CHEMISTRY

Chapter 2 ATOMS, MOLECULES, AND IONS

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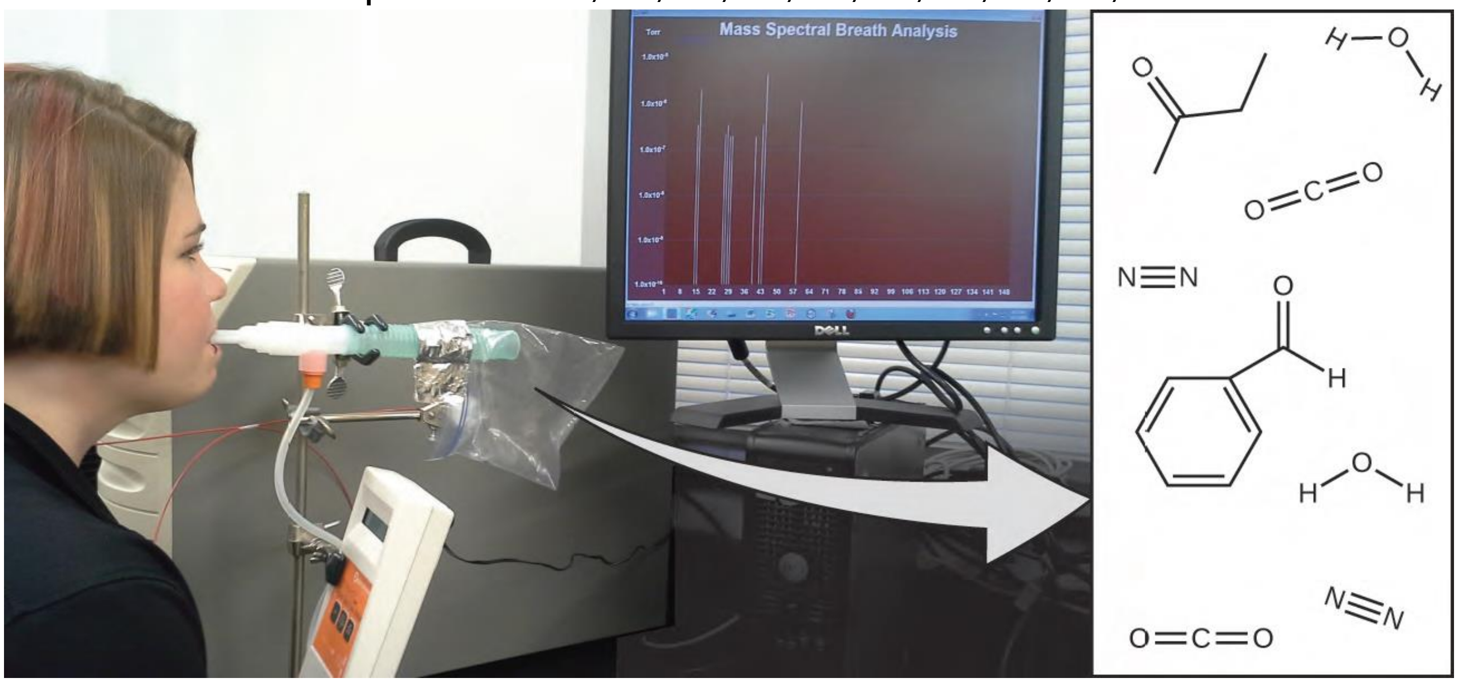
The Cooper Union

HW problems: 11, 18, 23, 31, 43, 47, 49, 55, 60, 61



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CH. 2 OUTLINE

2.1: Early Ideas in Atomic Theory

2.2: Evolution of Atomic Theory

2.3: Atomic Structure and Symbolism

2.4: Chemical Formulas

2.5: The Periodic Table

2.6: Molecular and Ionic Compounds

2.7: Chemical Nomenclature

2.1 EARLY IDEAS IN ATOMIC THEORY

- The concept of atoms was first proposed by the Greek philosophers, Leucippus and Democritus, in the fifth century BC.
    - ***atomos***, a term derived from the Greek word for "indivisible".

- Later, Aristotle and others believed that matter consisted of various combinations of the four "elements"—fire, earth, air, and water.

- In 1807, English schoolteacher, John Dalton, proposed his atomic theory.

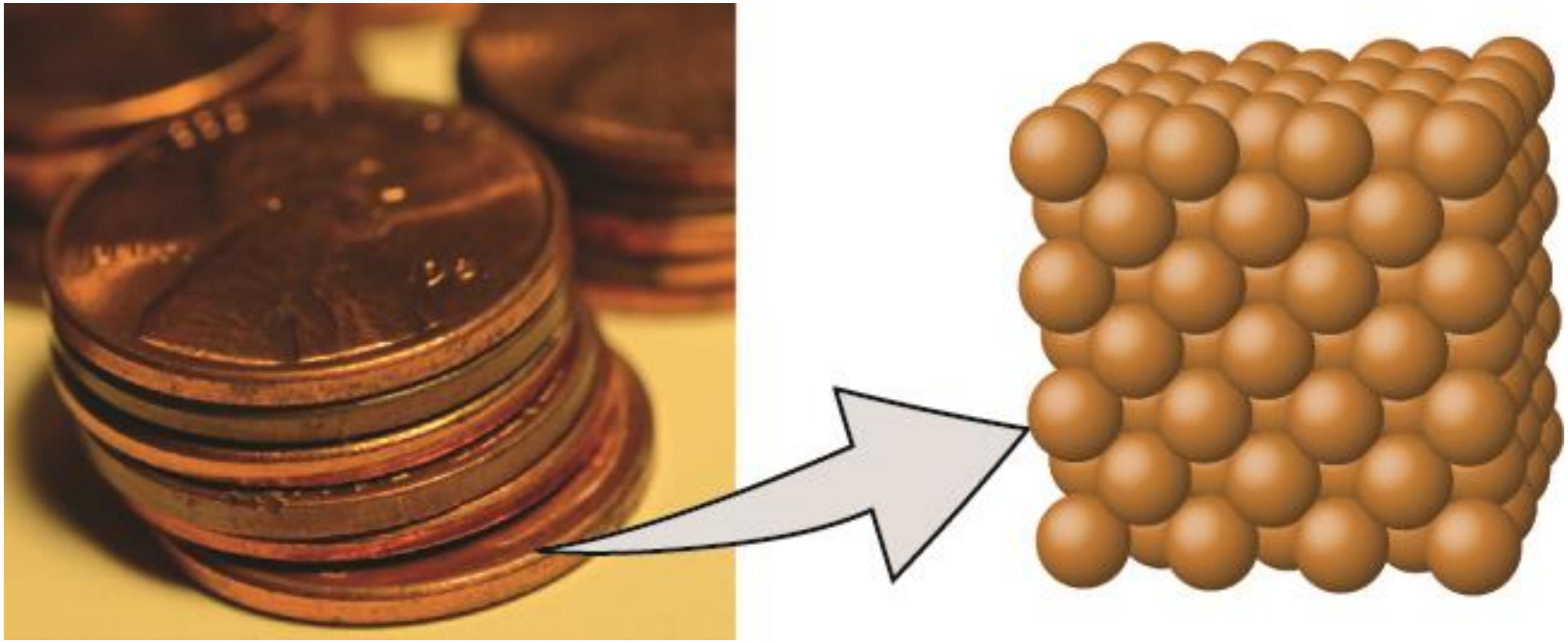
DALTON’S ATOMIC THEORY

- **Dalton’s atomic theory can be summarized in five postulates.**

1) ***Matter*** is composed of exceedingly small particles called atoms. An atom is the smallest unit of an element that can participate in a chemical change.

1) ***An element*** consists of only one type of atom, which has a mass that is characteristic of the element and is the same for all atoms of that element.

FIGURE 2.2



A pre-1982 copper penny (left) contains approximately 3×10^{22} copper atoms (several dozen are represented as brown spheres at the right), each of which has the same chemical properties. (credit: modification of work by “slgckgc”/Flickr)

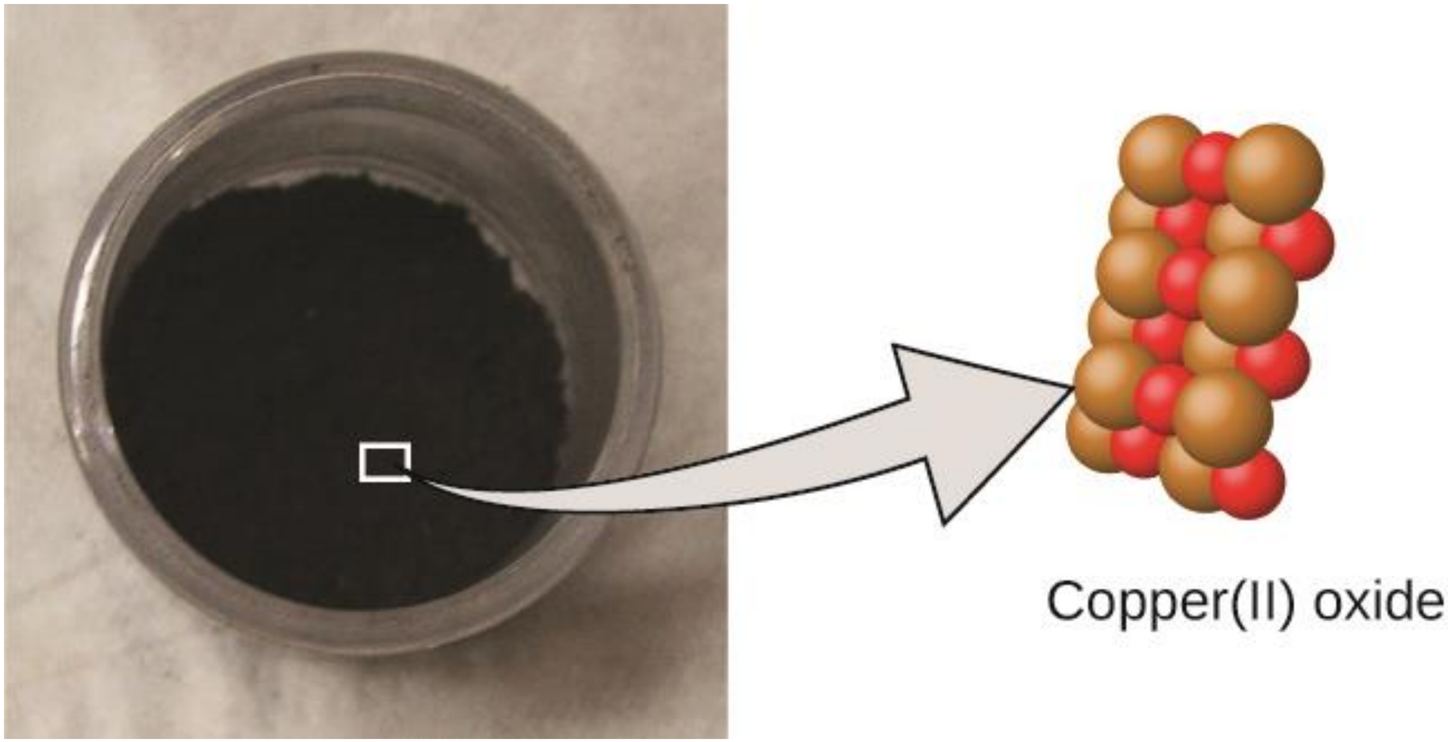
DALTON'S ATOMIC THEORY

3) Atoms of one element differ in properties from atoms of all other elements.

4) A ***Compound*** consists of atoms of two or more elements combined in a small, whole-number ratio. In a given compound, the number of atoms of each of its elements are always present in the same ratio.

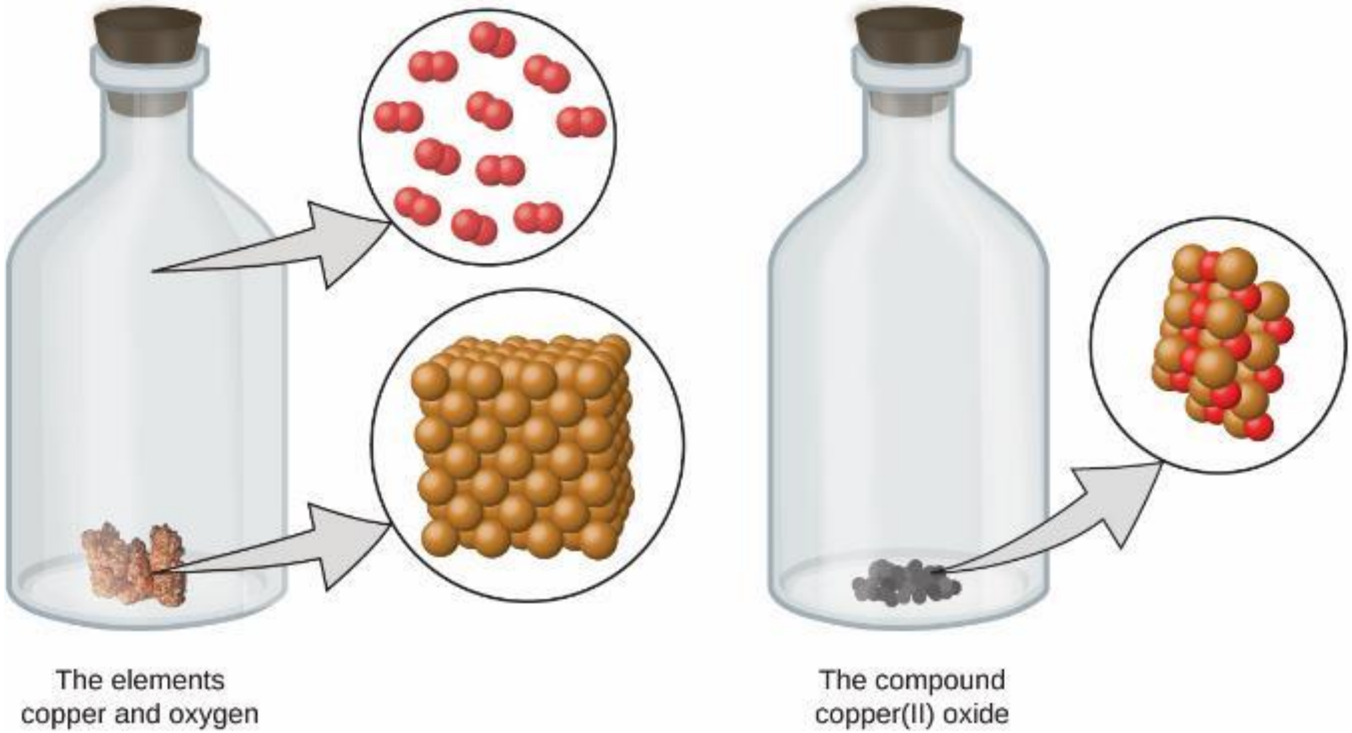
5) Atoms are neither created nor destroyed during a chemical change, but instead rearrange to yield a different type(s) of matter.

FIGURE 2.3



Copper(II) oxide, a powdery, black compound, results from the combination of two types of atoms—copper (brown spheres) and oxygen (red spheres)—in a 1:1 ratio. (credit: modification of work by “Chemicalinterest”/Wikimedia Commons)

FIGURE 2.4



When the elements copper (a shiny, red-brown solid, shown here as brown spheres) and oxygen (a clear and colorless gas, shown here as red spheres) react, their atoms rearrange to form a compound containing copper and oxygen (a powdery, black solid). (credit copper: modification of work by http://imagesof-elements.com/copper.php)

LAW OF CONSERVATION OF MATTER

- Dalton’s atomic theory provides a microscopic explanation of the many macroscopic properties of matter that you’ve learned about.

- If atoms are neither created nor destroyed during a chemical change, then the total mass of matter present when matter changes from one type to another will remain constant ***(the law of conservation of matter).***

LAW OF DEFINITE PROPORTIONS

- ***Law of definite proportions or the law of constant composition -*** All samples of a pure compound contain the same elements in the same proportion by mass.
  - Illustrated by experiments performed by French chemist, Joseph Proust.

Constant Composition of Isooctane

Sample	Carbon	Hydrogen	Mass Ratio
A	14.82 g	2.78 g	5.33 g carbon/1.00 g hydrogen
B	22.33 g	4.19 g	5.33 g carbon/1.00 g hydrogen
C	19.40 g	3.64 g	5.33 g carbon/1.00 g hydrogen

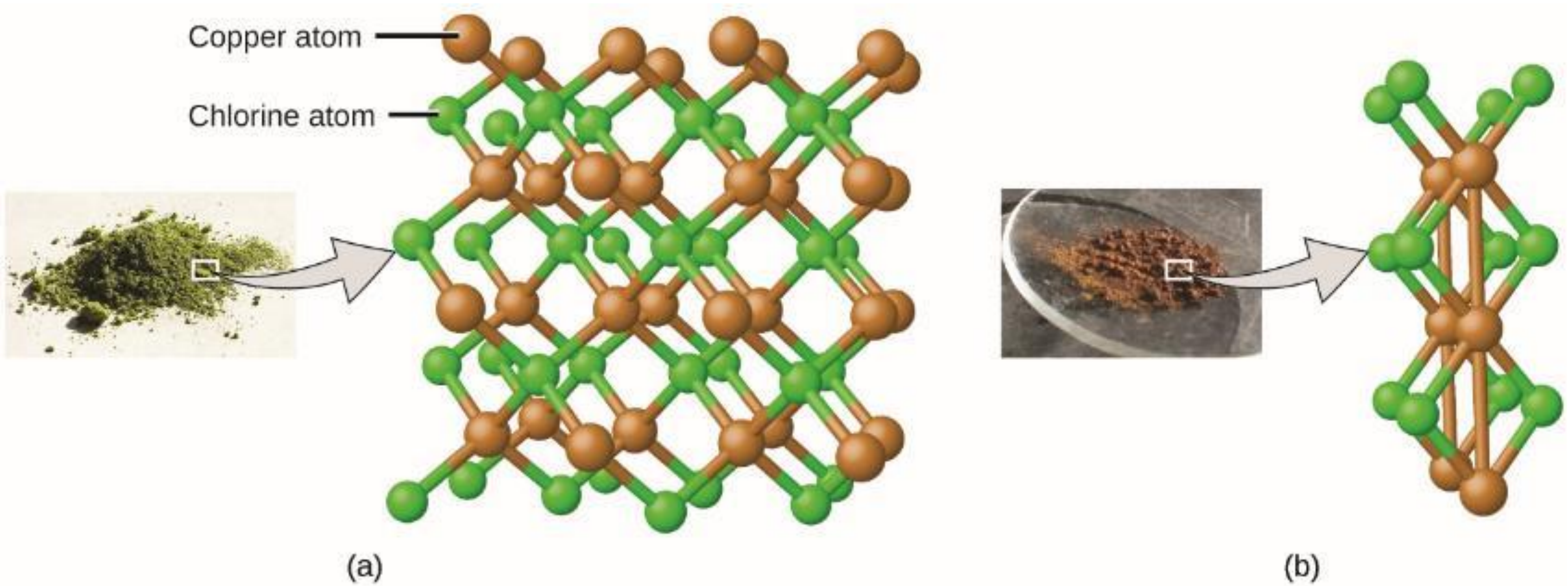
LAW OF MULTIPLE PROPORTIONS

- ***The law of multiple proportions*** states that when two elements react to form more than one compound, a fixed mass of one element will react with masses of the other element in a ratio of small, whole numbers.

- For example: compounds containing chlorine and copper.
  - A green solid contains 0.558 g Cl to 1 g Cu.
  - A brown solid contains 1.116 g Cl to 1 g Cu.

$$\frac{\dfrac{1.116\ g\ Cl}{1\ g\ Cu}}{\dfrac{0.558\ g\ Cl}{1\ g\ Cu}} = \frac{2}{1}$$

FIGURE 2.5



Compared to the copper chlorine compound in (a), where copper is represented by brown spheres and chlorine by green spheres, the copper chlorine compound in (b) has twice as many chlorine atoms per copper atom. (credit a: modification of work by "Benjah-bmm27"/Wikimedia Commons; credit b: modification of work by "Walkerma"/Wikimedia Commons)

2.2 EVOLUTION OF ATOMIC THEORY

- In the two centuries since Dalton developed his ideas, scientists have made significant progress in furthering our understanding of atomic theory.

- What were atoms composed of?

- Was there something smaller than an atom?

- Here, we will discuss some of these key developments.

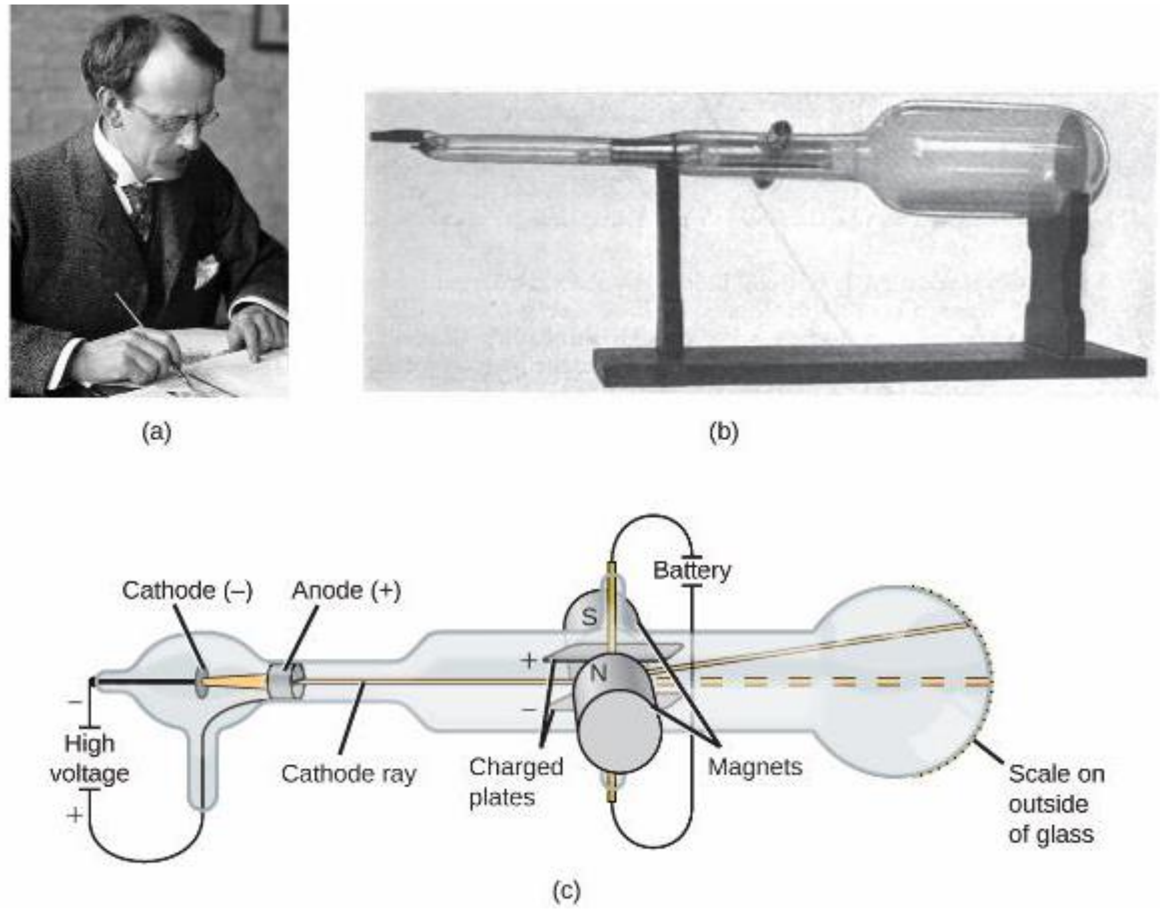
DISCOVERY OF THE ELECTRON

- **J.J. Thomson experimented with cathode ray tubes.**
- **Cathode ray tube:**
  - A sealed glass tube from which almost all the air had been removed.
  - Contained two metal electrodes.
  - When a high voltage was applied across the electrodes, a visible beam called a cathode ray appeared between them.
  - Regardless of the metals used, this beam was always deflected toward the positive charge and away from the negative charge.
  - Thompson was able to calculate the charge-to-mass ratio of the cathode ray particles.

DISCOVERY OF THE ELECTRON

- **Thompson's results:**
  - The cathode ray particles were much lighter than atoms.
  - These particles are negatively charged.
  - These particles are indistinguishable, regardless of the source material.
  - This cathode ray particle is what we now call an **electron -** a negatively charged, subatomic particle with a mass more than one thousand times less than that of an atom.

FIGURE 2.6



(a) J. J. Thomson produced a visible beam in a cathode ray tube.

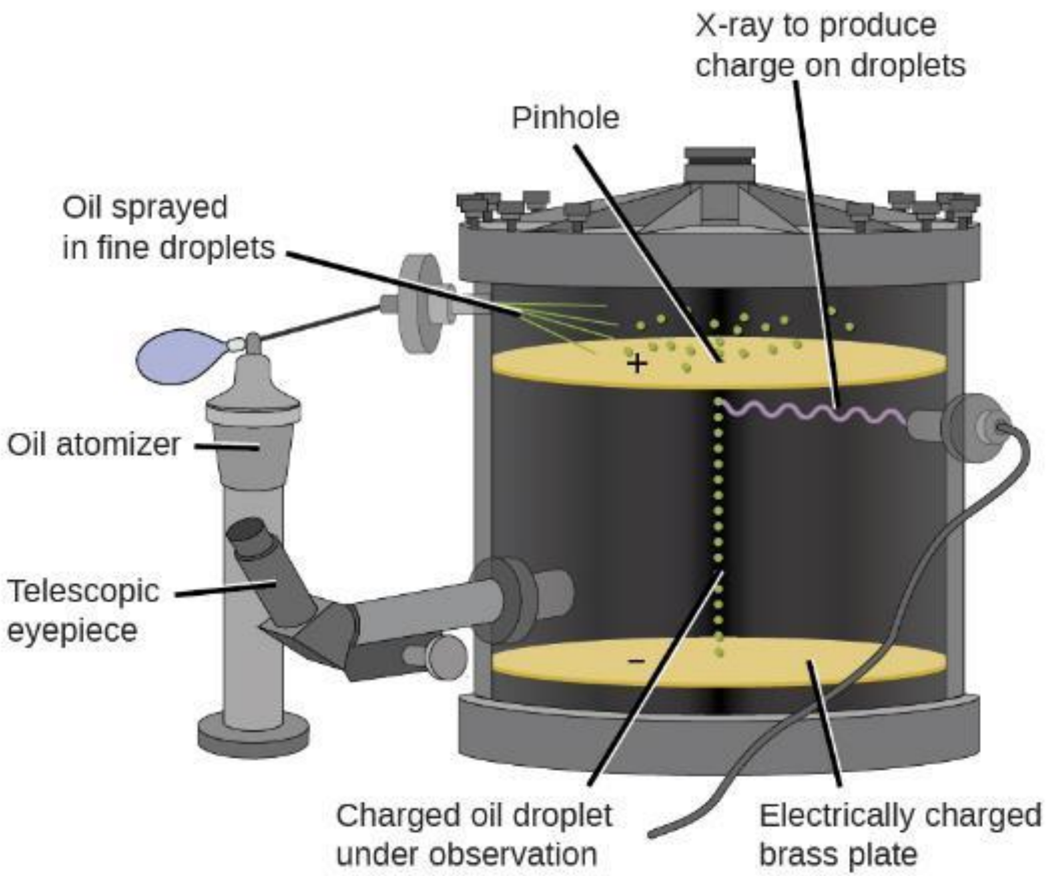
(b) This is an early cathode ray tube, invented in 1897 by Ferdinand Braun.

(c) In the cathode ray, the beam (shown in yellow) comes from the cathode and is accelerated past the anode toward a fluorescent scale at the end of the tube. Simultaneous deflections by applied electric and magnetic fields permitted Thomson to calculate the mass-to-charge ratio of the particles composing the cathode ray. (credit a: modification of work by Nobel Foundation; credit b: modification of work by Eugen Nesper; credit c: modification of work by “Kurzon”/Wikimedia Commons)

DISCOVERY OF THE ELECTRON

- **Robert A. Millikan's Oil Drop Experiment (1909)**
  - Millikan created microscopic oil droplets, which were electrically charged.
  - These drops could be slowed or reversed by an electric field.
  - Millikan was able to determine the charge on individual drops.

FIGURE 2.7



Oil drop	Charge in coulombs (C)
A	4.8×10^{-19} C
B	3.2×10^{-19} C
C	6.4×10^{-19} C
D	1.6×10^{-19} C
E	4.8×10^{-19} C

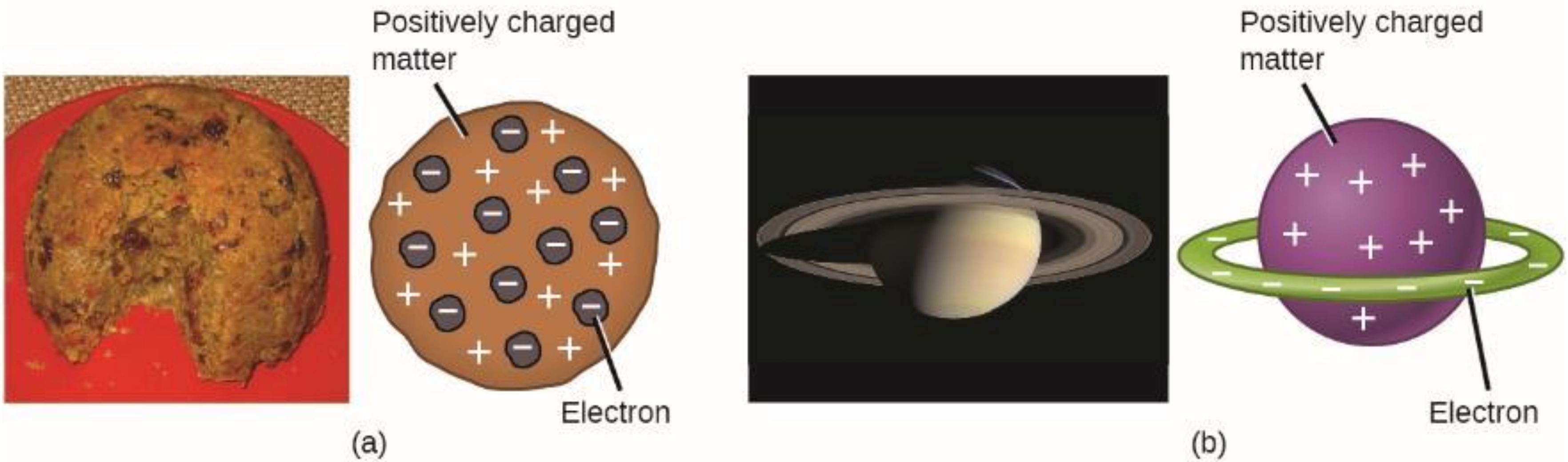
Millikan's experiment measured the charge of individual oil drops. The tabulated data are examples of a few possible values.

DISCOVERY OF THE ELECTRON

- **Millikan's results:**
  - The charge of an oil drop was always a multiple of a specific charge, 1.6×10^{-19} C.
  - Millikan concluded that 1.6×10^{-19} C was the charge of a single electron.
  - Thompson already showed the charge to mass ratio of an electron to be 1.759×10^{11} C/kg.

$$Mass\ of\ electron = 1.602 \times 10^{-19}\,C \left(\frac{1\,kg}{1.759 \times 10^{11}\,C} \right) = 9.107 \times 10^{-31}\,kg$$

FIGURE 2.8



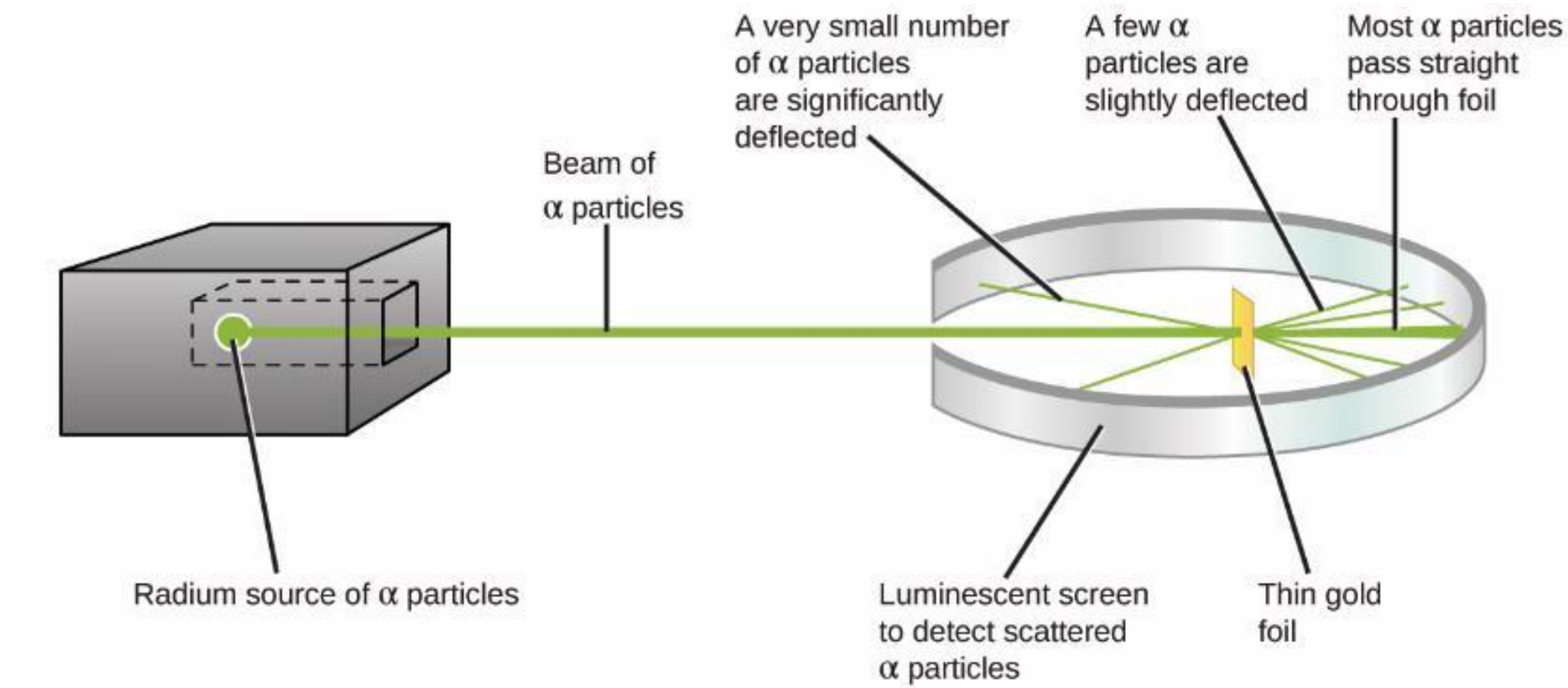
(a) Thomson suggested that atoms resembled plum pudding, an English dessert consisting of moist cake with embedded raisins (“plums”).

(b) Nagaoka proposed that atoms resembled the planet Saturn, with a ring of electrons surrounding a positive “planet.” (credit a: modification of work by “Man vyi”/Wikimedia Commons; credit b: modification of work by “NASA”/Wikimedia Commons)

DISCOVERY OF THE NUCLEUS

- **Ernest Rutherford's Gold Foil Scattering Experiment:**
  - Aimed a beam of alpha particles (α particles) at a very thin piece of gold foil.
  - α particles are positively charged.
  - The scattering of these α particles was examined using a luminescent screen that would glow briefly when hit.

FIGURE 2.9

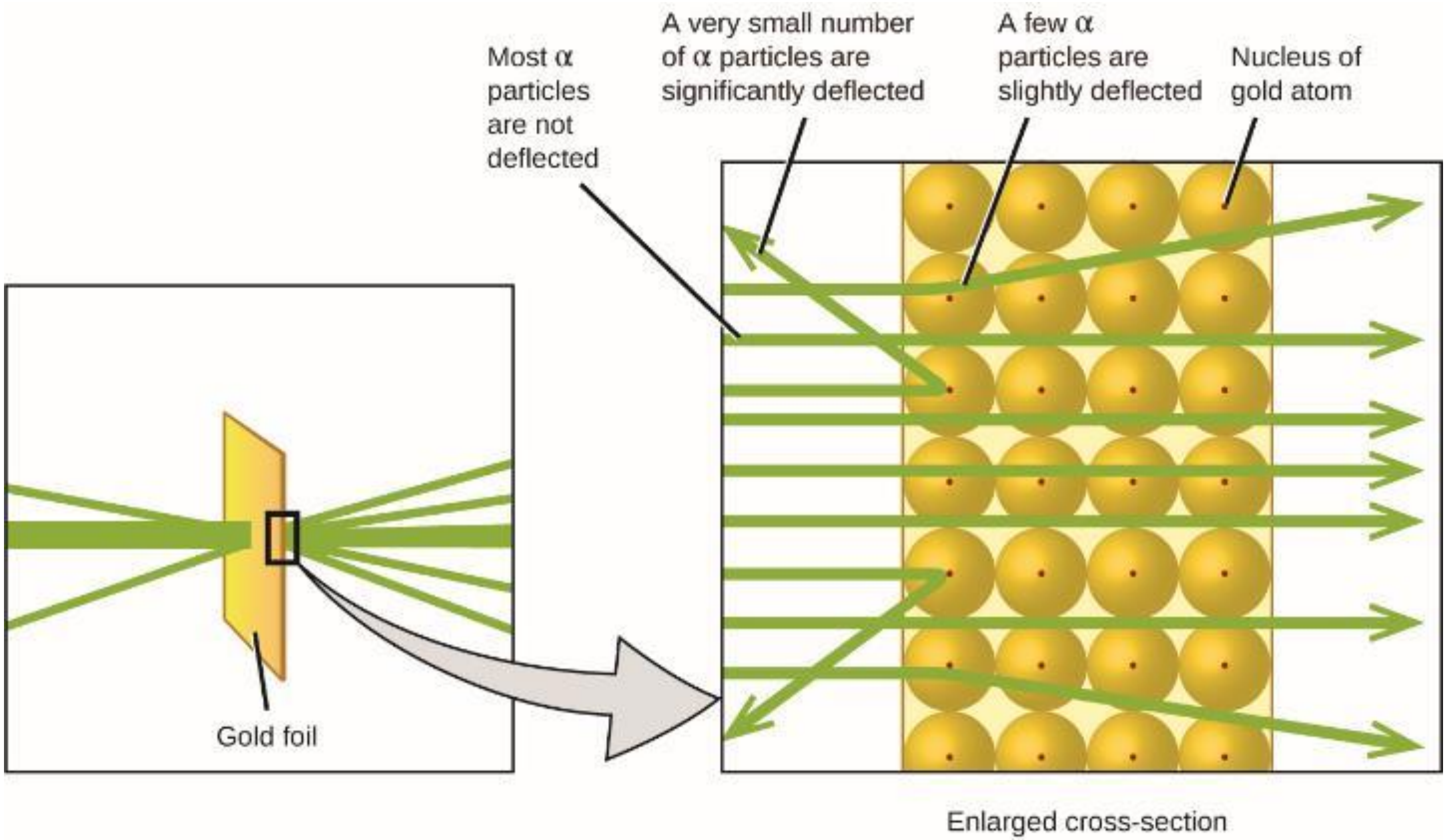


Geiger and Rutherford fired α particles at a piece of gold foil and detected where those particles went, as shown in this schematic diagram of their experiment. Most of the particles passed straight through the foil, but a few were deflected slightly and a very small number were significantly deflected.

DISCOVERY OF THE NUCLEUS

- **Rutherford's results:**
  - The volume occupied by an atom must consist of a large amount of empty space.
  - A small, relatively heavy, positively charged body, the ***nucleus***, must be at the center of each atom.
  - The nucleus contains most of the atom's mass.
  - Negatively charged electrons surround the nucleus.
  - The ***proton***, a positively charged, subatomic particle is located in the nucleus.

FIGURE 2.10



The α particles are deflected only when they collide with or pass close to the much heavier, positively charged gold nucleus. Because the nucleus is very small compared to the size of an atom, very few α particles are deflected. Most pass through the relatively large region occupied by electrons, which are too light to deflect the rapidly moving particles.

OTHER IMPORTANT DISCOVERIES OF THE 20TH CENTURY

- ***Isotopes* —** atoms of the same element that differ in mass.
  - Frederick Soddy of England. Noble Prize in 1921.

- ***Neutrons* —** uncharged, subatomic particles with a mass approximately the same as that of protons.
  - Discovered by James Chadwick in 1932.
  - Neutrons are also found in the nucleus.

2.3 ATOMIC STRUCTURE AND SYMBOLISM

- The nucleus contains the majority of an atom's mass.
- Protons and neutrons are much heavier than electrons.
- Electrons occupy almost all of an atom's volume.
- Diameter of an atom ~ 10^{-10} m
- Diameter of a nucleus is 100,000 times smaller ~ 10^{-15} m

FIGURE 2.11

If an atom could be expanded to the size of a football stadium, the nucleus would be the size of a single blueberry. (credit middle: modification of work by “babyknight”/Wikimedia Commons; credit right: modification of work by Paxson Woelber)

UNITS

- Atoms and subatomic particles are very small.
  - Example: A carbon atom weighs less than 2×10^{-23} g.

- Electrons have a charge of less than 2×10^{-19} C.

- Small units are needed.
  - Atomic mass unit (amu).
    - 1 amu = 1.6605×10^{-24} g.
    - Mass of a carbon-12 atom = 12 amu

  - Fundamental unit of charge (e).
    - $e = 1.602 \times 10^{-19}$ C

PROPERTIES OF SUBATOMIC PARTICLES

- **Proton**
  - Mass = 1.0073 amu
  - Charge = +1
- **Neutron**
  - Mass = 1.0087 amu (slightly heavier than a proton)
  - Charge = 0
- **Electron**
  - Mass = 0.00055 amu
  - Charge = -1

ATOMIC NUMBER (Z)

- The number of protons in the nucleus of an atom is its ***atomic number (Z).***

- This is the defining trait of an element: Its value determines the identity of the atom.

- For example, any atom that contains six protons is the element carbon and has the atomic number 6, regardless of how many neutrons or electrons it may have.

NEUTRAL ATOMS

- A neutral atom must contain the same number of positive and negative charges.

- The number of protons equals the number of electrons.

- Therefore, the atomic number also indicates the number of electrons in a neutral atom.

MASS NUMBER (A)

- The total number of protons and neutrons in an atom is called its ***mass number (A).***

- The number of neutrons is therefore the difference between the mass number and the atomic number.

atomic number (Z) = number of protons

mass number (A) = number of protons + number of neutrons

A – Z = number of neutrons

IONS

• When the number of protons and electrons are ***NOT equal***, the atom is electrically charged and called an ***ion***.

Charge of an atom = number of protons – number of electrons

• Atoms (and molecules) acquire charge by losing or gaining electrons.

CATIONS AND ANIONS

- An atom that gains one or more electrons will exhibit a ***negative charge*** and is called an ***anion.***
  - ***Example:*** A neutral oxygen atom (Z = 8) has eight electrons, and if it gains two electrons it will become an anion with a 2− charge (8 − 10 = 2−).

- An atom that loses one or more electrons will exhibit a ***positive charge*** and is called an ***cation.***
  - ***Example:*** a neutral sodium atom (Z = 11) has 11 electrons. If this atom loses one electron, it will become a cation with a 1+ charge (11 − 10 = 1+).

CHEMICAL SYMBOLS

- A chemical symbol is an abbreviation that we use to indicate an element or an atom of an element.

- For example, the symbol for mercury is Hg.

- Some symbols are derived from the common name of the element; others are abbreviations of the name in another language.

- Most symbols have one or two letters, but three-letter symbols have been used to describe some elements that have atomic numbers greater than 112.

- Only the first letter of a chemical symbol is capitalized.

FIGURE 2.13

The symbol Hg represents the element mercury regardless of the amount; it could represent one atom of mercury or a large amount of mercury.

SOME COMMON ELEMENTS AND THEIR SYMBOLS

Element	Symbol	Element	Symbol
aluminum	Al	iron	Fe (from ferrum)
bromine	Br	lead	Pb (from plumbum)
calcium	Ca	magnesium	Mg
carbon	C	mercury	Hg (from hydrargyrum)
chlorine	Cl	nitrogen	N
chromium	Cr	oxygen	O
cobalt	Co	potassium	K (from kalium)
copper	Cu (from cuprum)	silicon	Si
fluorine	F	silver	Ag (from argentum)
gold	Au (from aurum)	sodium	Na (from natrium)
helium	He	sulfur	S
hydrogen	H	tin	Sn (from stannum)
iodine	I	zinc	Zn

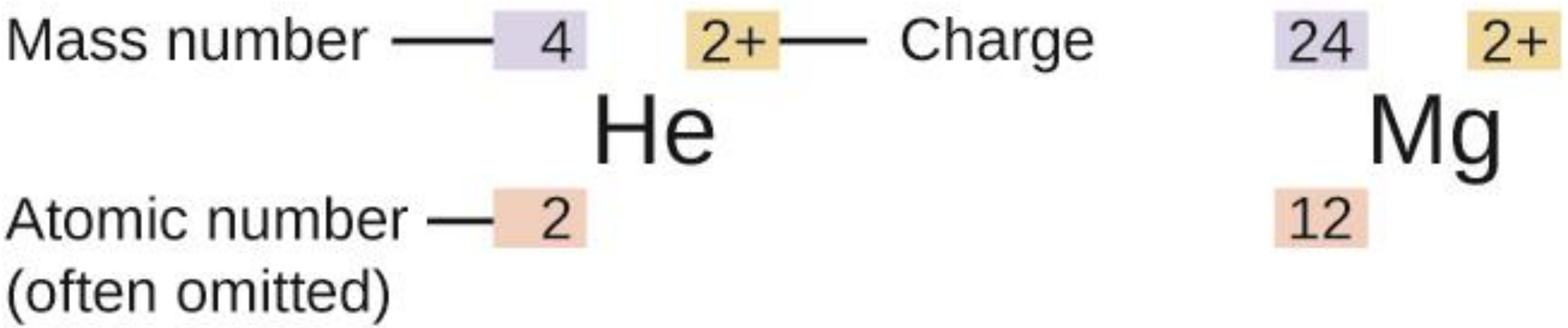
ISOTOPES

- The symbol for a specific isotope of any element is written by placing the mass number as a superscript to the left of the element symbol.

- The atomic number is sometimes written as a subscript to the left of the element symbol.

- For example, magnesium exists as a mixture of three isotopes.
    - ^{24}Mg, ^{25}Mg, and ^{26}Mg
    - All isotopes have 12 protons, but the number of neutrons are different.

FIGURE 2.14



The symbol for an atom indicates the element via its usual two-letter symbol, the mass number as a left superscript, the atomic number as a left subscript (sometimes omitted), and the charge as a right superscript.

ISOTOPES OF HYDROGEN

- Hydrogen exists as a mixture of three isotopes.

Symbol	Atomic Number	Number of Protons	Number of Neutrons	Mass (amu)	% Natural Abundance
${}_{1}^{1}H$ (protium)	1	1	0	1.0078	99.989
${}_{1}^{2}H$ (deuterium)	1	1	1	2.0141	0.0115
${}_{1}^{3}H$ (tritium)	1	1	2	3.01605	trace

ATOMIC MASS

- Each proton and each neutron has a mass of ~ 1 amu.

- Each electron weighs far less.

- Therefore the **atomic mass** of a single atom in amu is ***approx.*** equal to its mass number.

- However, most elements exist naturally as a mixture of two or more isotopes.

- The periodic table lists the weighted, average mass of all the isotopes present in a naturally occurring sample of that element.

ATOMIC MASS

$$average\ mass = \sum_i (fractional\ abundance \times isotopic\ mass)_i$$

- **For example, the element boron is composed of two isotopes:**
  - 19.9% ^{10}B with a mass of 10.0129 amu
  - 80.1% ^{11}B with a mass of 11.0093 amu.

$$boron\ average\ mass$$

$$= (0.199 \times 10.0129\ amu) + (0.801 \times 11.0093\ amu)$$

$$= 10.81\ amu$$

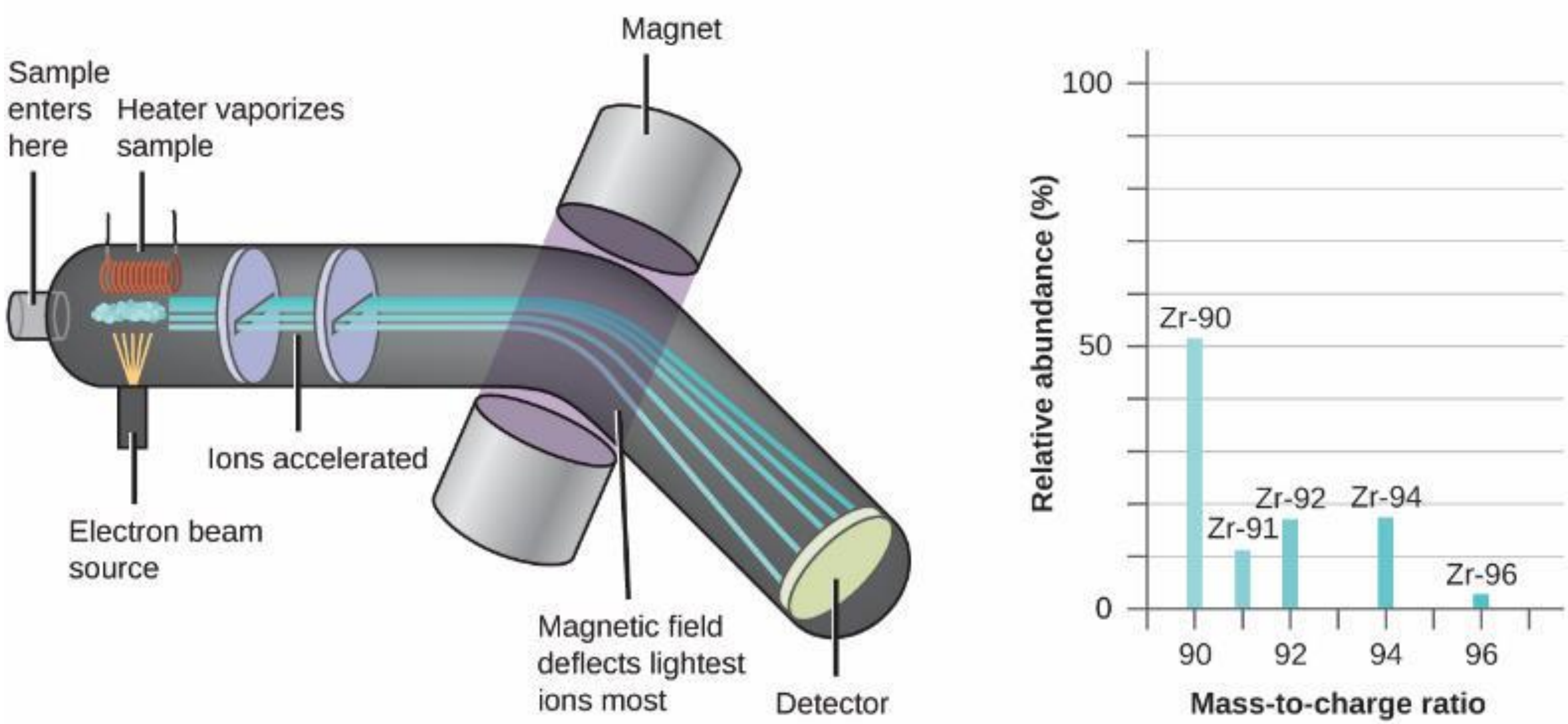
MASS SPECTROMETRY (MS)

- The occurrence and natural abundances of isotopes can be experimentally determined using an instrument called a mass spectrometer.

- The sample is vaporized and exposed to a high-energy electron beam that causes the sample’s atoms (or molecules) to become electrically charged, typically by losing one or more electrons.

- These cations are then separated by their mass and charge.

FIGURE 2.15



Analysis of zirconium in a mass spectrometer produces a mass spectrum with peaks showing the different isotopes of Zr.

2.4 CHEMICAL FORMULAS

- **Molecular Formula –** A representation of a molecule or compound which consists of the following:

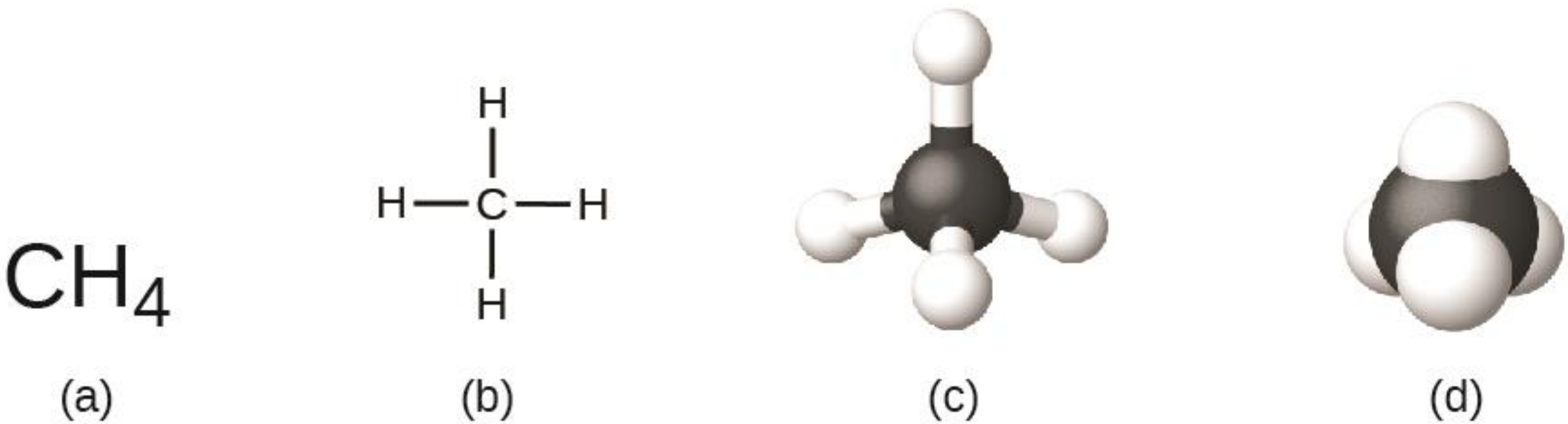
  1) Chemical symbols to indicate the types of atoms.

  2) Subscripts after the symbol to indicate the number of each type of atom in the molecule.

  Subscripts are used only when more than one atom of a given type is present.

- A **Structural Formula** shows the same information as a molecular formula but also shows how the atoms are connected.

FIGURE 2.16

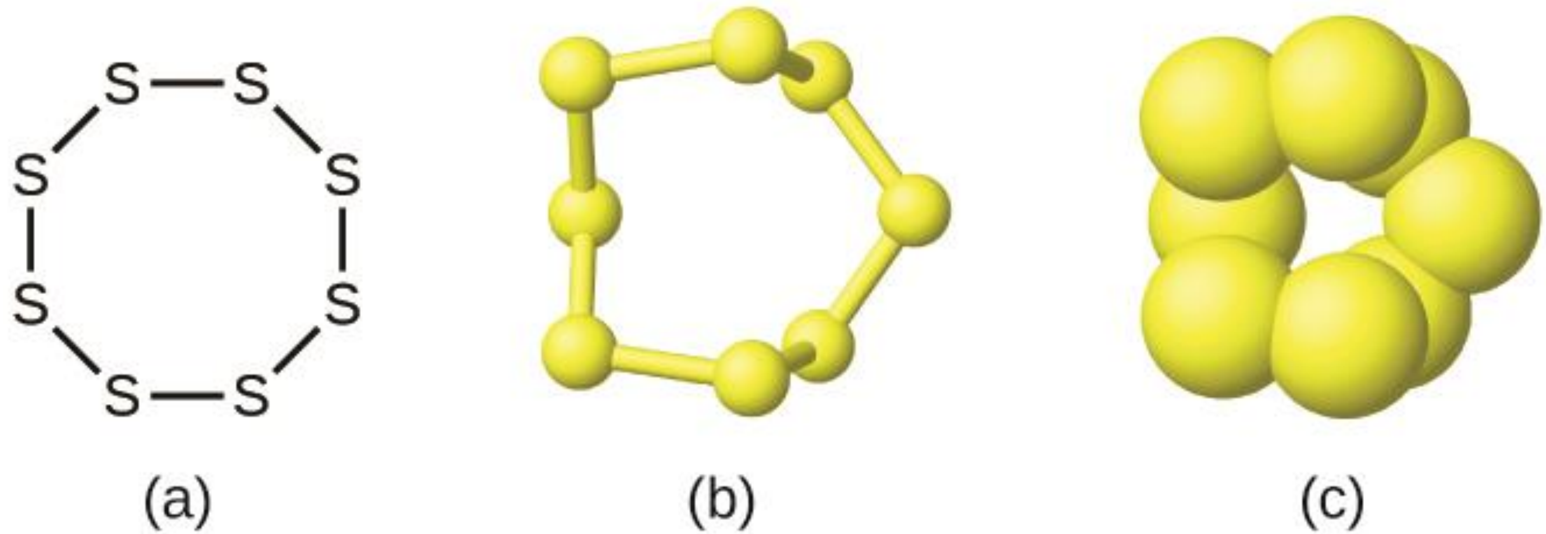


A methane molecule can be represented as (a) a molecular formula, (b) a structural formula, (c) a ball-and-stick model, and (d) a space-filling model. Carbon and hydrogen atoms are represented by black and white spheres, respectively.

ELEMENTS THAT EXIST AS MOLECULES

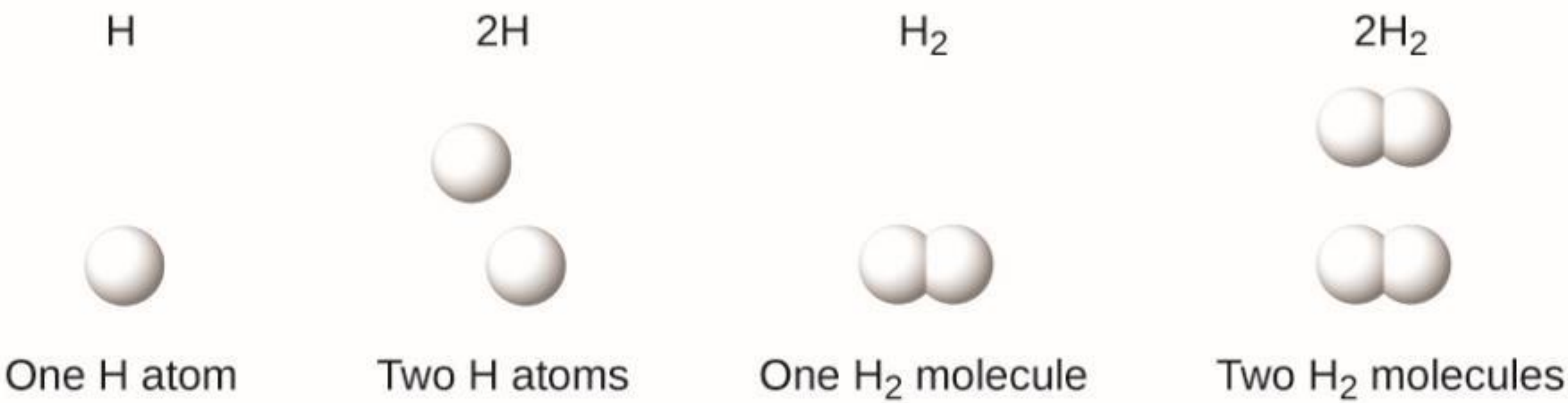
- Many elements consist of discrete, individual atoms.
- Some elements exist as molecules.
- **Diatomic molecules:** H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2
- The most common form of elemental sulfur exists as S_8.

FIGURE 2.17



A molecule of sulfur is composed of eight sulfur atoms and is therefore written as S_8. It can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. Sulfur atoms are represented by yellow spheres.

FIGURE 2.18



The symbols H, 2H, H_2, and $2H_2$ represent very different entities.

EMPIRICAL FORMULA

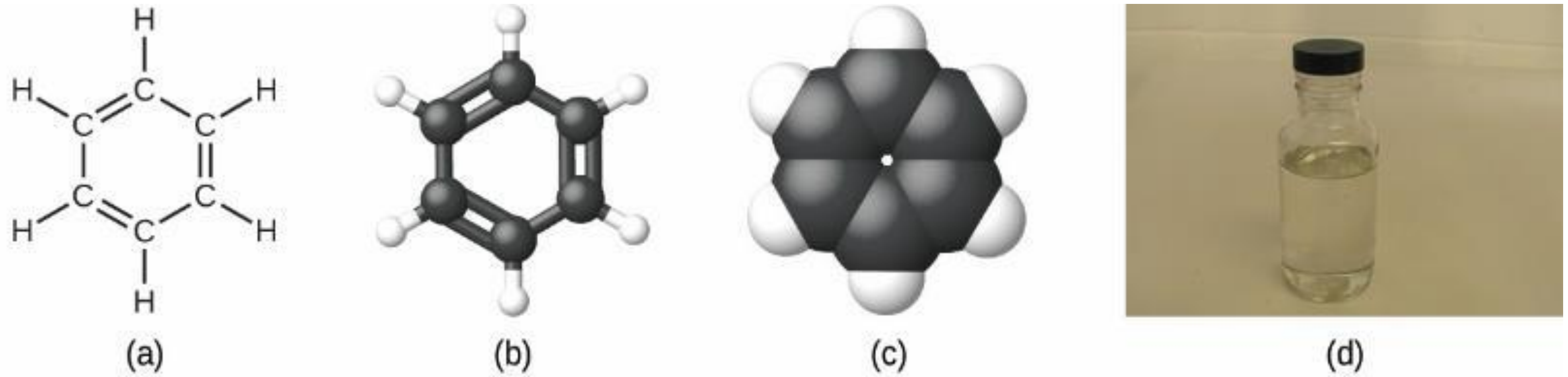
- An ***empirical formula*** indicates *the simplest whole-number ratio of the number of atoms (or ions) in the compound.*

- A ***molecular formula*** *indicates the actual numbers of atoms of each element in a molecule of the compound.*

- ***Example:*** Benzene
  - Molecular formula = C_6H_6 Empirical formula = CH

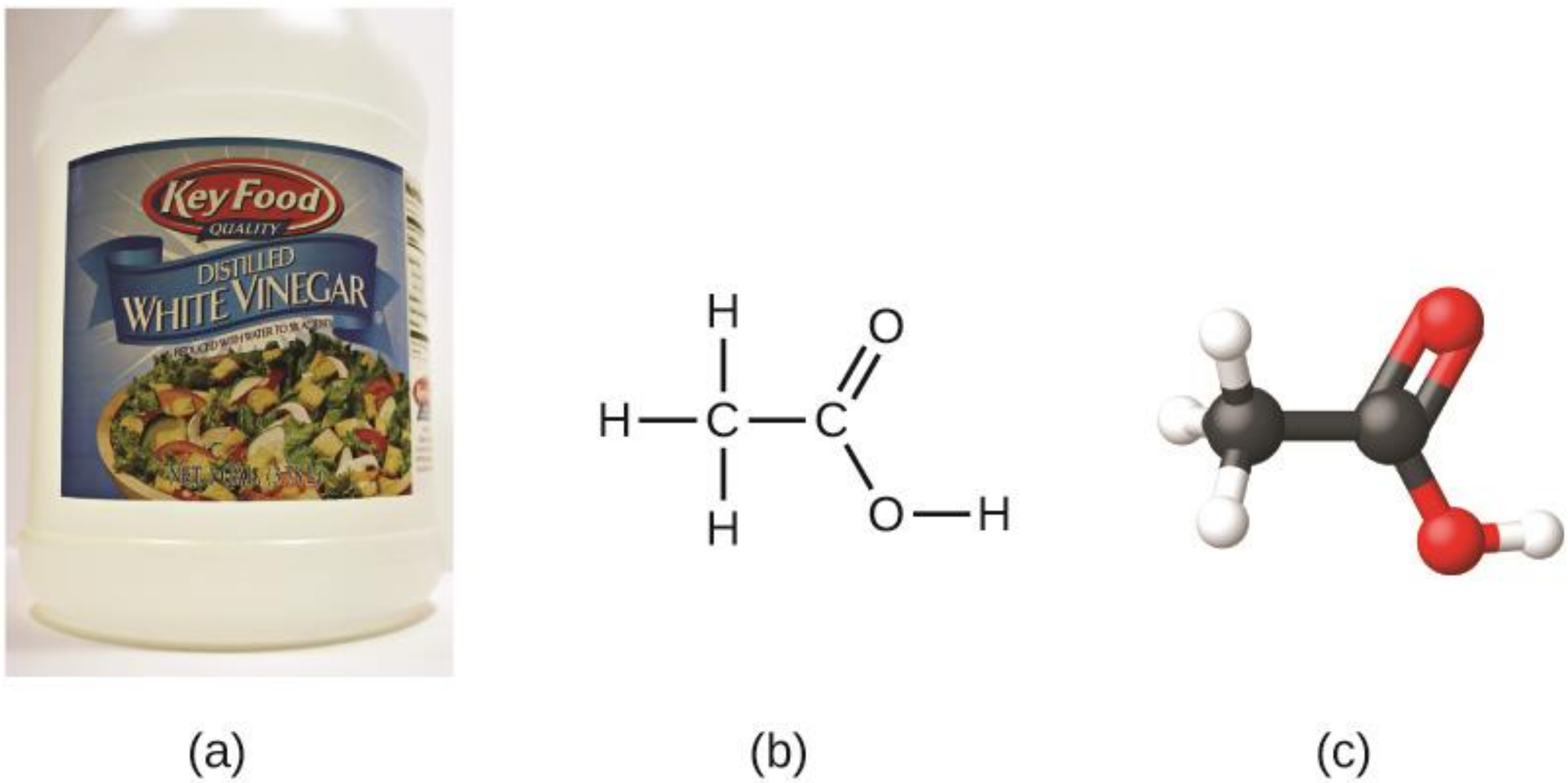
- ***Example:*** Acetic acid
  - Molecular formula = $C_2H_4O_2$ Empirical formula = CH_2O

FIGURE 2.20



Benzene, C_6H_6, is produced during oil refining and has many industrial uses. A benzene molecule can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. (d) Benzene is a clear liquid. (credit d: modification of work by Sahar Atwa)

FIGURE 2.21



(a) Vinegar contains acetic acid, $C_2H_4O_2$, which has an empirical formula of CH_2O. It can be represented as (b) a structural formula and (c) as a ball-and-stick model. (credit a: modification of work by “HomeSpot HQ”/Flickr)

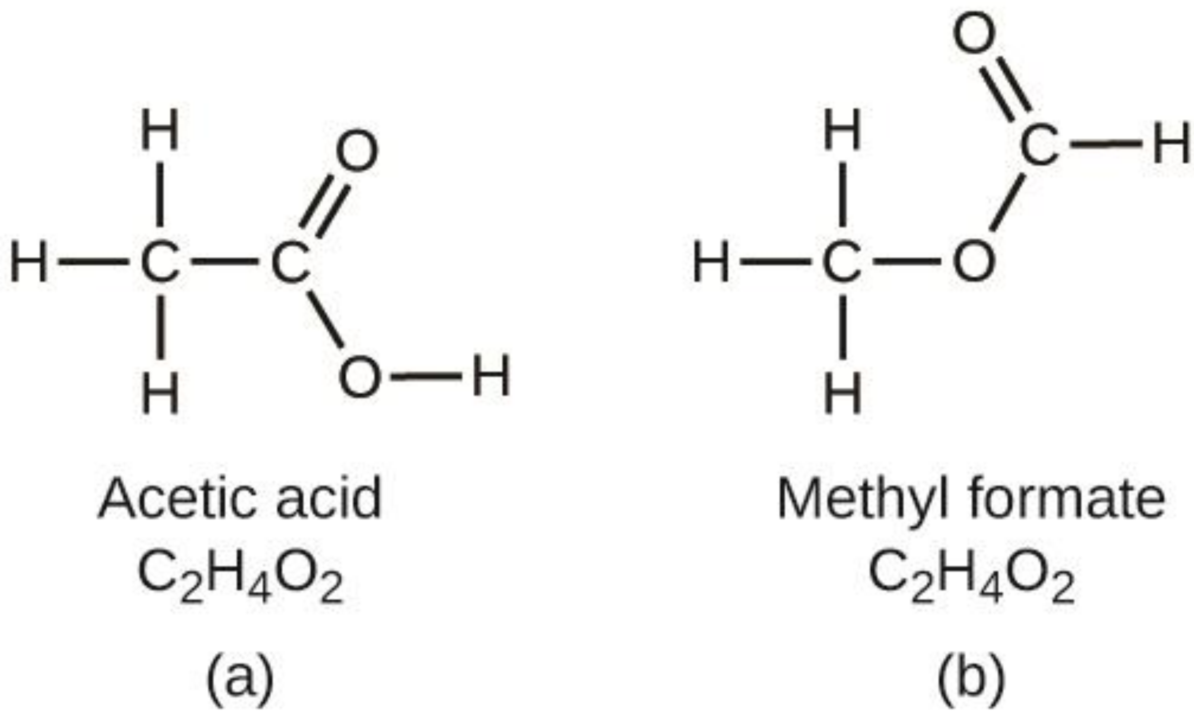
ISOMERS

- It may be possible for the same atoms to be arranged in different ways.

- **Isomers** — compounds with the same chemical formula but different molecular structures.

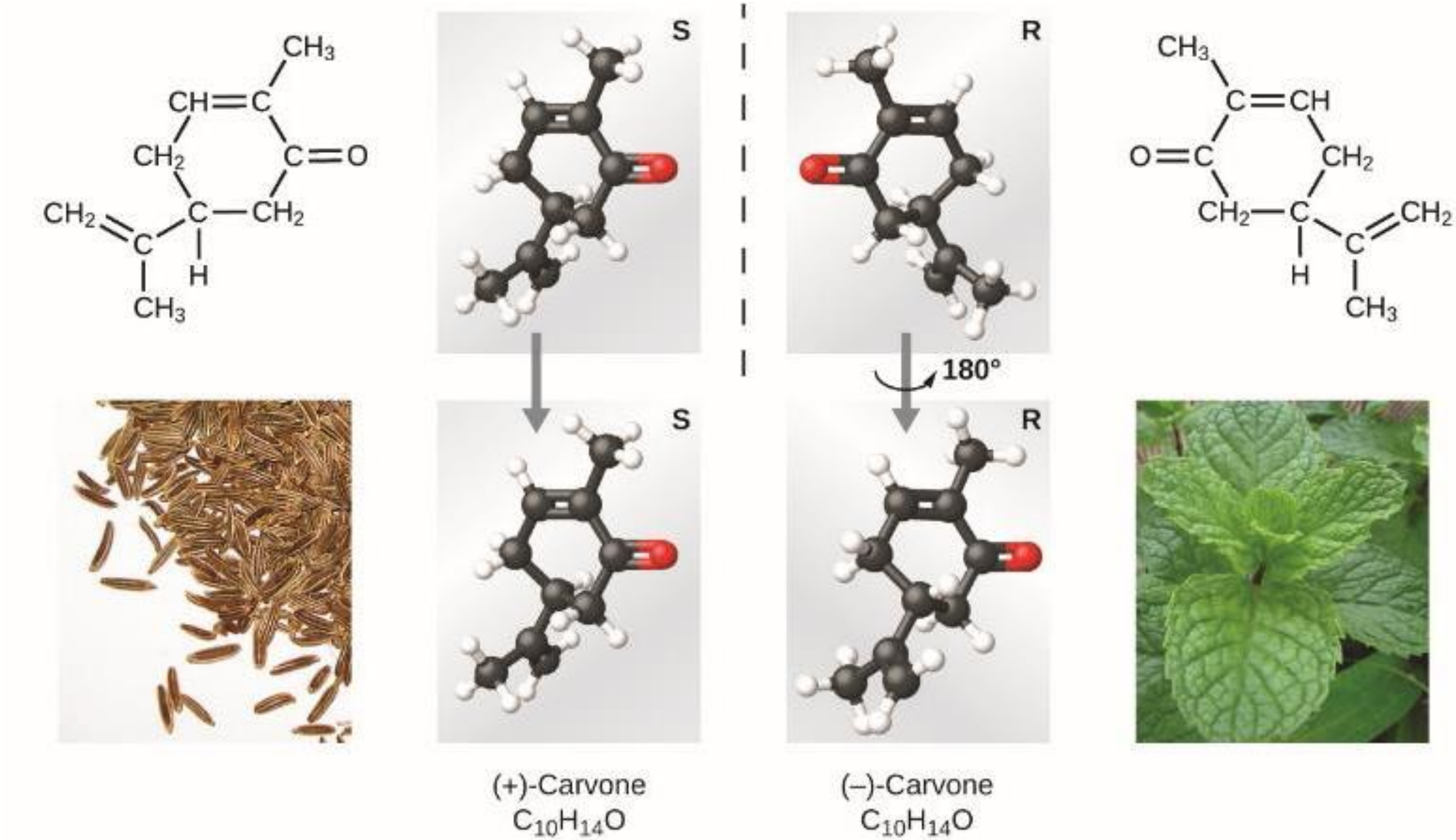
- Example: acetic acid and methyl formate both have the molecular formula $C_2H_4O_2$, but they have different structures and properties.

FIGURE 2.23



Molecules of (a) acetic acid and methyl formate (b) are structural isomers; they have the same formula ($C_2H_4O_2$) but different structures (and therefore different chemical properties).

FIGURE 2.24



Molecules of carvone are spatial isomers; they only differ in the relative orientations of the atoms in space. (credit bottom left: modification of work by “Miansari66”/Wikimedia Commons; credit bottom right: modification of work by Forest & Kim Starr)

2.5 THE PERIODIC TABLE

- **Dimitri Mendeleev in Russia (1869) and Lothar Meyer in Germany (1870)** independently recognized that there was a periodic relationship among the properties of the elements known at that time.

- **For Example:**
  - Lithium (Li), sodium (Na), and potassium (K) are all shiny, conduct heat and electricity well, and have similar chemical properties.

  - Calcium (Ca), strontium (Sr), and barium (Ba) are also shiny, conduct heat and electricity well, but are less reactive than Li, Na, and K.

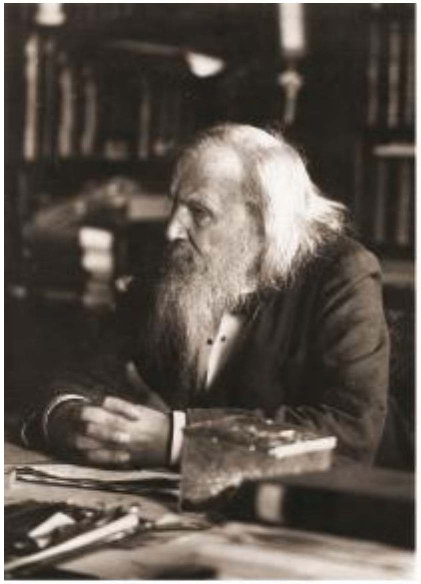
THE FIRST PERIODIC TABLE

• Both **Mendeleev and Lothar Meyer** published tables with the elements arranged according to increasing atomic mass.

• Mendeleev used his table to predict the existence of elements that would have the properties similar to aluminum and silicon, but were not yet known.

• The discoveries of gallium (1875) and germanium (1886) provided great support for Mendeleev's work.

FIGURE 2.25



(a)

Reihen	Gruppo I. — R^2O	Gruppo II. — RO	Gruppo III. — R^2O^3	Gruppo IV. RH^4 RO^2	Gruppo V. RH^3 R^2O^5	Gruppo VI. RH^2 RO^3	Gruppo VII. RH R^2O^7	Gruppo VIII. — RO^4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

(b)

(a) Dimitri Mendeleev is widely credited with creating (b) the first periodic table of the elements. (credit a: modification of work by Serge Lachinov; credit b: modification of work by “Den fjättrade ankan”/Wikimedia Commons)

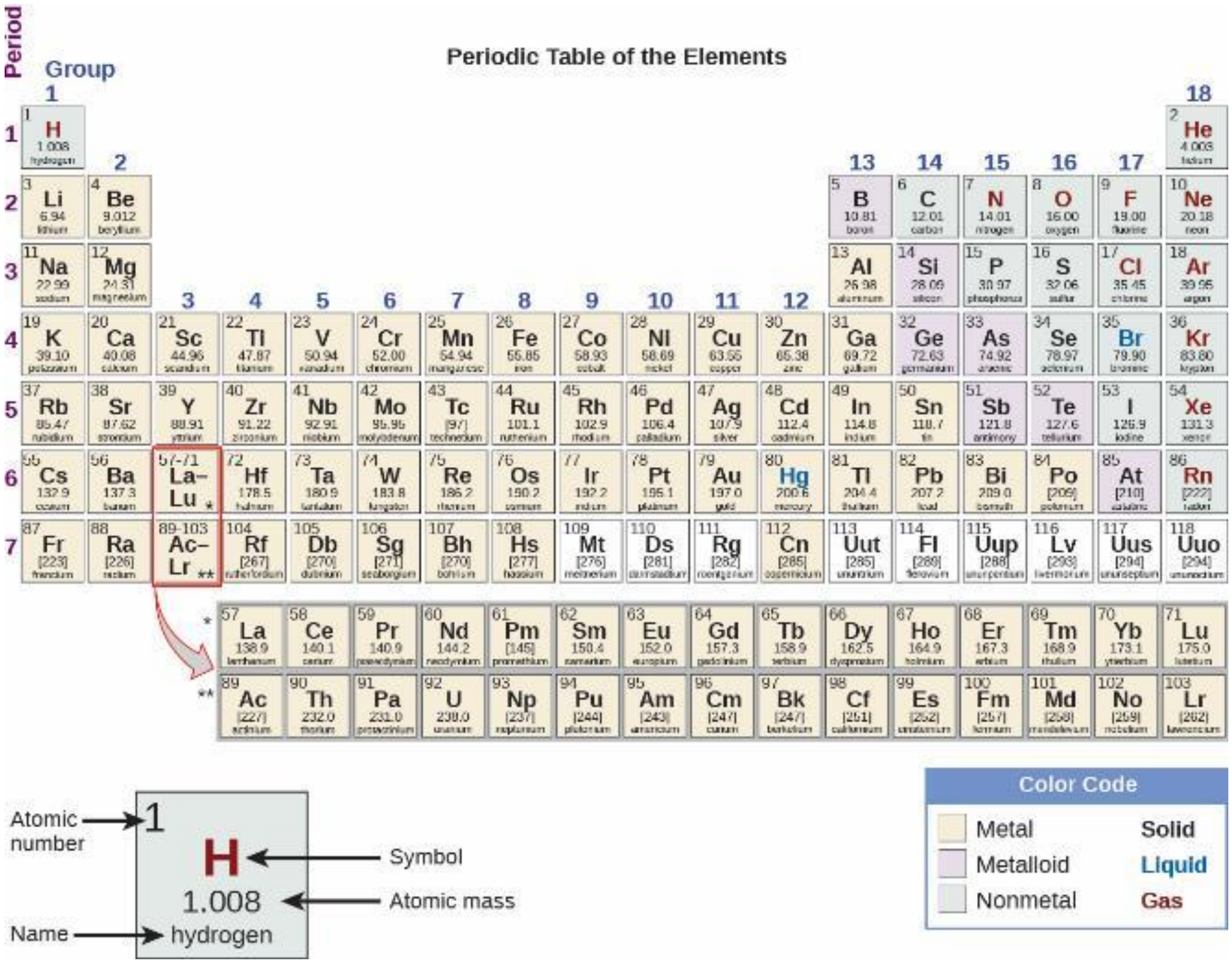
THE MODERN PERIODIC TABLE

- By the twentieth century, it became apparent that the periodic relationship involved atomic numbers rather than atomic masses.

- **Periodic Law –** The properties of the elements are periodic functions of their atomic numbers.

- A modern periodic table arranges the elements in increasing order of their atomic numbers and groups atoms with similar properties in the same vertical column
  - **Periods or series** – horizontal rows
  - **Groups –** vertical columns (numbered 1-18)

FIGURE 2.26



Elements in the periodic table are organized according to their properties.

CLASSIFICATIONS OF ELEMENTS

- **Metals** are shiny, malleable, good conductors of heat and electricity.

- **Nonmetals** appear dull, poor conductors of heat and electricity.

- **Metalloids** conduct heat and electricity moderately well, and possess some properties of metals and some properties of nonmetals.

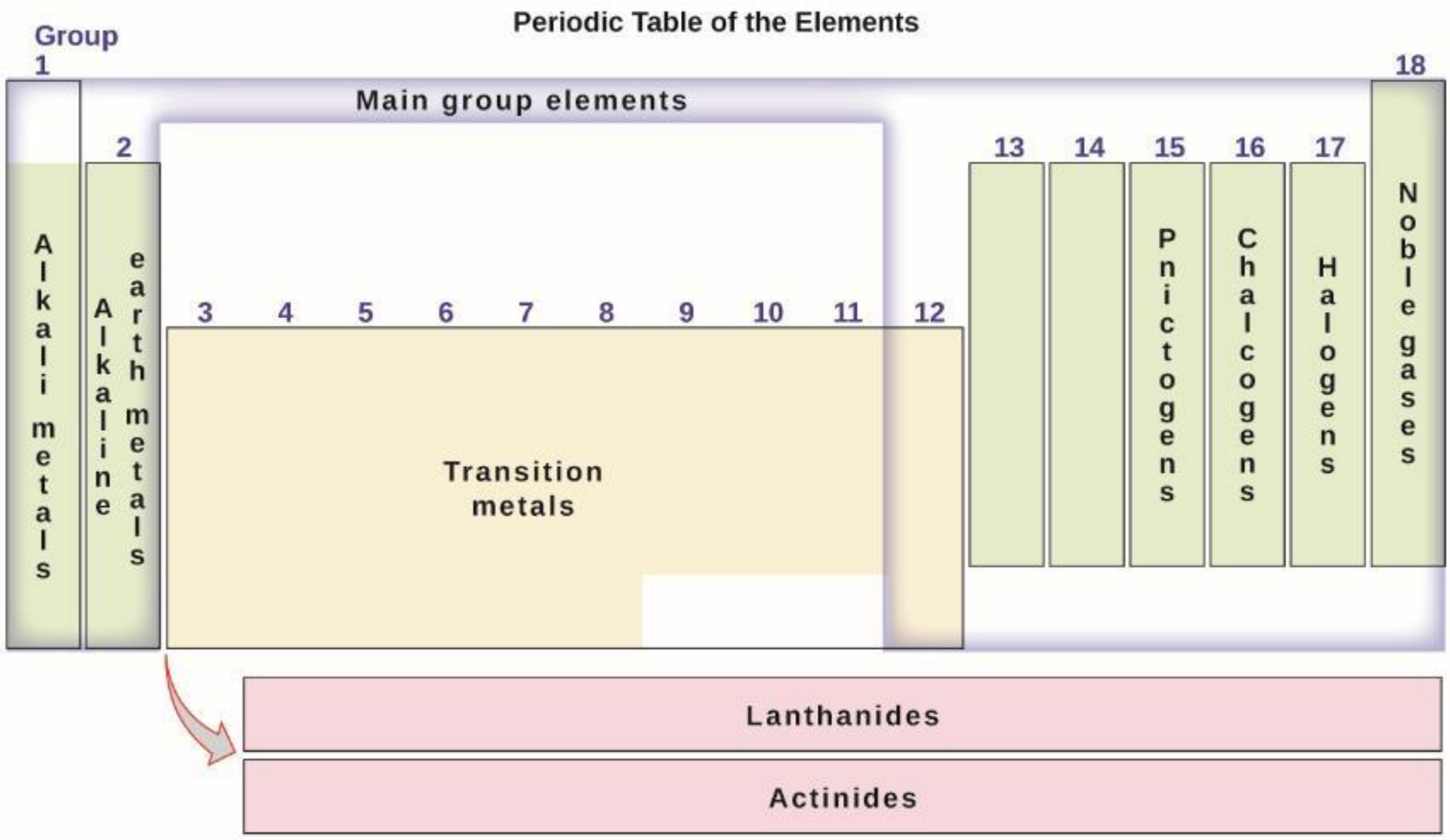
CLASSIFICATIONS OF ELEMENTS

- **Main group elements (or representative elements)**
  - Groups: 1, 2, 13-18
- **Transition metals**
  - Groups: 3-12
- **Inner transition metals**
  - Two rows at the bottom of the periodic table.
    - Lanthanides – top row
    - Actinides – bottom row

CLASSIFICATIONS OF ELEMENTS

- Alkali metals – Group 1 (except hydrogen)
- Alkaline earth metals – Group 2
- Pnictogens – Group 15
- Chalcogens – Group 16
- Halogens – Group 17
- Noble Gases (or inert gases) – Group 18

FIGURE 2.27



The periodic table organizes elements with similar properties into groups.

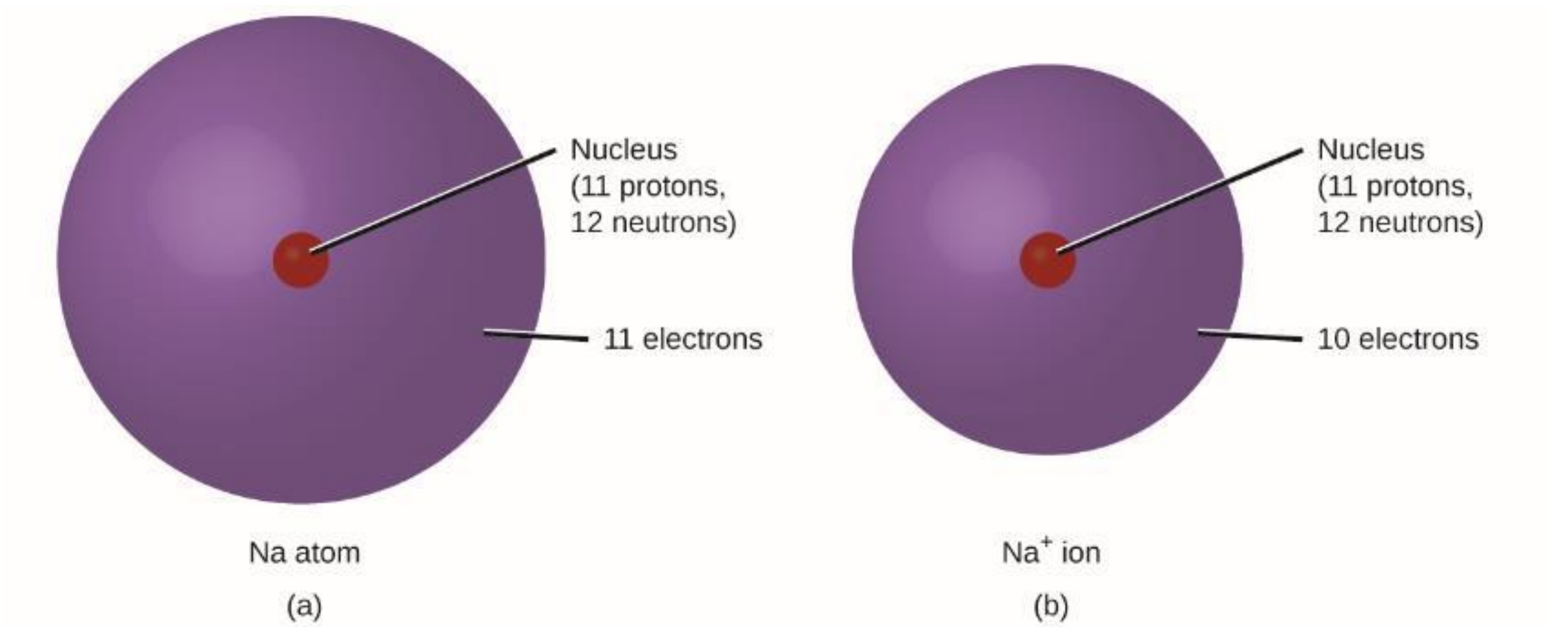
2.6 MOLECULAR AND IONIC COMPOUNDS

- In ordinary chemical reactions, the nucleus of each atom (and thus the identity of the element) remains unchanged.

- Electrons participate in chemical reactions by being
  - gained
  - lost
  - shared

- The gain or lose of electrons, results in the formation of ions.

FIGURE 2.28



(a) A sodium atom (Na) has equal numbers of protons and electrons (11) and is uncharged. (b) A sodium cation (Na^+) has lost an electron, so it has one more proton (11) than electrons (10), giving it an overall positive charge, signified by a superscripted plus sign.

PREDICTING ION CHARGE

- The periodic table can serve as a guide for predicting the ionic charge of main-group elements.

- Many main-group ***metals lose enough electrons*** to leave them with the same number of electrons as an atom of the preceding noble gas.
  - Group 1 – lose one electron, form a cation with a 1+ charge.
  - Group 2 – lose two electrons, form a cation with a 2+ charge.

PREDICTING ION CHARGE

- Many ***non-metals gain enough electrons*** to give them the same number of electrons as an atom of the next noble gas.
  - Group 17 – gain one electron, form an anion with a 1- charge.
  - Group 16 – gain two electrons, form an anion with a 2- charge.

PREDICTING ION CHARGE

- **Example: Ca (group 2)**
  - Ca atom (20 protons, 20 electrons)
  - Loses 2 electrons
  - Now a Ca^{2+} ion (20 protons, 18 electrons)
  - Same number of electrons as the preceding noble gas, Ar.

- **Example: Br (group 17)**
  - Br atom (35 protons, 35 electrons)
  - Gains 1 electron
  - Now a Br^{-} ion (35 protons, 36 electrons)
  - Same number of electrons as the next noble gas, Kr.

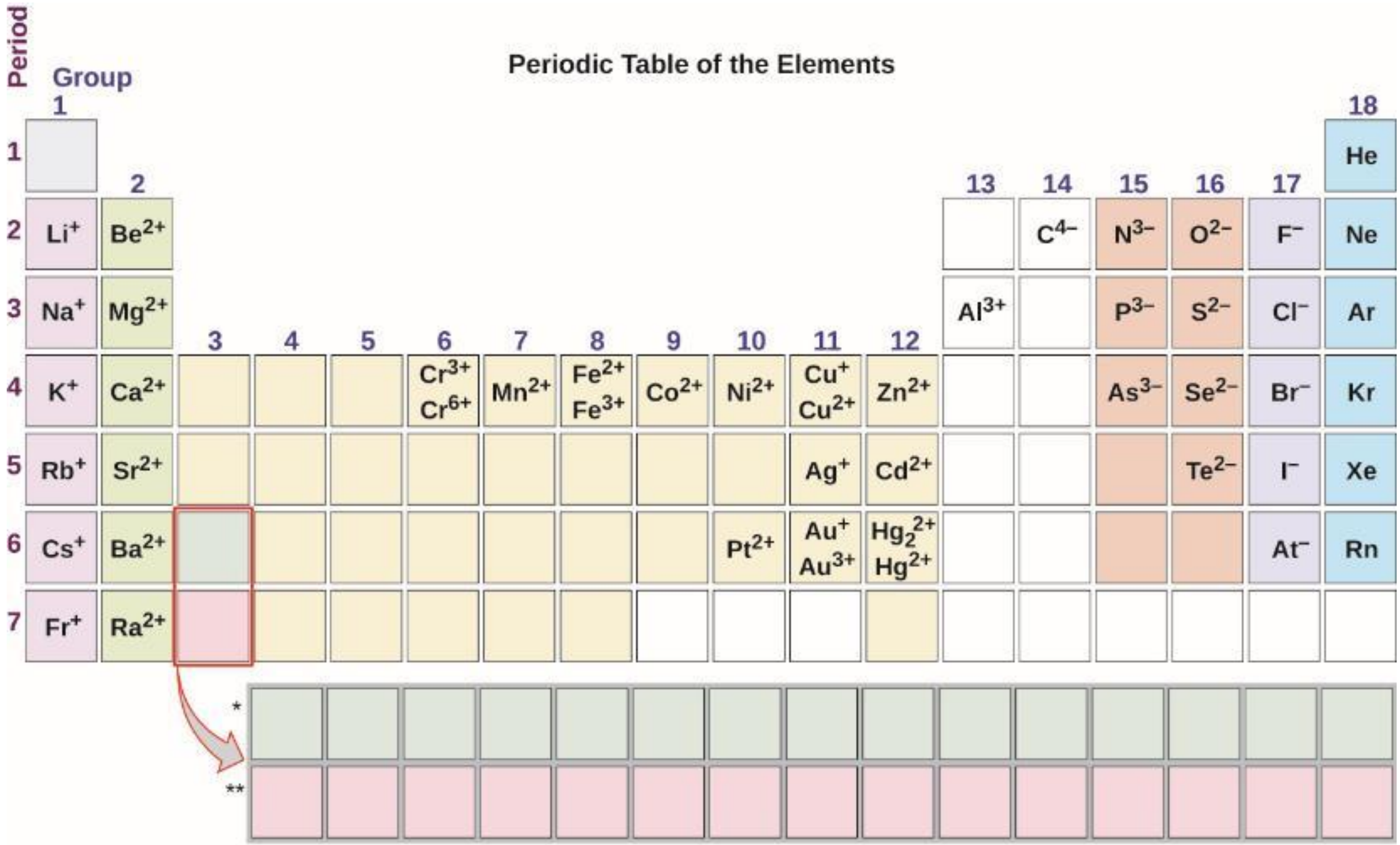
PREDICTING ION CHARGE

- **Moving from far left to far right in the periodic table:**
  - Positive charges of cations are equal to the group number.

- **Moving from the far right to the far left in the periodic table:**
  - Negative charges of anions are equal to the number of groups moved left from the noble gas.

- **This method is less reliable for transition metals.**
  - Cu forms ions of 1+ and 2+ charge.
  - Fe forms ions of 2+ and 3+ charge.

FIGURE 2.29



Some elements exhibit a regular pattern of ionic charge when they form ions.

POLYATOMIC IONS

- The ions that we have discussed so far are called ***monatomic ions***, that is, they are ions formed from only one atom.

- ***Polyatomic ions*** are electrically charged molecules (a group of bonded atoms with an overall charge).

- ***Oxyanions*** are polyatomic ions that contain one or more oxygen atoms.

COMMON POLYATOMIC IONS

Name	Formula	Name	Formula
ammonium	NH_4^+	nitrate	NO_3^-
hydronium	H_3O^+	nitrite	NO_2^-
peroxide	O_2^{2-}	sulfate	SO_4^{2-}
hydroxide	OH^-	hydrogen sulfate	HSO_4^-
acetate	CH_3COO^-	sulfite	SO_3^{2-}
cyanide	CN^-	hydrogen sulfite	HSO_3^-
azide	N_3^-	phosphate	PO_4^{3-}
carbonate	CO_3^{2-}	hydrogen phosphate	HPO_4^{2-}
bicarbonate	HCO_3^-		

COMMON POLYATOMIC IONS

Name	Formula
dihydrogen phosphate	$H_2PO_4^-$
perchlorate	ClO_4^-
chlorate	ClO_3^-
chlorite	ClO_2^-
hypochlorite	ClO^-
chromate	CrO_4^{2-}
dichromate	$Cr_2O_7^{2-}$
permanganate	MnO_4^-

NAMING OXYANIONS

- There is a system for naming oxyanions.

- **When a nonmetal forms two oxyanions**
  - ***-ate*** is the suffix used for the ion with the larger number of oxygen atoms.
  - ***-ite*** is the suffix used for the ion with the smaller number of oxygen atoms.

- **When a nonmetal forms more than two oxyanions, prefixes are used in addition to *-ate* and *-ite***
  - ***per-*** (largest number of oxygens)
  - ***hypo-*** (smallest number of oxygens)

TYPES OF CHEMICAL BONDS

- When electrons are ***transferred***, ions form, and an ***ionic bond*** results.

- ***Ionic bonds*** are electrostatic forces of attraction.

- When electrons are ***shared*** and molecules form, a ***covalent bond*** results.

- Compounds are classified as ionic or molecular (covalent) on the basis of the bonds present in them.

IONIC COMPOUNDS

- Metals readily lose electrons – form cations.
- Nonmetals readily gain electrons – form anions.
- When a metal and nonmetal react, a transfer of electrons usually takes place.
- Metals and nonmetals generally form ionic compounds.
- A compound that contains ions and is held together by ionic bonds is called an ***ionic compound.***

IONIC COMPOUND EXAMPLES

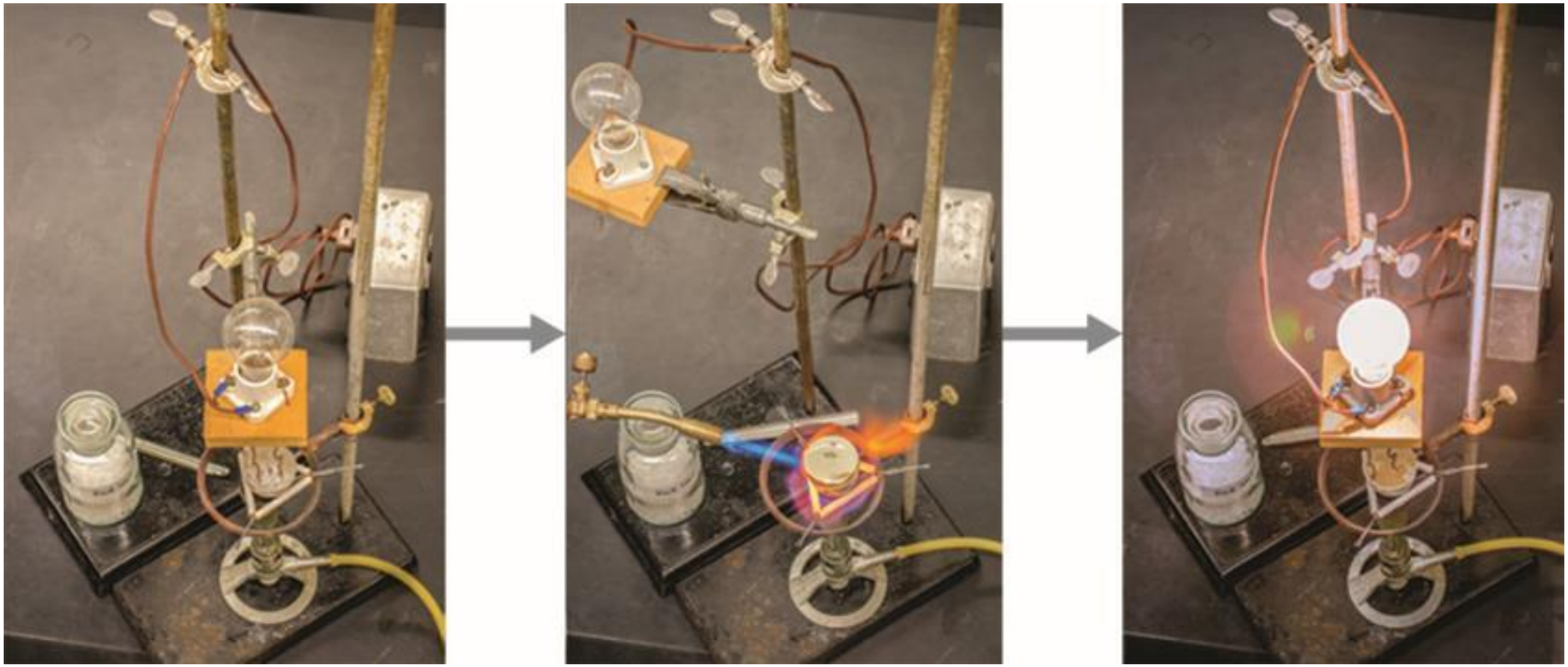
- Na and Cl
  - One Na atom gives up one electron forming a Na^+ ion.
  - One Cl atom accepts that electron forming a Cl^- ion.
  - The ionic compound, NaCl forms.

- Ca and Cl
  - One Ca atom gives up two electrons forming a Ca^{2+} ion.
  - Two Cl atoms each accept one electron forming two Cl^- ions.
  - The ionic compound, $CaCl_2$ forms.

PROPERTIES OF IONIC COMPOUNDS

- Typically solids with high melting and boiling points.
- Non-conductive in solid form.
- Conductive in molten form.

FIGURE 2.30



Sodium chloride melts at 801 °C and conducts electricity when molten. (credit: modification of work by Mark Blaser and Matt Evans)

FORMULAS OF IONIC COMPOUNDS

- Ionic compounds are electrically neutral overall.

- The formula of an ionic compound must have a ratio of ions such that the numbers of positive and negative charges are equal.

- These formulas are not molecular formulas.

- **Example:** Al^{3+} and O^{2-} forms Al_2O_3
  - Two Al^{3+} ions gives six positive charges.
  - Three O^{2-} ions gives six negative charges.

FIGURE 2.31

Although pure aluminum oxide is colorless, trace amounts of iron and titanium give blue sapphire its characteristic color. (credit: modification of work by Stanislav Doronenko)

FORMULAS OF IONIC COMPOUNDS

- Many ionic compounds contain polyatomic ions as the cation, the anion, or both.
- Treat polyatomic ions as discrete units.
- Parentheses in a formula are used to indicate a group of atoms that behave as a unit.
- **Example:** Ca^{2+} and PO_4^{3-} forms $Ca_3(PO_4)_2$
  - Three Ca^{2+} ions gives six positive charges.
  - Two PO_4^{3-} ions gives six negative charges.

MOLECULAR COMPOUNDS

- Molecular compounds (covalent compounds) result when atoms share electrons.
- Exist as discrete, neutral molecules.
- Usually formed by a combination of nonmetals.
- Often exist as gases, low-boiling liquids, and low-melting solids.

2.7 CHEMICAL NOMENCLATURE

- **Nomenclature –** A collection of rules for naming things.

- Compounds are identified by both their formula and name.

- **We will learn how to name the following types of inorganic compounds:**
  - Ionic and molecular **binary compounds –** composed of two elements.
  - Ionic compounds containing polyatomic ions.
  - Acids

NAMING IONIC COMPOUNDS

- Name the cation first, followed by the name of the anion.
- A monoatomic cation is just given the name of the element.
- A monoatomic anion is given the name of the element with its ending replaced by the suffix –*ide*.
- A polyatomic ion is just given the name of the ion.

NAMES OF SOME IONIC COMPOUNDS CONTAINING ONLY MONOATOMIC IONS

NaCl, sodium chloride	Na_2O, sodium oxide
KBr, potassium bromide	CdS, cadmium sulfide
CaI_2, calcium iodide	Mg_3N_2, magnesium nitride
CsF, cesium fluoride	Ca_3P_2, calcium phosphide
LiCl, lithium chloride	Al_4C_3, aluminum carbide

NAMES OF SOME IONIC COMPOUNDS CONTAINING POLYATOMIC IONS

$KC_2H_3O_2$, potassium acetate

NH_4Cl, ammonium chloride

$NaHCO_3$, sodium bicarbonate

$CaSO_4$, calcium sulfate

$Al_2(CO_3)_3$, aluminum carbonate

$Mg_3(PO_4)_2$, magnesium phosphate

NAMING IONIC COMPOUNDS CONTAINING A METAL ION WITH A VARIABLE CHARGE

- Most of the transition metals can form two or more cations with different charges.

- The charge of the metal ion is specified by a Roman numeral in parentheses after the name of the metal.

Transition Metal Ionic Compound	**Name**
$FeCl_3$	iron(III) chloride
Hg_2O	mercury(I) oxide
HgO	mercury(II) oxide
$Cu_3(PO_4)_2$	copper(II) phosphate

NAMING BINARY MOLECULAR (COVALENT) COMPOUNDS

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- Molecular compounds are name using a different set of rules.
- Covalent bonding allows for significant variation in the ratios of the atoms in a molecule.
- The names for molecular compounds must explicitly identify these ratios.

NAMING BINARY MOLECULAR (COVALENT) COMPOUNDS

- The name of the more metallic element (the one farther to the left and/or bottom of the periodic table) is named first.

- Followed by the name of the more nonmetallic element (the one farther to the right and/or top) with its ending changed to the suffix *–ide*.

- The numbers of atoms of each element are designated by Greek prefixes.

NOMENCLATURE PREFIXES

Number	Prefix
1 (sometimes omitted)	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

NOMENCLATURE PREFIXES

- When only one atom of the first element is present, the prefix *mono*- is usually not used.

- When two vowels are adjacent, the *a* in the Greek prefix is usually dropped.

NAMES OF SOME BINARY MOLECULAR COMPOUNDS

Compound	Name
SO_2	sulfur dioxide
BCl_3	boron trichloride
SO_3	sulfur trioxide
SF_6	sulfur hexafluoride
NO_2	nitrogen dioxide
PF_5	phosphorus pentafluoride
N_2O_4	dinitrogen tetroxide
P_4O_{10}	tetraphosphorus decaoxide
N_2O_5	dinitrogen pentoxide
IF_7	iodine heptafluoride

NAMING ACIDS

- Some compounds containing hydrogen are members of an important class of substances known as acids.
- Many acids release hydrogen ions, H^+, when dissolved in water.
- A mixture of an acid with water is given a special name to denote this property.

NAMING BINARY ACIDS

1) The word “hydrogen” is changed to the prefix *hydro-*

2) The other nonmetallic element name is modified by adding the suffix *–ic*

3) The word “acid” is added as a second word

Names of Some Simple Acids

Name of Gas	**Name of Acid**
HF(g), hydrogen fluoride	HF(aq), hydrofluoric acid
HCl(g), hydrogen chloride	HCl(aq), hydrochloric acid
HBr(g), hydrogen bromide	HBr(aq), hydrobromic acid
HI(g), hydrogen iodide	HI(aq), hydroiodic acid
H_2S(g), hydrogen sulfide	H_2S(aq), hydrosulfuric acid

NAMING OXYACIDS

- **Oxyacids** - compounds that contain hydrogen, oxygen, and at least one other element, and are bonded in such a way as to impart acidic properties to the compound.

- Typical oxyacids consist of hydrogen combined with a polyatomic, oxygen-containing ion.

- **To name oxyacids:**

1) Omit “hydrogen”

2) Start with the root name of the anion

3) Replace –*ate* with –*ic*, or –*ite* with –*ous*

4) Add “acid”

NAMING OXYACIDS

Names of Common Oxyacids

Formula	**Anion Name**	**Acid Name**
$HC_2H_3O_2$	acetate	acetic acid
HNO_3	nitrate	nitric acid
HNO_2	nitrite	nitrous acid
$HClO_4$	perchlorate	perchloric acid
H_2CO_3	carbonate	carbonic acid
H_2SO_4	sulfate	sulfuric acid
H_2SO_3	sulfite	sulfurous acid
H_3PO_4	phosphate	phosphoric acid