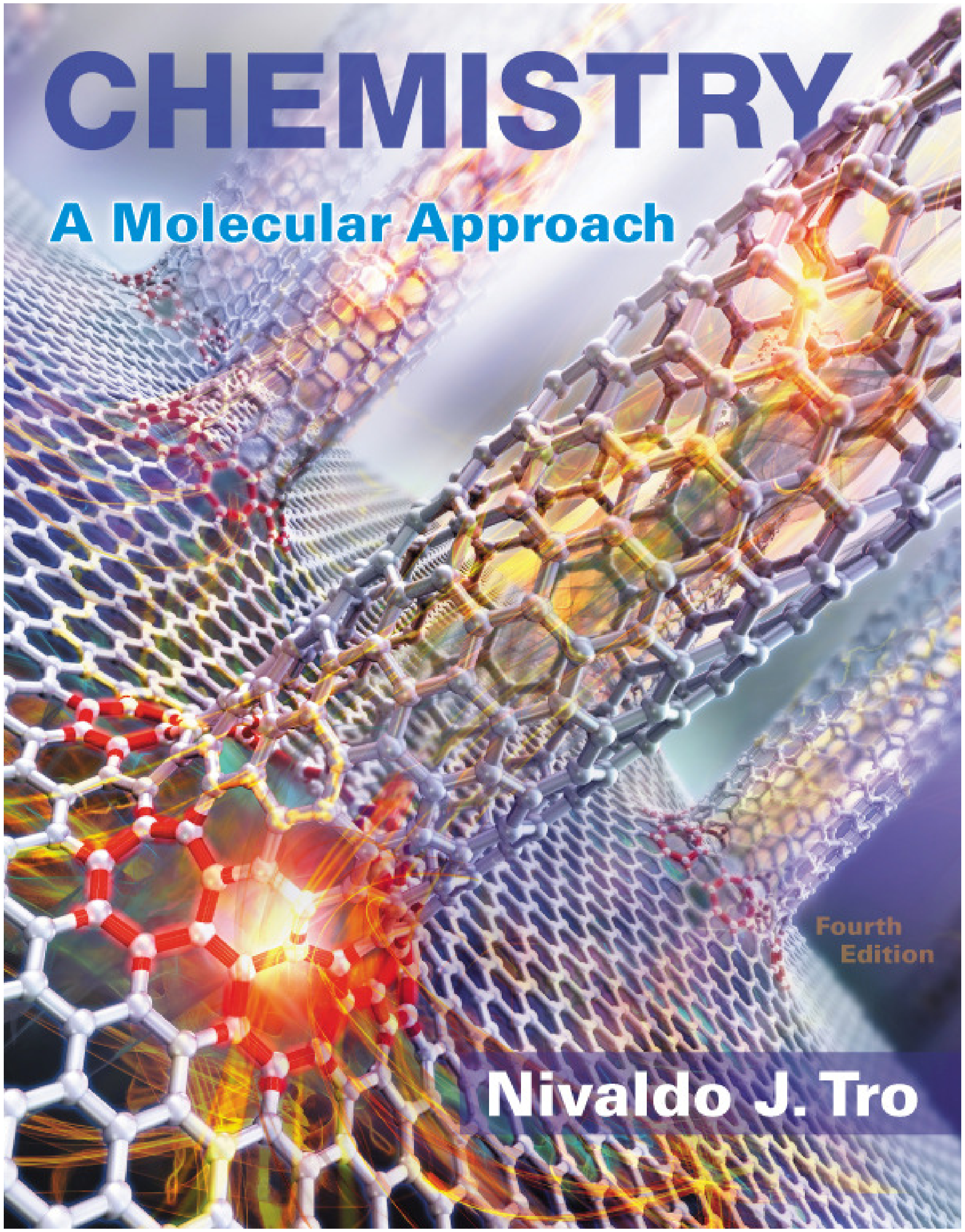


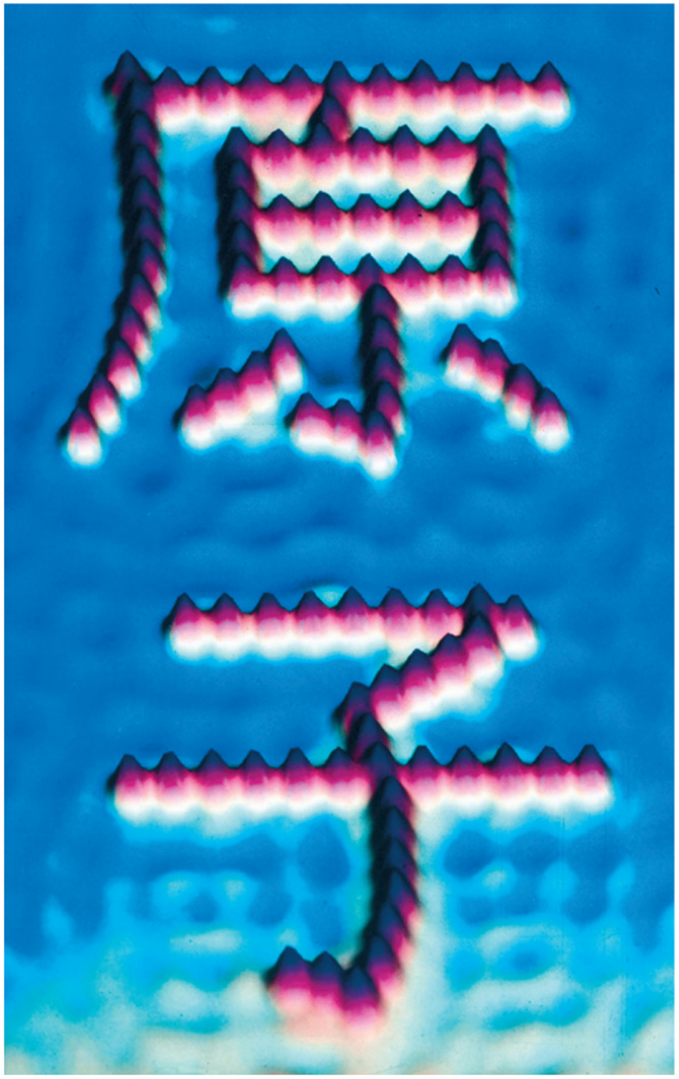
Lecture Presentation

Lecture 02.1

Atoms and Elements

Imaging and Moving Individual Atoms

- **Scanning tunneling microscopy (STM)** is a technique that can image, and even move, individual atoms and molecules.
- The image below, obtained by STM, shows iron atoms (red) on a copper surface (blue).



Early Ideas about the Building Blocks of Matter

- Leucippus (fifth century B.C.) and his student Democritus (460–370 B.C.) were the first to propose that matter was composed of small, indestructible particles.
  - Democritus wrote, “Nothing exists except atoms and empty space; everything else is opinion.”
- They proposed that many different kinds of atoms existed, each different in shape and size, and that they moved randomly through empty space.

Early Building Blocks of Matter Ideas

- Plato and Aristotle did not embrace the atomic ideas of Leucippus and Democritus.
- They held that
  - matter had no smallest parts.
  - different substances were composed of various proportions of fire, air, earth, and water.

Early Building Blocks of Matter Ideas

- Later, the scientific approach became the established way to learn about the physical world.

- An English chemist, John Dalton (1766–1844), offered convincing evidence that supported the early atomic ideas of Leucippus and Democritus.

Modern Atomic Theory and the Laws That Led to It

- The theory that all matter is composed of atoms grew out of observations and laws.
- The three most important laws that led to the development and acceptance of the atomic theory were
  - the law of conservation of mass,
  - the law of definite proportions, and
  - the law of multiple proportions.

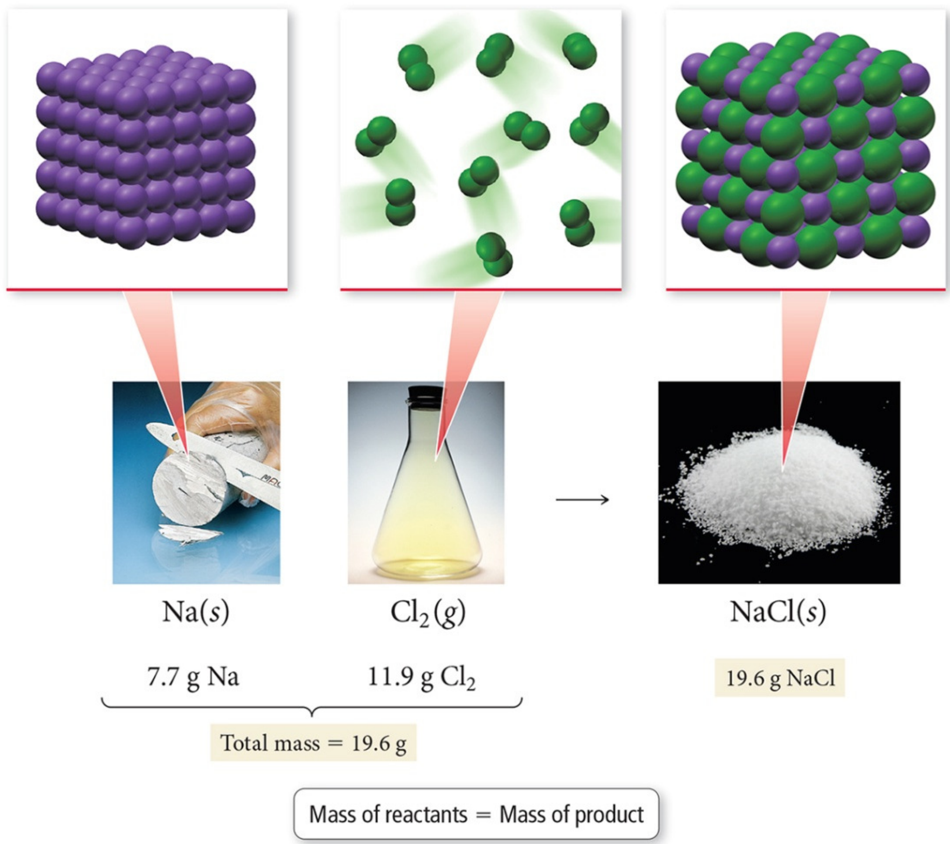
The Law of Conservation of Mass

- Antoine Lavoisier formulated the **law of conservation of mass**, which states that ***in a chemical reaction, matter is neither created nor destroyed.***

- Hence, when a chemical reaction occurs, the total mass of the substances involved in the reaction does not change.

The Law of Conservation of Mass

- This law is consistent with the idea that matter is composed of small, indestructible particles.



The Law of Definite Proportions

- In 1797, a French chemist, Joseph Proust, made observations on the composition of compounds.
- He summarized his observations in the **law of definite proportions**.
  - ***All samples of a given compound, regardless of their source or how they were prepared, have the same proportions of their constituent elements.***

The Law of Definite Proportions

- The *law of definite proportions* is sometimes called the *law of constant composition*.
  - For example, the decomposition of 18.0 g of water results in 16.0 g of oxygen and 2.0 g of hydrogen, or an oxygen-to-hydrogen mass ratio of the following:

$$\text{Mass ratio} = \frac{16.0\ \text{g O}}{2.0\ \text{g H}} = 8.0 \text{ or } 8{:}1$$

The Law of Multiple Proportions

- In 1804, John Dalton published his **law of multiple proportions**.
  - ***When two elements (call them A and B) form two different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers.***
- An atom of A combines with either one, two, three, or more atoms of B (AB_1, AB_2, AB_3, etc.).

The Law of Multiple Proportions

- Carbon monoxide and carbon dioxide are two compounds composed of the same two elements: carbon and oxygen.
  - The mass ratio of oxygen to carbon in carbon dioxide is 2.67:1; therefore, 2.67 g of oxygen reacts with 1 g of carbon.
  - In carbon monoxide, however, the mass ratio of oxygen to carbon is 1.33:1, or 1.33 g of oxygen to every 1 g of carbon.

The Law of Multiple Proportions

Carbon dioxide

Mass oxygen that combines with 1 g carbon = 2.67 g

Carbon monoxide

Mass oxygen that combines with 1 g carbon = 1.33 g

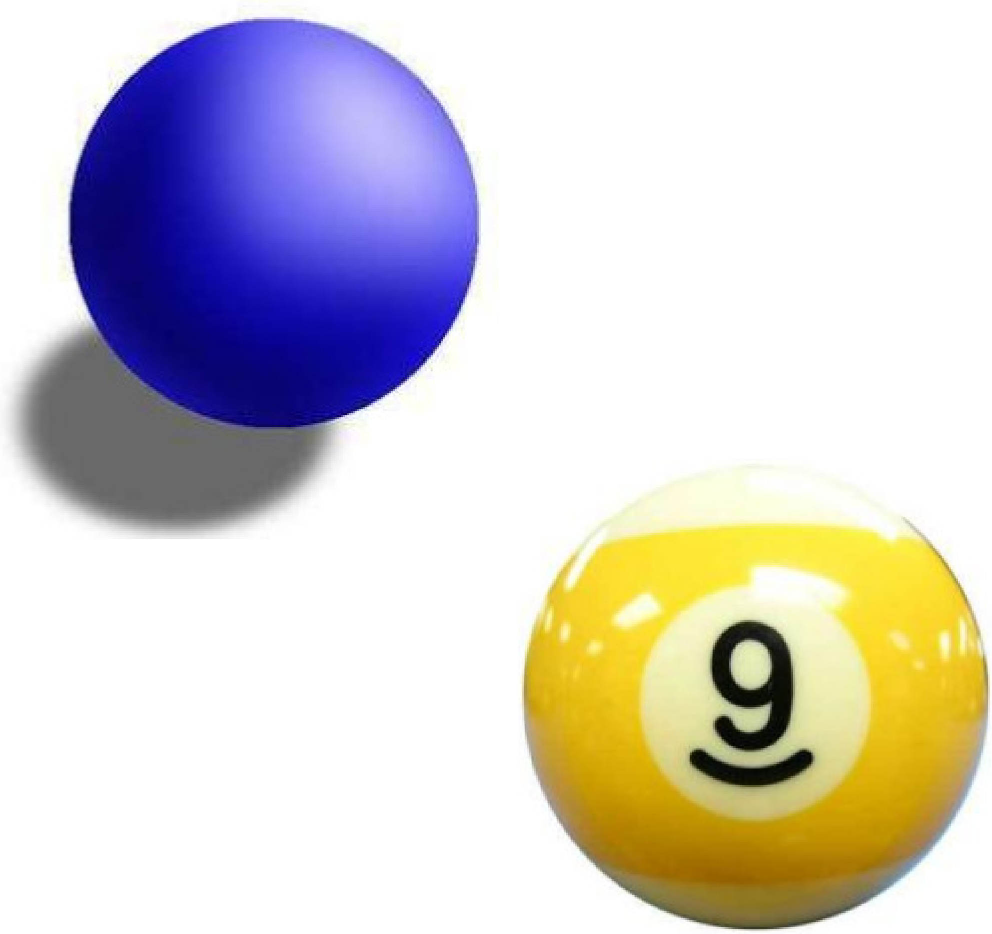
- The ratio of these two masses is itself a small whole number.

$$\frac{\text{Mass oxygen to 1 g carbon in carbon dioxide}}{\text{Mass oxygen to 1 g carbon in carbon monoxide}} = \frac{2.67}{1.33} = 2$$

John Dalton and the Atomic Theory

- Dalton's **atomic theory** explained the laws as follows**:**
  1. Each element is composed of tiny, indestructible particles called atoms.
  2. All atoms of a given element have the same mass and other properties that distinguish them from the atoms of other elements.
  3. Atoms combine in simple, whole-number ratios to form compounds.
  4. Atoms of one element cannot change into atoms of another element. In a chemical reaction, atoms only change the way that they are bound together with other atoms.

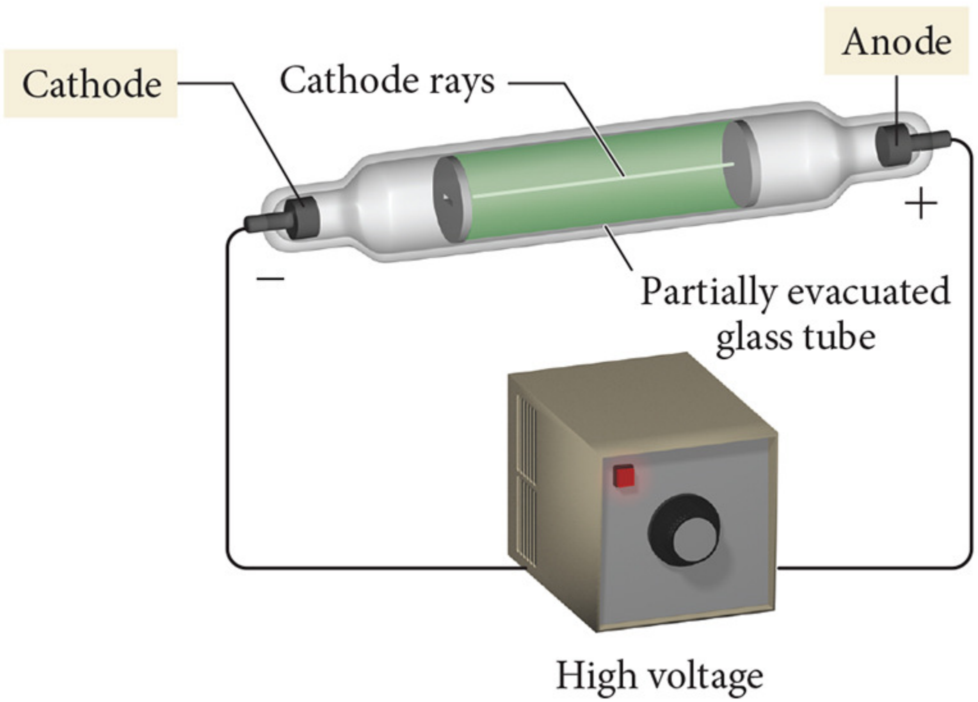
John Dalton and the Atomic Theory

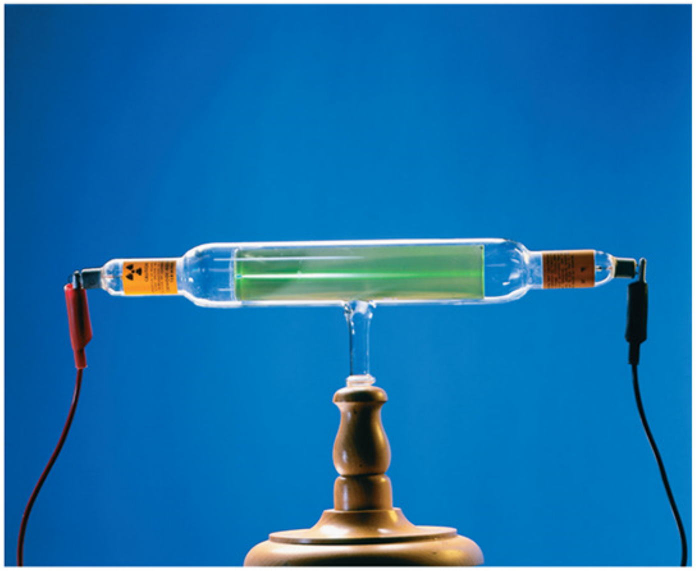


The Discovery of the Electron

- J. J. Thomson (1856–1940) conducted **cathode ray** experiments.
- Thomson constructed a partially evacuated glass tube called a **cathode ray tube**.
- He found that a beam of particles, called cathode rays, traveled from the negatively charged electrode (called the cathode) to the positively charged one (called the anode).

The Discovery of the Electron





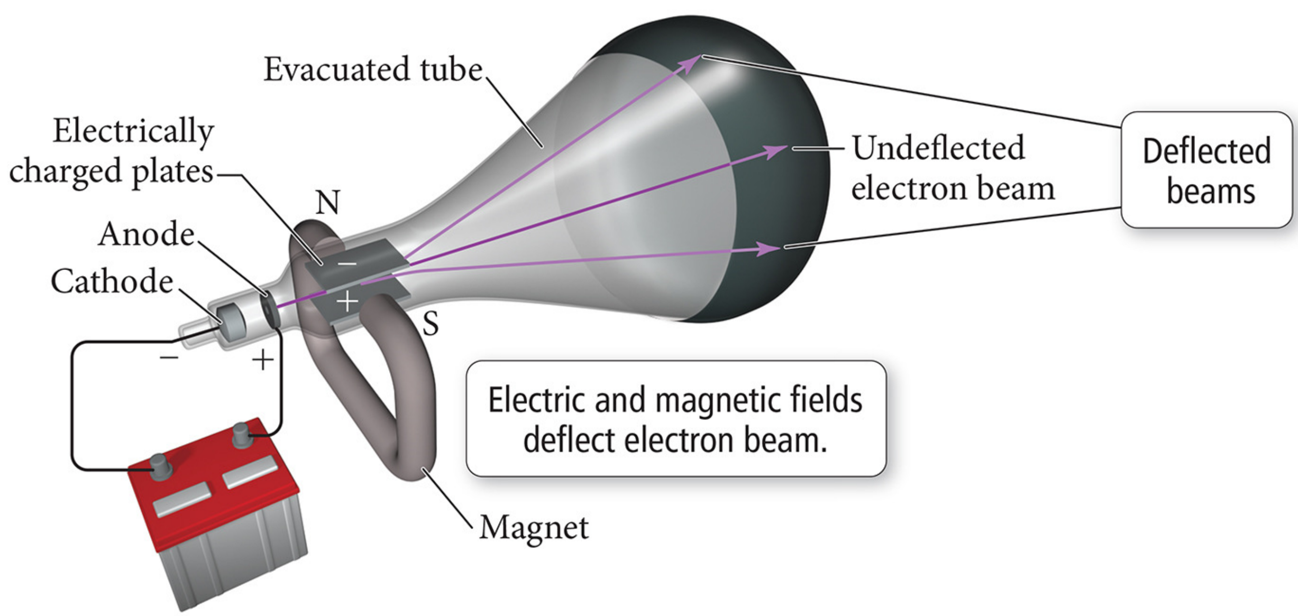
The Discovery of the Electron

- Thomson found that the particles that compose the cathode ray have the following properties:
  - They travel in straight lines.
  - They are independent of the composition of the material from which they originate (the cathode).
  - They carry a negative **electrical charge**.

The Discovery of the Electron

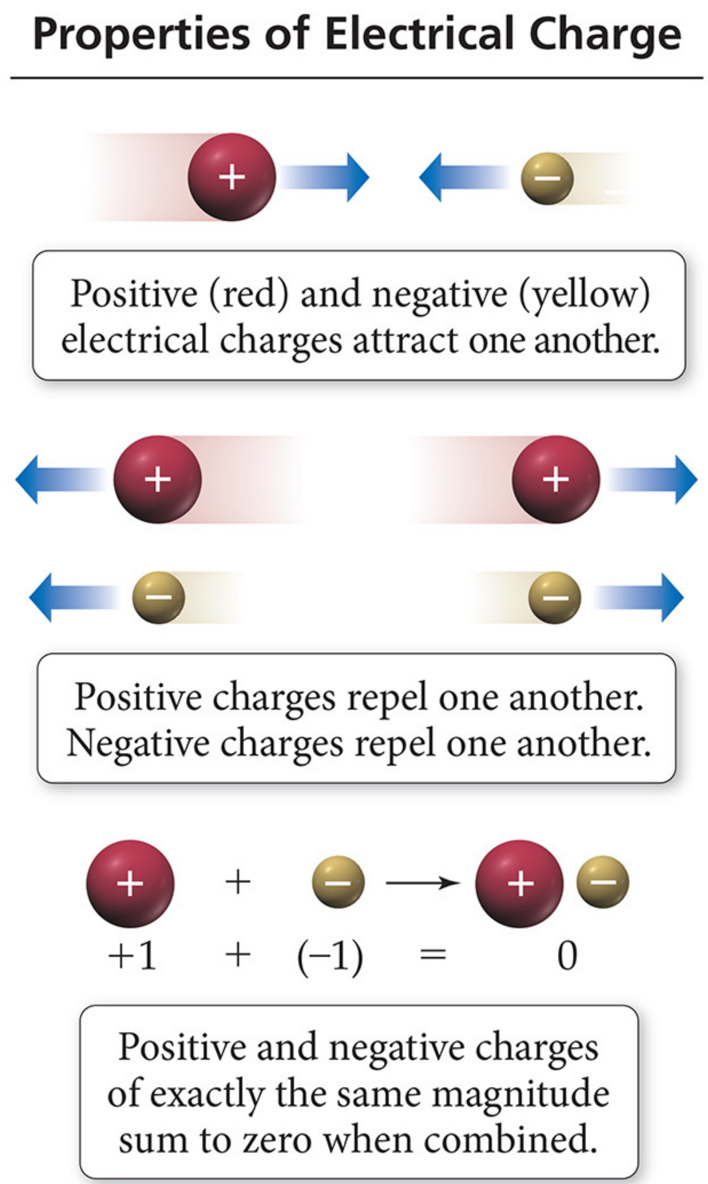
- J. J. Thomson measured the charge-to-mass ratio of the cathode ray particles by deflecting them using electric and magnetic fields, as shown in the figure.
- The value he measured was -1.76×10^3 coulombs (C) per gram.

Charge-to-Mass Ratio of the Electron



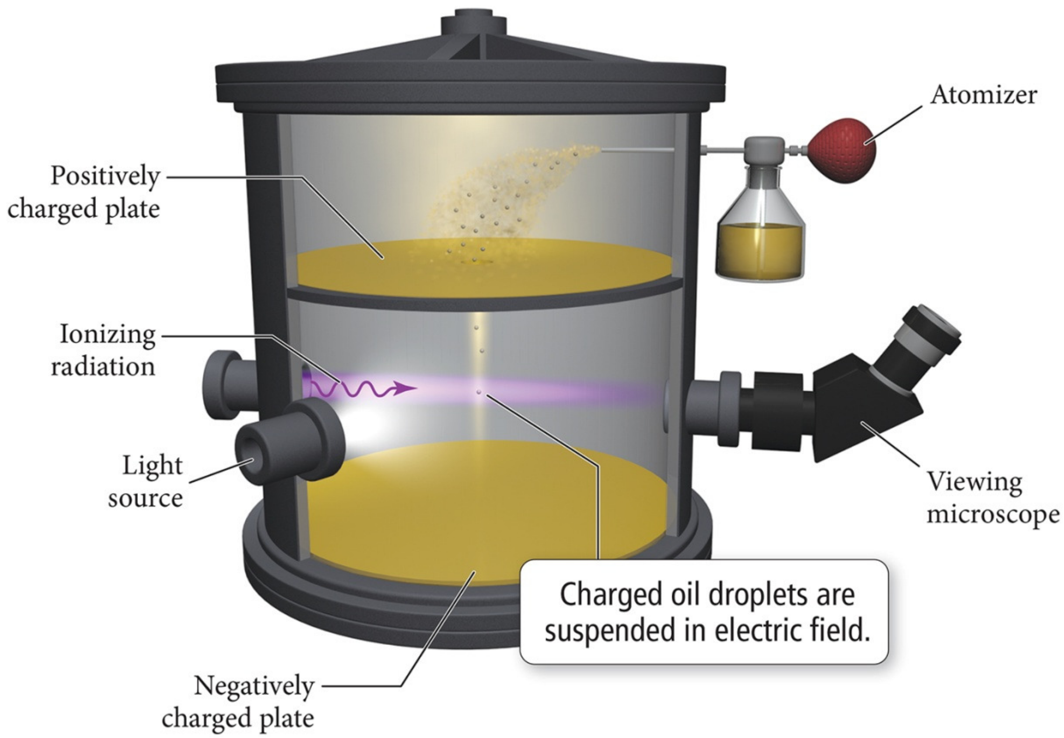
The Discovery of the Electron

- J. J. Thomson had discovered the **electron**, a negatively charged, low-mass particle present within all atoms.



Millikan's Oil Drop Experiment: The Charge of the Electron

- American physicist Robert Millikan (1868–1953) performed his now famous oil drop experiment in which he deduced the charge of a single electron.



Millikan’s Oil Drop Experiment

- By measuring the strength of the electric field required to halt the free fall of the drops, and by figuring out the masses of the drops themselves (determined from their radii and density), Millikan calculated the charge of each drop.
- The measured charge on any drop was always a whole-number multiple of -1.96×10^{-19}, the fundamental charge of a single electron.

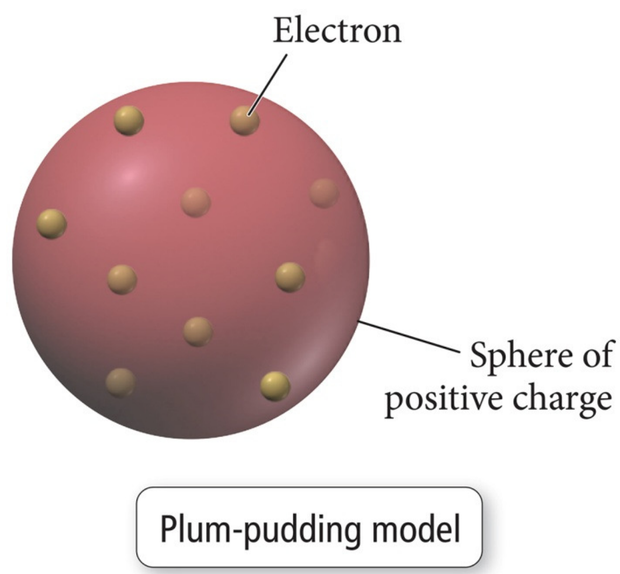
Millikan's Oil Drop Experiment

- With this number in hand, and knowing Thomson's mass-to-charge ratio for electrons, we can deduce the mass of an electron:

$$\cancel{\text{Charge}} \times \frac{\text{mass}}{\cancel{\text{charge}}} = \text{mass}$$

The Structure of the Atom

- J. J. Thomson proposed that the negatively charged electrons were small particles held within a positively charged sphere.

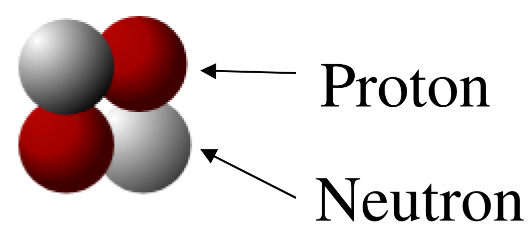


Analogy

- This model, the most popular of the time, became known as the plum-pudding model.

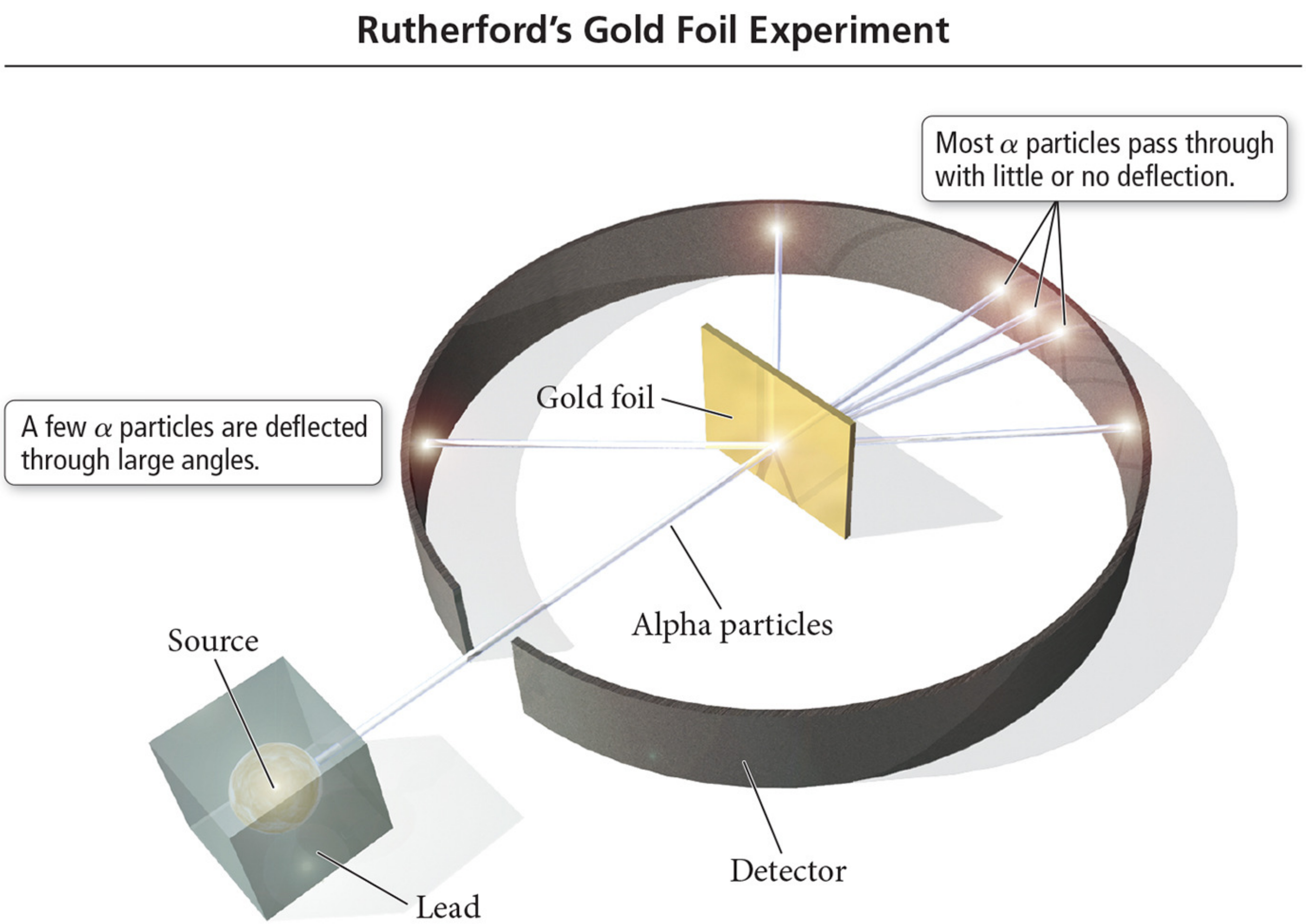
Rutherford's Gold Foil Experiment

- In 1909, Ernest Rutherford (1871–1937), who had worked under Thomson and subscribed to his plum-pudding model, performed an experiment in an attempt to confirm Thomson's model.
- In the experiment, Rutherford directed the positively charged alpha particles at an ultra thin sheet of gold foil.



$$\alpha \text{ (alpha particle)} = \mathrm{He}_2^4$$

Rutherford’s Gold Foil Experiment



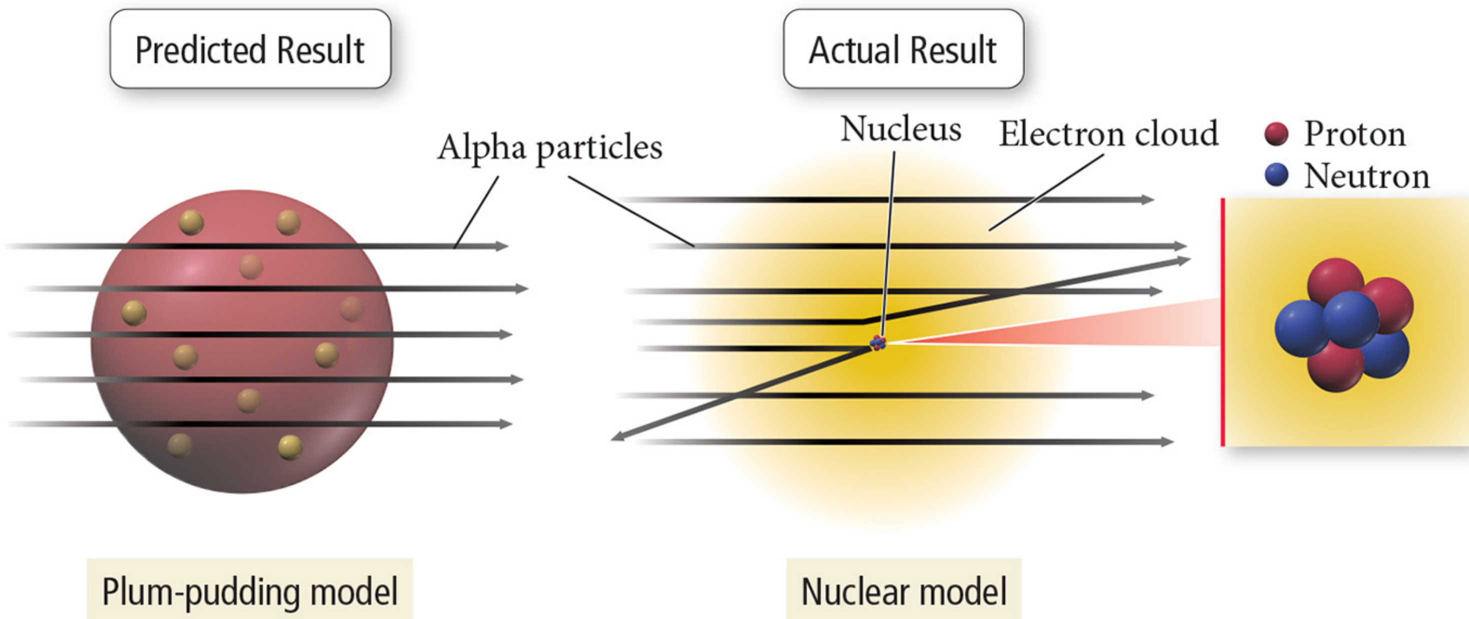
Rutherford’s Gold Foil Experiment

- The Rutherford experiment gave an unexpected result. The majority of particles did pass directly through the foil, but some particles were deflected, and some (approximately 1 in 20,000) even bounced back.

- Rutherford created a new model—a modern version of which is shown in Figure 2.6 alongside the plum-pudding model—to explain his results.

Rutherford's Gold Foil Experiment

- He concluded that matter must not be as uniform as it appears. It must contain large regions of empty space dotted with small regions of very dense matter.



Rutherford's Gold Foil Experiment

- Building on this idea, he proposed the **nuclear theory** of the atom, with three basic parts:
  1. Most of the atom's mass and all of its positive charge are contained in a small core called a **nucleus**.
  2. Most of the volume of the atom is empty space, throughout which tiny, negatively charged electrons are dispersed.
  3. There are as many negatively charged electrons outside the nucleus as there are positively charged particles (named **protons**) within the nucleus, so that the atom is electrically neutral.

The Neutrons

- Although Rutherford’s model was highly successful, scientists realized that it was incomplete.

- Later work by Rutherford and one of his students, British scientist James Chadwick (1891–1974), demonstrated that the previously unaccounted for mass was due to **neutrons**, neutral particles within the nucleus.

The Neutrons

- The mass of a neutron is similar to that of a proton.
- However, a neutron has no electrical charge.
  - The helium atom is four times as massive as the hydrogen atom because it contains two protons and *two neutrons.*
- Hydrogen, on the other hand, contains only one proton and no neutrons.

Subatomic Particles

- All atoms are composed of the same subatomic particles:
  - Protons
  - Neutrons
  - Electrons
- Protons and neutrons, as we saw earlier, have nearly identical masses.
  - The mass of the proton is 1.67262×10^{-27} kg.
  - The mass of the neutron is 1.67493×10^{-27} kg.
  - The mass of the electron is 0.00091×10^{-27} kg.

Subatomic Particles

- *The charge of the proton and the electron are equal in magnitude but opposite in sign*. The neutron has no charge.

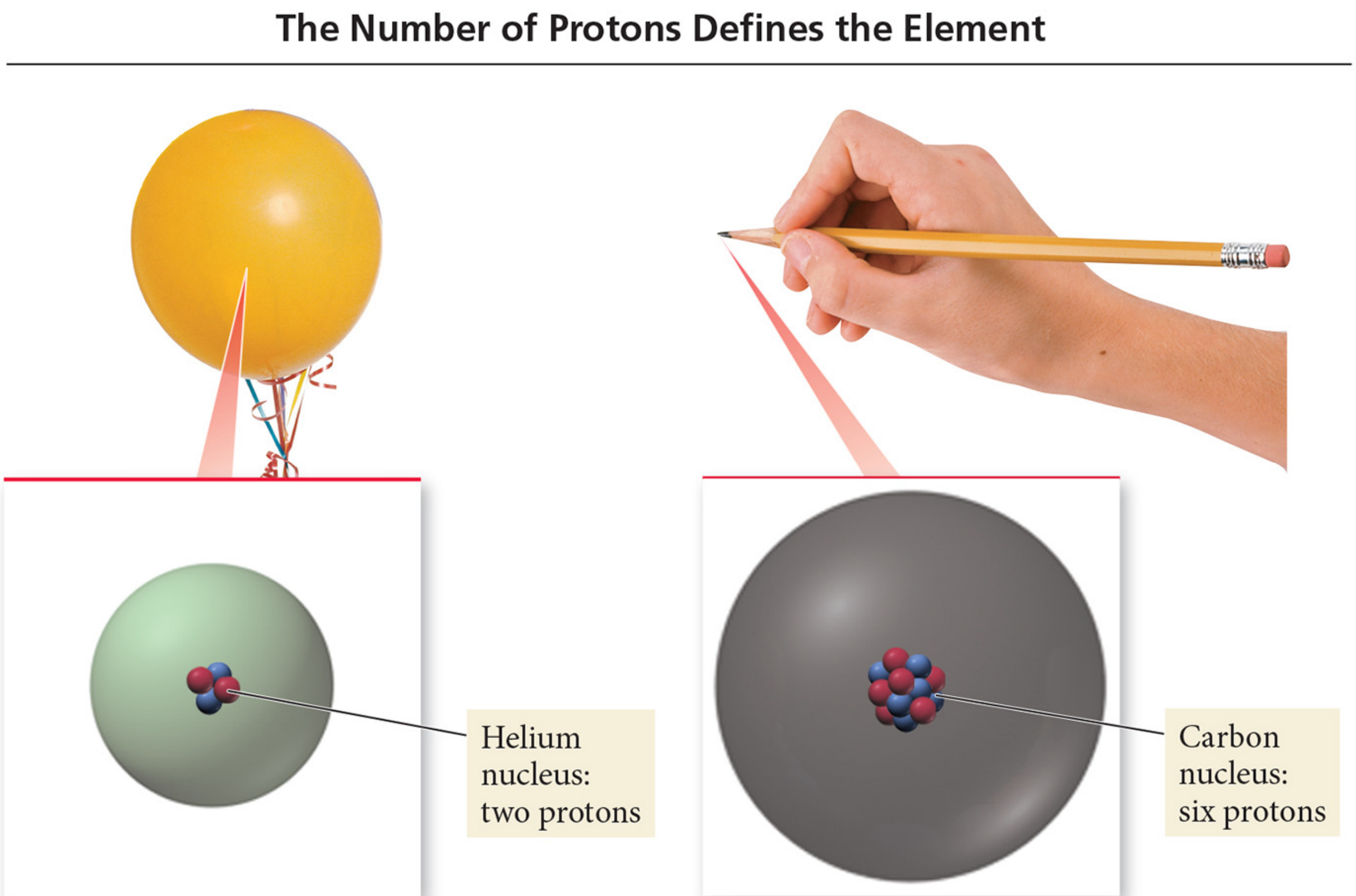
TABLE 2.1 Subatomic Particles

	Mass (kg)	Mass (amu)	Charge (relative)	Charge (C)
Proton	1.67262×10^{-27}	1.00727	+1	$+1.60218 \times 10^{-19}$
Neutron	1.67493×10^{-27}	1.00866	0	0
Electron	0.00091×10^{-27}	0.00055	−1	-1.60218×10^{-19}

Elements: Defined by Their Numbers of Protons

- The most important number to the identity of an atom is the number of protons in its nucleus.
- *The number of protons defines the element*.
- The number of protons in an atom's nucleus is its **atomic number** and is given the symbol ***Z***.

Elements: Defined by Their Numbers of Protons



Periodic Table

The Periodic Table

4 **Be** beryllium	
4	Atomic number (Z)
Be	Chemical symbol
beryllium	Name

1 **H** hydrogen																	2 **He** helium
3 **Li** lithium	4 **Be** beryllium											5 **B** boron	6 **C** carbon	7 **N** nitrogen	8 **O** oxygen	9 **F** fluorine	10 **Ne** neon
11 **Na** sodium	12 **Mg** magnesium											13 **Al** aluminum	14 **Si** silicon	15 **P** phosphorus	16 **S** sulfur	17 **Cl** chlorine	18 **Ar** argon
19 **K** potassium	20 **Ca** calcium	21 **Sc** scandium	22 **Ti** titanium	23 **V** vanadium	24 **Cr** chromium	25 **Mn** manganese	26 **Fe** iron	27 **Co** cobalt	28 **Ni** nickel	29 **Cu** copper	30 **Zn** zinc	31 **Ga** gallium	32 **Ge** germanium	33 **As** arsenic	34 **Se** selenium	35 **Br** bromine	36 **Kr** krypton
37 **Rb** rubidium	38 **Sr** strontium	39 **Y** yttrium	40 **Zr** zirconium	41 **Nb** niobium	42 **Mo** molybdenum	43 **Tc** technetium	44 **Ru** ruthenium	45 **Rh** rhodium	46 **Pd** palladium	47 **Ag** silver	48 **Cd** cadmium	49 **In** indium	50 **Sn** tin	51 **Sb** antimony	52 **Te** tellurium	53 **I** iodine	54 **Xe** xenon
55 **Cs** cesium	56 **Ba** barium	57 **La** lanthanum	72 **Hf** hafnium	73 **Ta** tantalum	74 **W** tungsten	75 **Re** rhenium	76 **Os** osmium	77 **Ir** iridium	78 **Pt** platinum	79 **Au** gold	80 **Hg** mercury	81 **Tl** thallium	82 **Pb** lead	83 **Bi** bismuth	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
87 **Fr** francium	88 **Ra** radium	89 **Ac** actinium	104 **Rf** rutherfordium	105 **Db** dubnium	106 **Sg** seaborgium	107 **Bh** bohrium	108 **Hs** hassium	109 **Mt** meitnerium	110 **Ds** darmstadtium	111 **Rg** roentgenium	112 **Cn** copernicium	113 **	114 **Fl** flerovium	115 **	116 **Lv** livermorium	117 **	118 **
				58 **Ce** cerium	59 **Pr** praseodymium	60 **Nd** neodymium	61 **Pm** promethium	62 **Sm** samarium	63 **Eu** europium	64 **Gd** gadolinium	65 **Tb** terbium	66 **Dy** dysprosium	67 **Ho** holmium	68 **Er** erbium	69 **Tm** thulium	70 **Yb** ytterbium	71 **Lu** lutetium
				90 **Th** thorium	91 **Pa** protactinium	92 **U** uranium	93 **Np** neptunium	94 **Pu** plutonium	95 **Am** americium	96 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fermium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium

Periodic Table

- Each element is identified by a unique atomic number and with a unique **chemical symbol**.
- The chemical symbol is either a one- or two-letter abbreviation listed directly below its atomic number on the periodic table.
  - The chemical symbol for helium is He.
  - The chemical symbol for carbon is C.
  - The chemical symbol for nitrogen is N.

Isotopes: Varied Number of Neutrons

- All atoms of a given element have the same number of protons; however, they do not necessarily have the same number of neutrons.
  – For example, all neon atoms contain 10 protons, but they may contain 10, 11, or 12 neutrons. All three types of neon atoms exist, and each has a slightly different mass.
- Atoms with the same number of protons but a different number of neutrons are called **isotopes**.

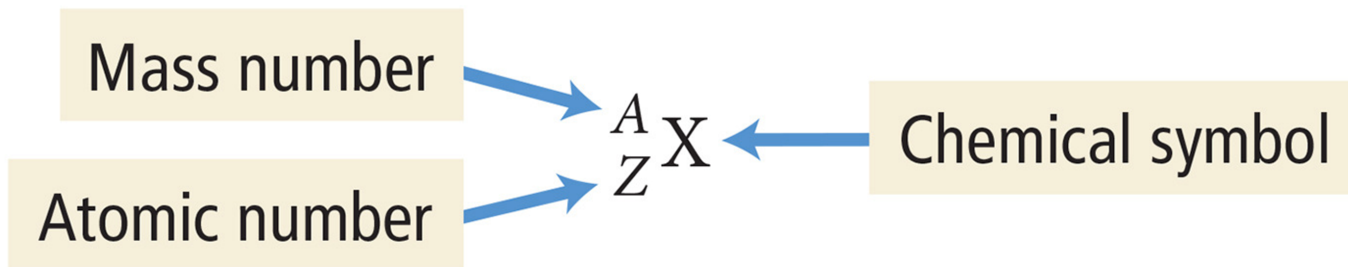
Isotopes: Varied Number of Neutrons

- The relative amount of each different isotope in a naturally occurring sample of a given element is roughly constant.
- These percentages are called the **natural abundance** of the isotopes.
  - Advances in mass spectrometry have allowed accurate measurements that reveal small but significant variations in the natural abundance of isotopes for many elements.

Isotopes: Varied Number of Neutrons

- The sum of the number of neutrons and protons in an atom is its **mass number** and is represented by the symbol ***A***.

A = number of protons (p) + number of neutrons (n)

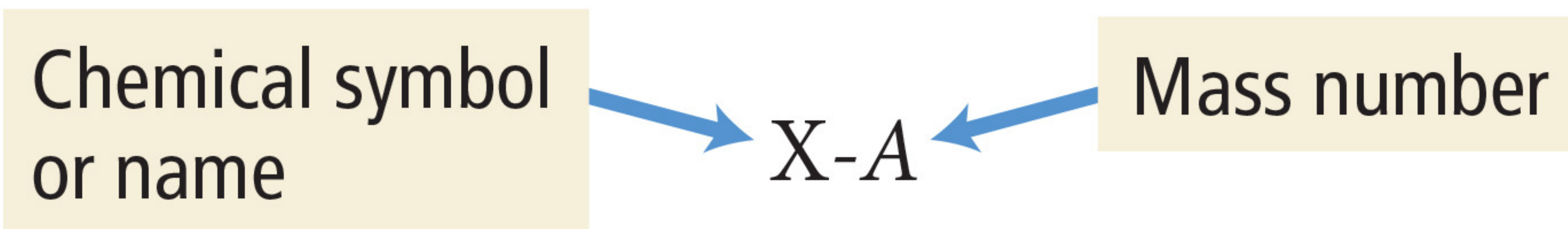


  – X is the chemical symbol, A is the mass number, and Z is the atomic number.

$$^{20}_{10}\mathrm{Ne} \quad ^{21}_{10}\mathrm{Ne} \quad ^{22}_{10}\mathrm{Ne}$$

Isotopes: Varied Number of Neutrons

- A second common notation for isotopes is the chemical symbol (or chemical name) followed by a dash and the mass number of the isotope.



Ne-20 Ne-21 Ne-22

neon-20 neon-21 neon-22

Isotopes: Varied Number of Neutrons

Symbol	Number of Protons	Number of Neutrons	*A* (Mass Number)	Natural Abundance (%)
Ne-20 or $^{20}_{10}Ne$	10	10	20	90.48
Ne-21 or $^{21}_{10}Ne$	10	11	21	0.27
Ne-22 or $^{22}_{10}Ne$	10	12	22	9.25

Ions: Losing and Gaining Electrons

- The number of electrons in a neutral atom is equal to the number of protons in its nucleus (designated by its atomic number *Z*).
- In a chemical change, however, atoms can lose or gain electrons and become charged particles called **ions**.
  - Positively charged ions, such as Na^+, are called **cations**.
  - Negatively charged ions, such as F^-, are called **anions**.

Atomic Mass: The Average Mass of an Element's Atoms

- Atomic mass is sometimes called *atomic weight* or *standard atomic weight.*
- The atomic mass of each element is directly beneath the element's symbol in the periodic table.
- It represents the average mass of the isotopes that compose that element, *weighted according to the natural abundance of each isotope*.

Example: Atomic Mass

- Naturally occurring chlorine consists of 75.77% chlorine-35 atoms (mass 34.97 amu) and 24.23% chlorine-37 atoms (mass 36.97 amu). Calculate its atomic mass.

- Solution:
  - Convert the percent abundance to decimal form and multiply each with its isotopic mass.

    Cl-37 = 0.2423(36.97 amu) = 8.9578 amu

    Cl-35 = 0.7577(34.97 amu) = 26.4968 amu

    Atomic Mass Cl = 8.9578 + 26.4968 = 35.45 amu

Atomic Mass

17
Cl
35.45
chlorine

- In general, we calculate the atomic mass with the following equation:

$$\begin{aligned} \text{Atomic mass} &= \sum_n (\text{fraction of isotope } n) \times (\text{mass of isotope } n) \\ &= (\text{fraction of isotope 1} \times \text{mass of isotope 1}) \\ &+ (\text{fraction of isotope 2} \times \text{mass of isotope 2}) \\ &+ (\text{fraction of isotope 3} \times \text{mass of isotope 3}) + \ \ldots \end{aligned}$$

Molar Mass: Counting Atoms by Weighing Them

- As chemists, we often need to know the number of atoms in a sample of a given mass. Why? *Because chemical processes happen between particles*.

- Therefore, if we want to know the number of atoms in anything of ordinary size, we count them by weighing.

The Mole: A Chemist's "Dozen"

- When we count large numbers of objects, we often use units such as
  - 1 dozen objects = 12 objects.
  - 1 gross objects = 144 objects.
- The chemist's "dozen" is the **mole** (abbreviated mol). A mole is the measure of material containing 6.02214×10^{23} particles:

  $1 \text{ mole} = 6.02214 \times 10^{23}$ particles
- This number is **Avogadro's number**.

The Mole

- First thing to understand about the mole is that it can specify Avogadro's number of anything.

- For example, 1 mol of marbles corresponds to 6.02214×10^{23} marbles.
- 1 mol of sand grains corresponds to 6.02214×10^{23} sand grains.

- *One mole of anything is* 6.02214×10^{23} *units of that thing.*

The Mole

- The second, and more fundamental, thing to understand about the mole is how it gets its specific value.

- **The value of the mole is equal to the number of atoms in exactly 12 grams of pure C-12.**

- **12 g C = 1 mol C atoms = 6.022×10^{23} C atoms**

Converting between Number of Moles and Number of Atoms

- Converting between number of moles and number of atoms is similar to converting between dozens of eggs and number of eggs.
- For atoms, you use the conversion factor 1 mol atoms = 6.022×10^{23} atoms.
- The conversion factors take the following forms:

$$\frac{1\text{ mol atoms}}{6.022 \times 10^{23}\text{ atoms}} \quad \text{or} \quad \frac{6.022 \times 10^{23}\text{ atoms}}{1\text{ mol atoms}}$$

Converting between Mass and Amount (Number of Moles)

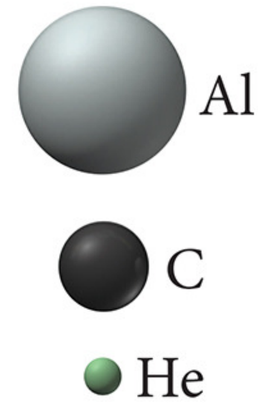
- To count atoms by weighing them, we need one other conversion factor—the mass of 1 mol of atoms.
- The mass of 1 mol of atoms of an element is the **molar mass**.
- **An element's molar mass in grams per mole is numerically equal to the element's atomic mass in atomic mass units (amu).**

Converting between Mass and Moles

26.98 g aluminum = 1 mol aluminum = 6.022×10^{23} Al atoms

12.01 g carbon = 1 mol carbon = 6.022×10^{23} C atoms

4.003 g helium = 1 mol helium = 6.022×10^{23} He atoms



- The lighter the atom, the less mass in 1 mol of atoms.

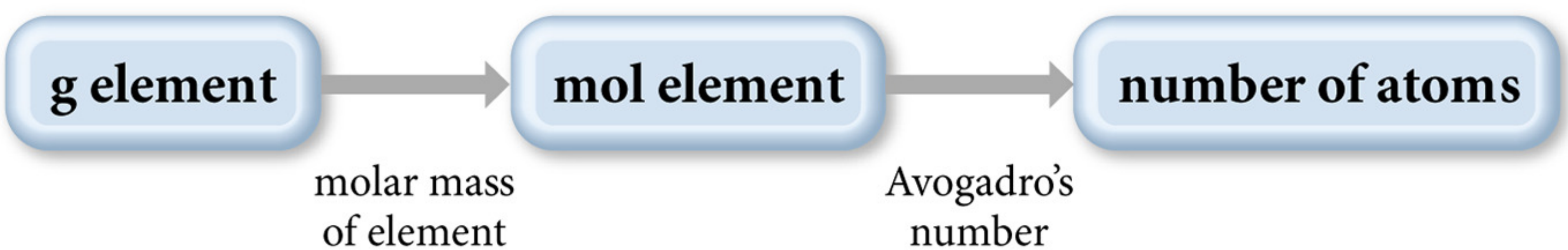
Converting between Mass and Moles

- The molar mass of any element is the conversion factor between the mass (in grams) of that element and the amount (in moles) of that element. For carbon,

$$12.01 \text{ g C} = 1 \text{ mol C} \text{ or } \frac{12.01 \text{ g C}}{\text{mol C}} \text{ or } \frac{1 \text{ mol C}}{12.01 \text{ g C}}$$

Conceptual Plan

- We now have all the tools to count the number of atoms in a sample of an element by weighing it.
  - First, we obtain the mass of the sample.
  - Then, we convert it to the amount in moles using the element's molar mass.
  - Finally, we convert it to the number of atoms using Avogadro's number.
- The conceptual plan for these kinds of calculations takes the following form:

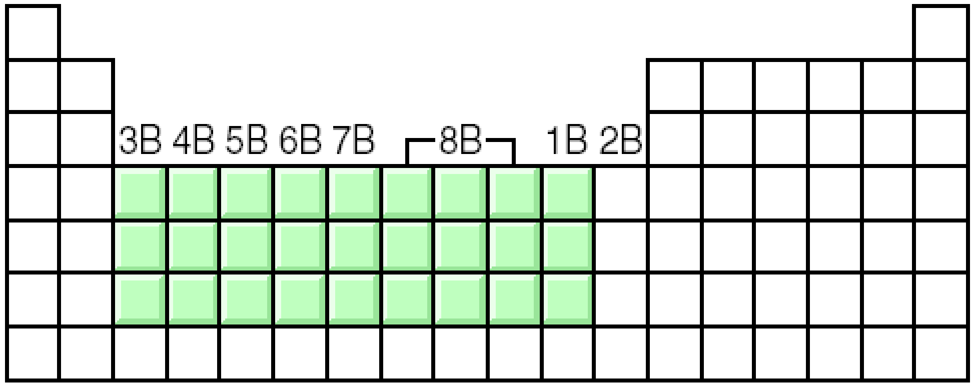


Chemical Nomenclature

- **Ionic Compounds**
  - Often a metal + nonmetal
  - Anion (nonmetal), add "ide" to element name

$BaCl_2$	barium chloride
K_2O	potassium oxide
$Mg(OH)_2$	magnesium hydroxide
KNO_3	potassium nitrate

- Transition metal ionic compounds
  - indicate charge on metal with Roman numerals



$FeCl_2$	2 Cl^- -2 so Fe is +2	iron(II) chloride
$FeCl_3$	3 Cl^- -3 so Fe is +3	iron(III) chloride
Cr_2S_3	3 S^{-2} -6 so Cr is +3 (6/2)	chromium(III) sulfide

TABLE 2.2 The "-ide" Nomenclature of Some Common Monatomic Anions According to Their Positions in the Periodic Table

Group 4A	Group 5A	Group 6A	Group 7A
C carbide (C^{4-})*	N nitride (N^{3-})	O oxide (O^{2-})	F fluoride (F^-)
Si silicide (Si^{4-})	P phosphide (P^{3-})	S sulfide (S^{2-})	Cl chloride (Cl^-)
		Se selenide (Se^{2-})	Br bromide (Br^-)
		Te telluride (Te^{2-})	I iodide (I^-)

*The word "carbide" is also used for the anion C_2^{2-}.

TABLE 2.3 Names and Formulas of Some Common Inorganic Cations and Anions

Cation	Anion
aluminum (Al^{3+})	bromide (Br^-)
ammonium (NH_4^+)	carbonate (CO_3^{2-})
barium (Ba^{2+})	chlorate (ClO_3^-)
cadmium (Cd^{2+})	chloride (Cl^-)
calcium (Ca^{2+})	chromate (CrO_4^{2-})
cesium (Cs^+)	cyanide (CN^-)
chromium(III) or chromic (Cr^{3+})	dichromate ($Cr_2O_7^{2-}$)
cobalt(II) or cobaltous (Co^{2+})	dihydrogen phosphate ($H_2PO_4^-$)
copper(I) or cuprous (Cu^+)	fluoride (F^-)
copper(II) or cupric (Cu^{2+})	hydride (H^-)
hydrogen (H^+)	hydrogen carbonate or bicarbonate (HCO_3^-)
iron(II) or ferrous (Fe^{2+})	hydrogen phosphate (HPO_4^{2-})
iron(III) or ferric (Fe^{3+})	hydrogen sulfate or bisulfate (HSO_4^-)
lead(II) or plumbous (Pb^{2+})	hydroxide (OH^-)
lithium (Li^+)	iodide (I^-)
magnesium (Mg^{2+})	nitrate (NO_3^-)
manganese(II) or manganous (Mn^{2+})	nitride (N^{3-})
mercury(I) or mercurous (Hg_2^{2+})*	nitrite (NO_2^-)
mercury(II) or mercuric (Hg^{2+})	oxide (O^{2-})
potassium (K^+)	permanganate (MnO_4^-)
rubidium (Rb^+)	peroxide (O_2^{2-})
silver (Ag^+)	phosphate (PO_4^{3-})
sodium (Na^+)	sulfate (SO_4^{2-})
strontium (Sr^{2+})	sulfide (S^{2-})
tin(II) or stannous (Sn^{2+})	sulfite (SO_3^{2-})
zinc (Zn^{2+})	thiocyanate (SCN^-)

*Mercury(I) exists as a pair as shown.

- **Molecular compounds**
  - Nonmetals or nonmetals + metalloids
  - Common names
    - H_2O, NH_3, CH_4,
  - Element furthest to the left in a period and closest to the bottom of a group on periodic table is placed first in formula
  - If more than one compound can be formed from the same elements, use prefixes to indicate number of each kind of atom
  - Last element name ends in *ide*

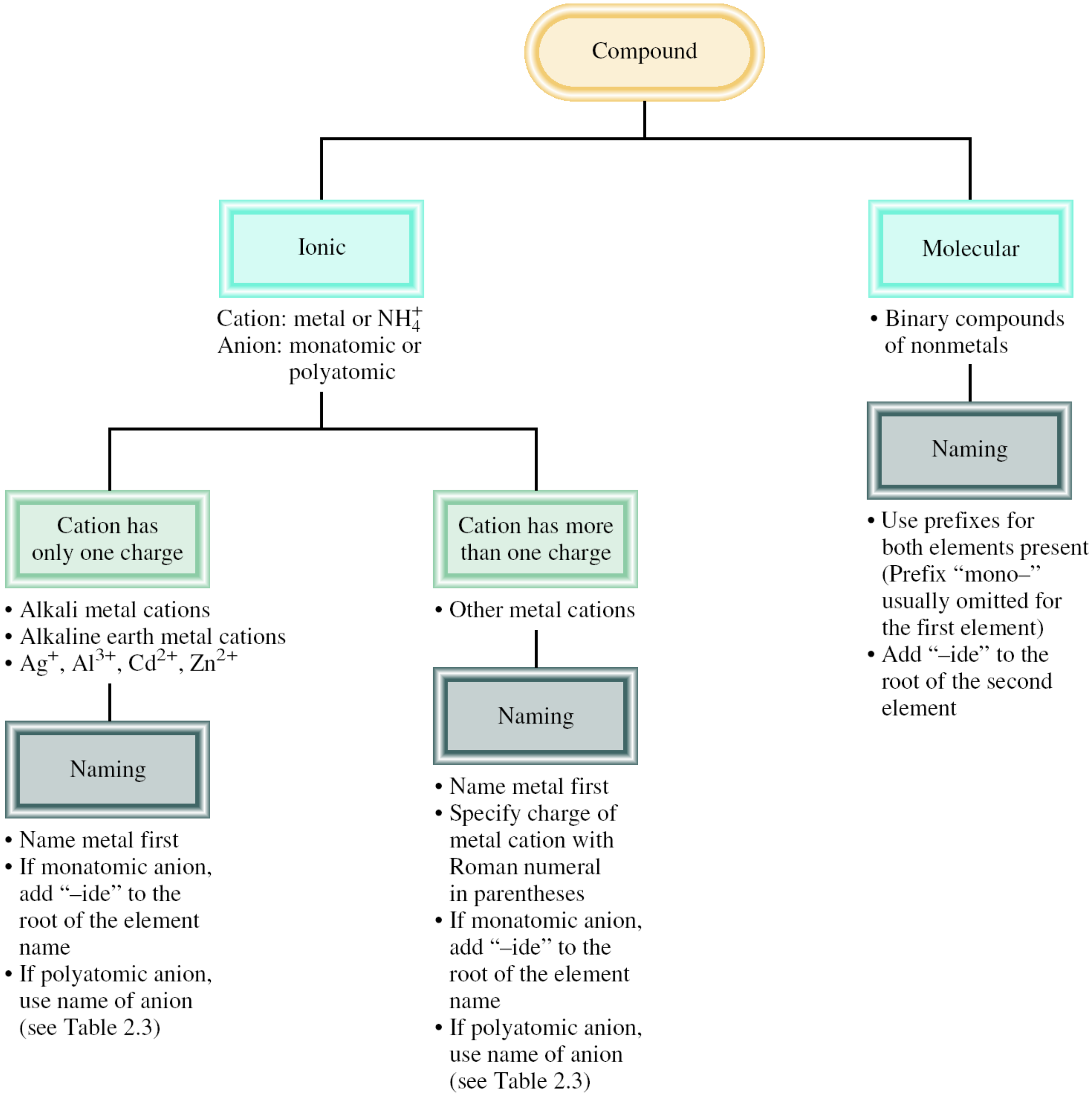
TABLE 2.4

Greek Prefixes Used in Naming Molecular Compounds

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

Molecular Compounds

HI	hydrogen iodide
NF_3	nitrogen trifluoride
SO_2	sulfur dioxide
N_2Cl_4	dinitrogen tetrachloride
NO_2	nitrogen dioxide
N_2O	dinitrogen monoxide


Compound
Ionic
Cation: metal or NH_4^+
Anion: monatomic or polyatomic
Cation has only one charge
• Alkali metal cations
• Alkaline earth metal cations
• Ag^+, Al^{3+}, Cd^{2+}, Zn^{2+}
Naming
• Name metal first
• If monatomic anion, add "–ide" to the root of the element name
• If polyatomic anion, use name of anion (see Table 2.3)
Cation has more than one charge
• Other metal cations
Naming
• Name metal first
• Specify charge of metal cation with Roman numeral in parentheses
• If monatomic anion, add "–ide" to the root of the element name
• If polyatomic anion, use name of anion (see Table 2.3)
Molecular
• Binary compounds of nonmetals
Naming
• Use prefixes for both elements present (Prefix "mono–" usually omitted for the first element)
• Add "–ide" to the root of the second element

An ***acid*** can be defined as a substance that yields hydrogen ions (H^+) when dissolved in water.

For example: HCl gas and HCl in water

- Pure substance, hydrogen chloride

- Dissolved in water (H_3O^+ and Cl^-), hydrochloric acid

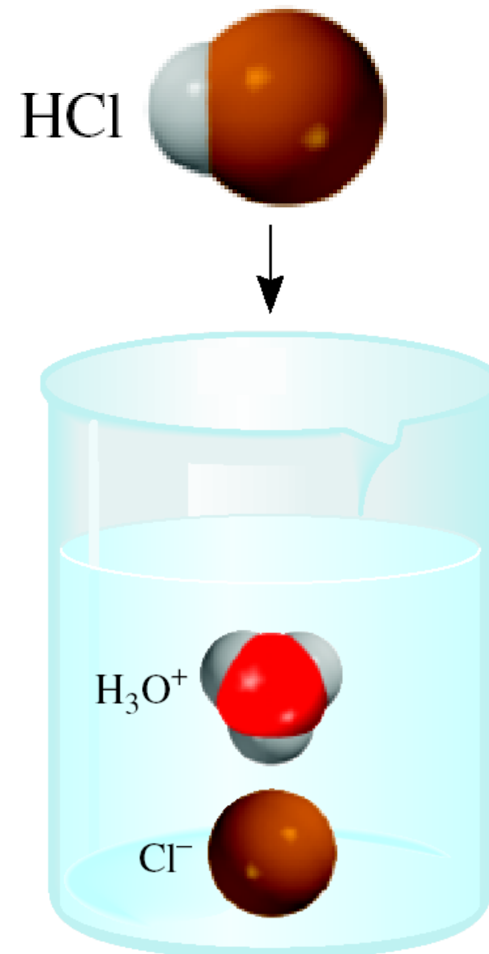
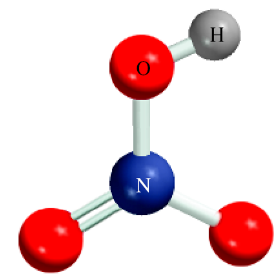
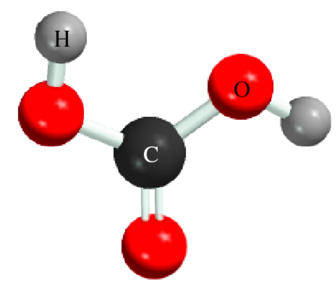
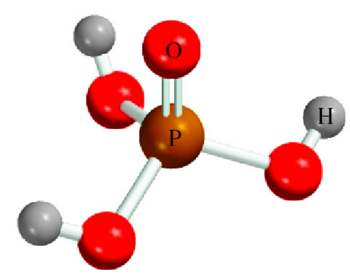


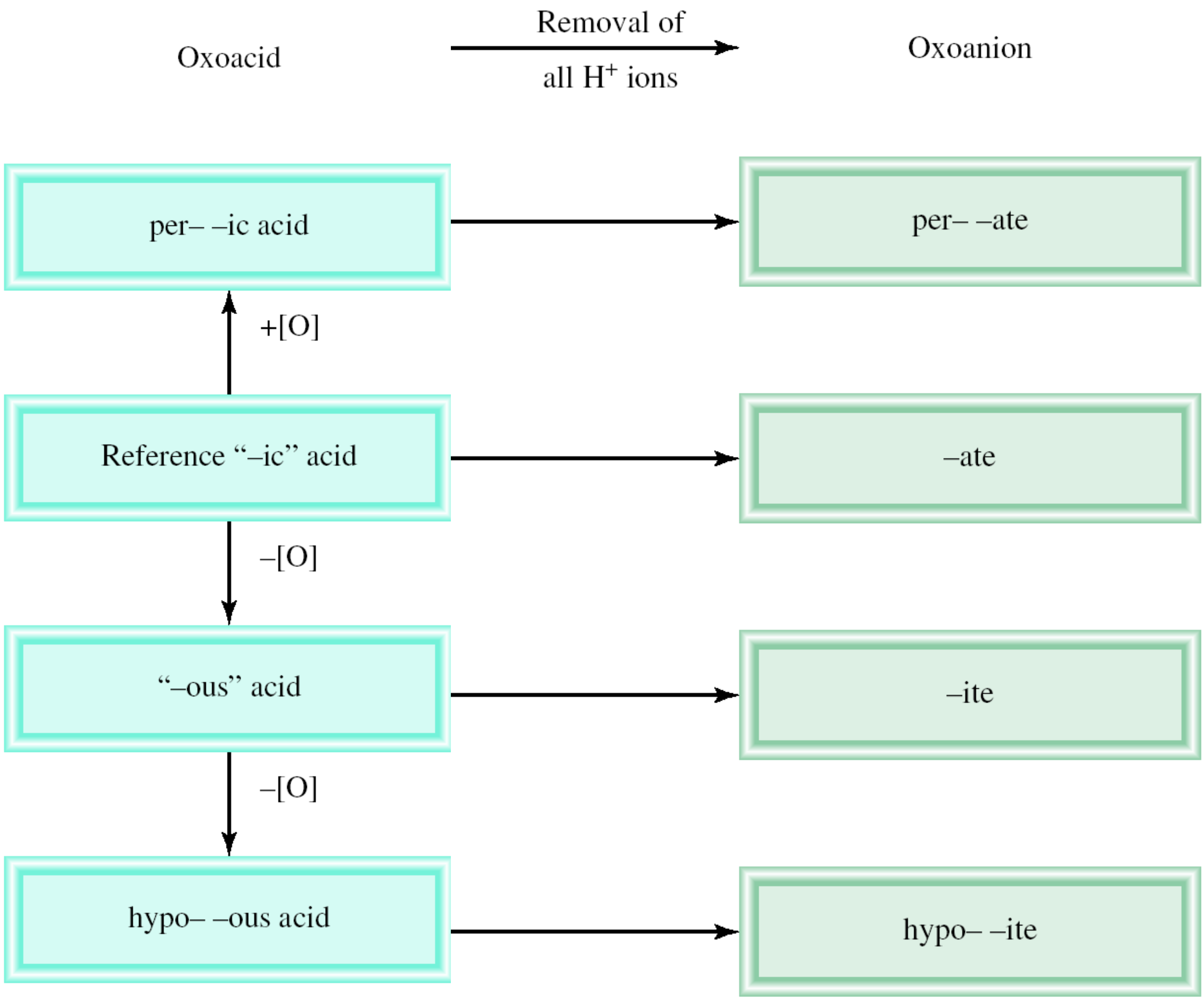
TABLE 2.5 Some Simple Acids

Anion	Corresponding Acid
F^- (fluoride)	HF (hydrofluoric acid)
Cl^- (chloride)	HCl (hydrochloric acid)
Br^- (bromide)	HBr (hydrobromic acid)
I^- (iodide)	HI (hydroiodic acid)
CN^- (cyanide)	HCN (hydrocyanic acid)
S^{2-} (sulfide)	H_2S (hydrosulfuric acid)

An ***oxoacid*** is an acid that contains hydrogen, oxygen, and another element.

HNO_3	nitric acid	 HNO_3
H_2CO_3	carbonic acid	 H_2CO_3
H_3PO_4	phosphoric acid	 H_3PO_4

Naming Oxoacids and Oxoanions



The rules for naming ***oxoanions,*** *anions of oxoacids,* are as follows:

1. When all the H ions are removed from the "-ic" acid, the anion's name ends with "-ate."
2. When all the H ions are removed from the "-ous" acid, the anion's name ends with "-ite."
3. The names of anions in which one or more but not all the hydrogen ions have been removed must indicate the number of H ions present.

   For example:

   – $H_2PO_4^-$ dihydrogen phosphate
   – HPO_4^{2-} hydrogen phosphate
   – PO_4^{3-} phosphate

TABLE 2.6 Names of Oxoacids and Oxoanions That Contain Chlorine

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