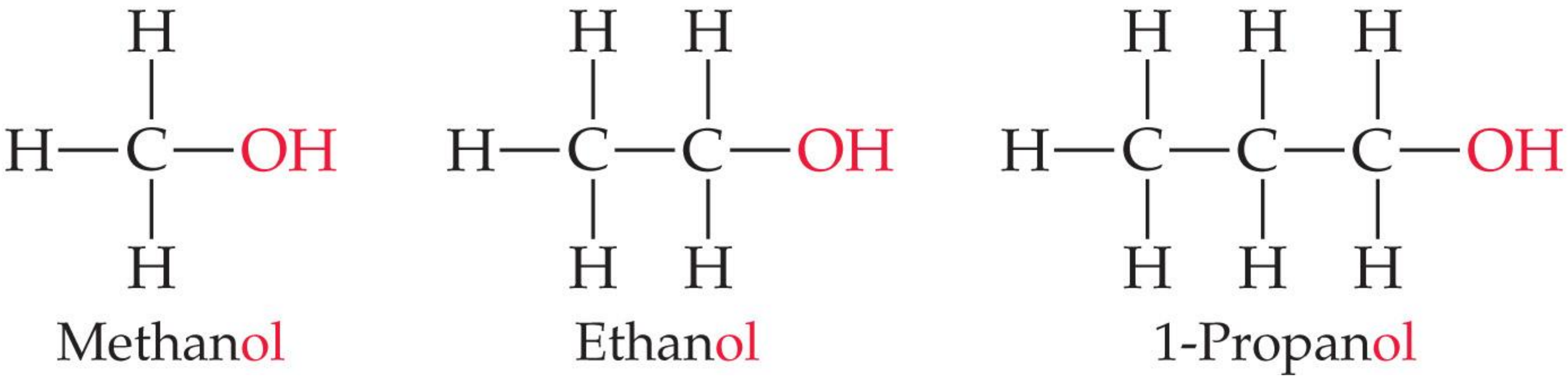
Ethane



Propane

- **Organic chemistry** is the study of carbon.
- Organic chemistry has its own system of nomenclature.
- The simplest hydrocarbons (compounds containing only carbon and hydrogen) are **alkanes**.
- The first part of the names just listed correspond to the number of carbons (***meth-*** = 1, ***eth-*** = 2, ***prop-*** = 3, etc.).

Nomenclature of Organic Compounds



- When a hydrogen in an alkane is replaced with something else (a **functional group**, like -OH in the compounds above), the name is derived from the name of the alkane.
- The ending denotes the type of compound.
  - An **alcohol** ends in *-ol*.

Sample Exercise 2.1 Atomic Size

The diameter of a U.S. dime is 17.9 mm, and the diameter of a silver atom is 2.88 Å. How many silver atoms could be arranged side by side across the diameter of a dime?

Solution

The unknown is the number of silver (Ag) atoms. Using the relationship 1 Ag atom = 2.88 Å as a conversion factor relating number of atoms and distance, we start with the diameter of the dime, first converting this distance into angstroms and then using the diameter of the Ag atom to convert distance to number of Ag atoms:

$$\text{Ag atoms} = (17.9\ \cancel{\text{mm}})\left(\frac{10^{-3}\ \cancel{\text{m}}}{1\ \cancel{\text{mm}}}\right)\left(\frac{1\ \cancel{\text{Å}}}{10^{-10}\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{Ag atom}}{2.88\ \cancel{\text{Å}}}\right) = 6.22 \times 10^{7}\ \text{Ag atoms}$$

That is, 62.2 million silver atoms could sit side by side across a dime!

Practice Exercise 1

Which of the following factors determines the size of an atom?
(a) The volume of the nucleus; **(b)** the volume of space occupied by the electrons of the atom; **(c)** the volume of a single electron, multiplied by the number of electrons in the atom; **(d)** The total nuclear charge; **(e)** The total mass of the electrons surrounding the nucleus.

Practice Exercise 2

The diameter of a carbon atom is 1.54 Å. **(a)** Express this diameter in picometers. **(b)** How many carbon atoms could be aligned side by side across the width of a pencil line that is 0.20 mm wide?

Sample Exercise 2.2 Determining the Number of Subatomic Particles in Atoms

How many protons, neutrons, and electrons are in an atom of **(a)** ^{197}Au, **(b)** strontium-90?

Solution

(a) The superscript 197 is the mass number 1protons + neutrons2. According to the list of elements given on the front inside cover, gold has atomic number 79. Consequently, an atom of ^{197}Au has 79 protons, 79 electrons, and 197 – 79 = 118 neutrons. **(b)** The atomic number of strontium is 38. Thus, all atoms of this element have 38 protons and 38 electrons. The strontium-90 isotope has 90 – 38 = 52 neutrons.

Practice Exercise 1

Which of these atoms has the largest number of neutrons in the nucleus? **(a)** ^{148}Eu, **(b)** ^{157}Dy, **(c)** ^{149}Nd, **(d)** ^{162}Ho, **(e)** ^{159}Gd.

Practice Exercise 2

How many protons, neutrons, and electrons are in an atom of **(a)** ^{138}Ba, **(b)** phosphorus-31?

Sample Exercise 2.3 Writing Symbols for Atoms

Magnesium has three isotopes with mass numbers 24, 25, and 26. **(a)** Write the complete chemical symbol (superscript and subscript) for each. **(b)** How many neutrons are in an atom of each isotope?

Solution

(a) Magnesium has atomic number 12, so all atoms of magnesium contain 12 protons and 12 electrons. The three isotopes are therefore represented by $^{24}_{12}\mathrm{Mg}$, $^{25}_{12}\mathrm{Mg}$, and $^{26}_{12}\mathrm{Mg}$. **(b)** The number of neutrons in each isotope is the mass number minus the number of protons. The numbers of neutrons in an atom of each isotope are therefore 12, 13, and 14, respectively.

Practice Exercise 1

Which of the following is an incorrect representation for a neutral atom:

$$\textbf{(a)}\ ^{6}_{3}\mathrm{Li},\ \textbf{(b)}\ ^{13}_{6}\mathrm{C},\ \textbf{(c)}\ ^{63}_{30}\mathrm{Cu},\textbf{(d)}\ ^{30}_{15}\mathrm{P},\textbf{(e)}\ ^{108}_{47}\mathrm{Ag}$$

Practice Exercise 2

Give the complete chemical symbol for the atom that contains 82 protons, 82 electrons, and 126 neutrons.

Sample Exercise 2.4 Calculating the Atomic Weight of an Element from Isotopic Abundances

Naturally occurring chlorine is 75.78% ^{35}Cl (atomic mass 34.969 amu) and 24.22% ^{37}Cl (atomic mass 36.966 amu). Calculate the atomic weight of chlorine.

Solution

We can calculate the atomic weight by multiplying the abundance of each isotope by its atomic mass and summing these products. Because 75.78% = 0.7578 and 24.22% = 0.2422, we have

$$\begin{aligned}\text{Atomic weight} &= (0.7578)\,(34.969\ \text{amu}) + (0.2422)(36.966\ \text{amu})\\ &= 26.50\ \text{amu} + 8.953\ \text{amu}\\ &= 35.45\ \text{amu}\end{aligned}$$

This answer makes sense: The atomic weight, which is actually the average atomic mass, is between the masses of the two isotopes and is closer to the value of ^{35}Cl, the more abundant isotope.

Practice Exercise 1

The atomic weight of copper, Cu, is listed as 63.546. Which of the following statements are untrue?

(a) Not all the atoms of copper have the same number of electrons.
(b) All the copper atoms have 29 protons in the nucleus.
(c) The dominant isotopes of Cu must be ^{63}Cu and ^{64}Cu.
(d) Copper is a mixture of at least two isotopes.
(e) The number of electrons in the copper atoms is independent of atomic mass.

Sample Exercise 2.4 Calculating the Atomic Weight of an Element from Isotopic Abundances

Continued

Practice Exercise 2

Three isotopes of silicon occur in nature: ^{28}Si (92.23%), atomic mass 27.97693 amu; ^{29}Si (4.68%), atomic mass 28.97649 amu; and ^{30}Si (3.09%), atomic mass 29.97377 amu. Calculate the atomic weight of silicon.

Sample Exercise 2.5 Using the Periodic Table

Which two of these elements would you expect to show the greatest similarity in chemical and physical properties: B, Ca, F, He, Mg, P?

Solution

Elements in the same group of the periodic table are most likely to exhibit similar properties. We therefore expect Ca and Mg to be most alike because they are in the same group (2A, the alkaline earth metals).

Practice Exercise 1

A biochemist who is studying the properties of certain sulfur (S)–containing compounds in the body wonders whether trace amounts of another nonmetallic element might have similar behavior. To which element should she turn her attention? **(a)** O, **(b)** As, **(c)** Se, **(d)** Cr, **(e)** P.

Practice Exercise 2

Locate Na (sodium) and Br (bromine) in the periodic table. Give the atomic number of each and classify each as metal, metalloid, or nonmetal.

Sample Exercise 2.6 Relating Empirical and Molecular Formulas

Write the empirical formulas for **(a)** glucose, a substance also known as either blood sugar or dextrose—molecular formula $C_6H_{12}O_6$; **(b)** nitrous oxide, a substance used as an anesthetic and commonly called laughing gas—molecular formula N_2O.

Solution

(a) The subscripts of an empirical formula are the smallest whole-number ratios. The smallest ratios are obtained by dividing each subscript by the largest common factor, in this case 6. The resultant empirical formula for glucose is CH_2O.

(b) Because the subscripts in N_2O are already the lowest integral numbers, the empirical formula for nitrous oxide is the same as its molecular formula, N_2O.

Practice Exercise 1

Tetracarbon dioxide is an unstable oxide of carbon with the following molecular structure:

What are the molecular and empirical formulas of this substance? **(a)** C_2O_2, CO_2, **(b)** C_4O, CO, **(c)** CO_2, CO_2, **(d)** C_4O_2, C_2O, **(e)** C_2O, CO_2.

Practice Exercise 2

Give the empirical formula for *decaborane*, whose molecular formula is $B_{10}H_{14}$.

Sample Exercise 2.7 Writing Chemical Symbols for Ions

Give the chemical symbol, including superscript indicating mass number, for **(a)** the ion with 22 protons, 26 neutrons, and 19 electrons; and **(b)** the ion of sulfur that has 16 neutrons and 18 electrons.

Solution

(a) The number of protons is the atomic number of the element. A periodic table or list of elements tells us that the element with atomic number 22 is titanium (Ti). The mass number (protons plus neutrons) of this isotope of titanium is 22 + 26 = 48. Because the ion has three more protons than electrons, it has a net charge of 3+ and is designated $^{48}Ti^{3+}$.

(b) The periodic table tells us that sulfur (S) has an atomic number of 16. Thus, each atom or ion of sulfur contains 16 protons. We are told that the ion also has 16 neutrons, meaning the mass number is 16 + 16 = 32. Because the ion has 16 protons and 18 electrons, its net charge is 2− and the ion symbol is $^{32}S^{2-}$.

In general, we will focus on the net charges of ions and ignore their mass numbers unless the circumstances dictate that we specify a certain isotope.

Practice Exercise 1

In which of the following species is the number of protons less than the number of electrons? **(a)** Ti^{2+}, **(b)** P^{3-}, **(c)** Mn, **(d)** Se_4^{-2}, **(e)** Ce^{4+}.

Practice Exercise 2

How many protons, neutrons, and electrons does the $^{79}Se^{2-}$ ion possess?

Sample Exercise 2.8 Predicting Ionic Charge

Predict the charge expected for the most stable ion of barium and the most stable ion of oxygen.

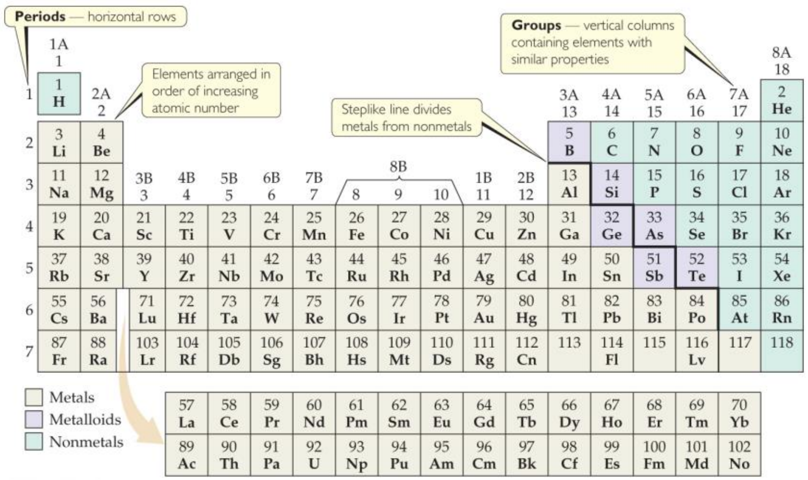
Solution

We will assume that barium and oxygen form ions that have the same number of electrons as the nearest noble-gas atom. From the periodic table, we see that barium has atomic number 56. The nearest noble gas is xenon, atomic number 54. Barium can attain a stable arrangement of 54 electrons by losing two electrons, forming the Ba^{2+} cation.

Oxygen has atomic number 8. The nearest noble gas is neon, atomic number 10. Oxygen can attain this stable electron arrangement by gaining two electrons, forming the O^{2-} anion.

Practice Exercise 1

Although it is helpful to know that many ions have the electron arrangement of a noble gas, many elements, especially among the metals, form ions that do not have a noble–gas electron arrangement. Use the periodic table, Figure 2.14, to determine which of the following ions has a noble–gas electron arrangement, and which do not. For those that do, indicate the noble–gas arrangement they match: **(a)** Ti^{4+}, **(b)** Mn^{2+}, **(c)** Pb^{2+}, **(d)** Te^{2-}, **(e)** Zn^{2+}.



Practice Exercise 2

Predict the charge expected for the most stable ion of **(a)** aluminum and **(b)** fluorine.

Sample Exercise 2.9 Identifying Ionic and Molecular Compounds

Which of these compounds would you expect to be ionic: N_2O, Na_2O, CaC_{l2}, SF_4?

Solution

We predict that Na_2O and CaC_{l2} are ionic compounds because they are composed of a metal combined with a nonmetal. We predict (correctly) that N_2O and SF_4 are molecular compounds because they are composed entirely of nonmetals.

Practice Exercise 1

Which of these compounds are molecular: CBr_4, FeS, P_4O_6, PbF_2?

Practice Exercise 2

Give a reason why each of the following statements is a safe prediction:

(a) Every compound of Rb with a nonmetal is ionic in character.
(b) Every compound of nitrogen with a halogen element is a molecular compound.
(c) The compound $MgKr_2$ does not exist.
(d) Na and K are very similar in the compounds they form with nonmetals.
(e) If contained in an ionic compound, calcium (Ca) will be in the form of the doubly charged ion, Ca^{2+}.

Sample Exercise 2.10 Using Ionic Charge to Write Empirical Formulas for Ionic Compounds

Write the empirical formula of the compound formed by **(a)** Al^{3+} and Cl^{-} ions, **(b)** Al^{3+} and O^{2-} ions, **(c)** Mg^{2+} and NO^{3-} ions.

Solution

(a) Three Cl^{-} ions are required to balance the charge of one Al^{3+} ion, making the empirical formula $AlCl_3$.
(b) Two Al^{3+} ions are required to balance the charge of three O^{2-} ions. A 2:3 ratio is needed to balance the total positive charge of 6+ and the total negative charge of 6–. The empirical formula is Al_2O_3.
(c) Two NO_3^{-} ions are needed to balance the charge of one Mg^{2+}, yielding $Mg(NO_3)_2$. Note that the formula for the polyatomic ion, NO_3^{-}, must be enclosed in parentheses so that it is clear that the subscript 2 applies to all the atoms of that ion.

Practice Exercise 1

For the following ionic compounds formed with S2–, what is the empirical formula for the positive ion involved? **(a)** MnS, **(b)** Fe_2S_3, **(c)** MoS_2, **(d)** K_2S, **(e)** Ag_2S.

Practice Exercise 2

Write the empirical formula for the compound formed by **(a)** Na^{+} and PO_4^{3-}, **(b)** Zn^{2+} and SO_4^{2-}, **(c)** Fe^{3+} and CO_3^{2-}.

Sample Exercise 2.11 Determining the Formula of an Oxyanion from Its Name

Based on the formula for the sulfate ion, predict the formula for **(a)** the selenate ion and **(b)** the selenite ion. (Sulfur and selenium are both in group 6A and form analogous oxyanions.)

Solution

(a) The sulfate ion is SO_4^{2-}. The analogous selenate ion is therefore SeO_4^{2-}.

(b) The ending *-ite* indicates an oxyanion with the same charge but one O atom fewer than the corresponding oxyanion that ends in *-ate*. Thus, the formula for the selenite ion is SeO_3^{2-}.

Practice Exercise 1

Which of the following oxyanions is incorrectly named? **(a)** ClO_2^-, chlorate; **(b)** IO_4^-, periodate; **(c)** SO_3^{2-}, sulfite; **(d)** IO_3^-, iodate; **(e)** SeO_4^{2-}, selenate.

Practice Exercise 2

The formula for the bromate ion is analogous to that for the chlorate ion. Write the formula for the hypobromite and bromite ions.

Sample Exercise 2.12 Determining the Names of Ionic Compounds from Their Formulas

Name the ionic compounds **(a)** K_2SO_4, **(b)** $Ba(OH)_2$, **(c)** $FeCl_3$.

Solution

In naming ionic compounds, it is important to recognize polyatomic ions and to determine the charge of cations with variabl charge.

(a) The cation is K^+, the potassium ion, and the anion is SO_4^{2-}, the sulfate ion, making the name potassium sulfate. (If you thought the compound contained S^{2-} and O^{2-} ions, you failed to recognize the polyatomic sulfate ion.)

(b) The cation is Ba^{2+}, the barium ion, and the anion is OH^-, the hydroxide ion: barium hydroxide.

(c) You must determine the charge of Fe in this compound because an iron atom can form more than one cation. Because the compound contains three chloride ions, Cl^-, the cation must be Fe^{3+}, the iron(III), or ferric, ion. Thus, the compound is iron(III) chloride or ferric chloride.

Practice Exercise 1

Which of the following ionic compounds is incorrectly named? **(a)** $Zn(NO_3)_2$, zinc nitrate; **(b)** $TeCl_4$, tellurium(IV) chloride; **(c)** Fe_2O_3, diiron oxide; **(d)** BaO, barium oxide; **(e)** $Mn_3(PO_4)_2$, manganese (II) phosphate.

Practice Exercise 2

Name the ionic compounds **(a)** NH_4Br, **(b)** Cr_2O_3, **(c)** $Co(NO_3)_2$.

Sample Exercise 2.13 Relating the Names and Formulas of Acids

Name the acids **(a)** HCN, **(b)** HNO_3, **(c)** H_2SO^4, **(d)** H_2SO_3.

Solution

(a) The anion from which this acid is derived is CN^-, the cyanide ion. Because this ion has an -*ide* ending, the acid is given a *hydro*- prefix and an -*ic* ending: hydrocyanic acid. Only water solutions of HCN are referred to as hydrocyanic acid. The pure compound, which is a gas under normal conditions, is called hydrogen cyanide. Both hydrocyanic acid and hydrogen cyanide are *extremely* toxic.

(b) Because NO_3^- is the nitrate ion, HNO_3 is called nitric acid (the -*ate* ending of the anion is replaced with an -*ic* ending in naming the acid).

(c) Because SO_4^{2-} is the sulfate ion, H_2SO_4 is called sulfuric acid.

(d) Because SO_3^{2-} is the sulfite ion, H_2SO_3 is sulfurous acid (the -*ite* ending of the anion is replaced with an -*ous* ending).

Practice Exercise 1

Which of the following acids are incorrectly named? For those that are, provide a correct name or formula.
(a) hydrocyanic acid, HCN; **(b)** nitrous acid, HNO_3; **(c)** perbromic acid, $HBrO_4$; **(d)** iodic acid, HI;
(e) selenic acid, $HSeO_4$.

Practice Exercise 2

Give the chemical formulas for **(a)** hydrobromic acid, **(b)** carbonic acid.

Sample Exercise 2.14 Relating the Names and Formulas of Binary Molecular Compounds

Name the compounds **(a)** SO_2, **(b)** PCl_5, **(c)** Cl_2O_3.

Solution

The compounds consist entirely of nonmetals, so they are molecular rather than ionic. Using the prefixes in Table 2.6, we have **(a)** sulfur dioxide, **(b)** phosphorus pentachloride, **(c)** dichlorine trioxide.

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*

Practice Exercise 1

Give the name for each of the following binary compounds of carbon:
(a) CS_2, **(b)** CO, **(c)** C_3O_2, **(d)** CBr_4, **(e)** CF.

Practice Exercise 2

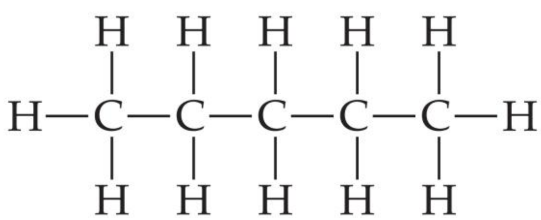
Give the chemical formulas for **(a)** silicon tetrabromide, **(b)** disulfur dichloride, **(c)** diphosphorus hexaoxide.

Sample Exercise 2.15 Writing Structural and Molecular Formulas for Hydrocarbons

Assuming the carbon atoms in pentane are in a linear chain, write **(a)** the structural formula and **(b)** the molecular formula for this alkane.

Solution

(a) Alkanes contain only carbon and hydrogen, and each carbon is attached to four other atoms. The name pentane contains the prefix *penta-* for five (Table 2.6), and we are told that the carbons are in a linear chain. If we then add enough hydrogen atoms to make four bonds to each carbon, we obtain the structural formula



This form of pentane is often called *n*-pentane, where the *n*- stands for "normal" because all five carbon atoms are in one line in the structural formula.

(b) Once the structural formula is written, we determine the molecular formula by counting the atoms present. Thus, *n*-pentane has the molecular formula C_5H_{12}.

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

Sample Exercise 2.15 Writing Structural and Molecular Formulas for Hydrocarbons

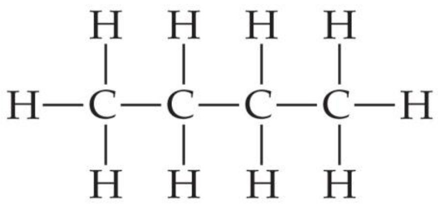
Continued

Practice Exercise 1

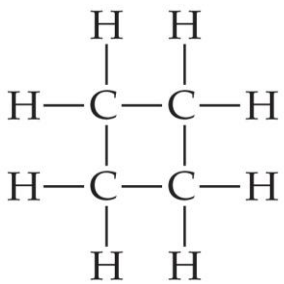
(a) What is the molecular formula of hexane, the alkane with six carbons? **(b)** What are the name and molecular formula of an alcohol derived from hexane?

Practice Exercise 2

These two compounds have "butane" in their name. Are they isomers?



butane



cyclobutane