Chemistry, Raymond Chang
10th edition, 2010
McGraw-Hill

Chapter 2

Atoms, Molecules and Ions

Ahmad Aqel Ifseisi
Assistant Professor of Analytical Chemistry
College of Science, Department of Chemistry
King Saud University
P.O. Box 2455 Riyadh 11451 Saudi Arabia
Building: 05, Office: 2A/149 & AA/53
Tel. 014674198, Fax: 014675992
Web site: http://fac.ksu.edu.sa/aifseisi
E-mail: ahmad3qel@yahoo.com
aifseisi@ksu.edu.sa

كرسي أبحاث المواد المتقدمة
Advanced Materials Research Chair
AMRC

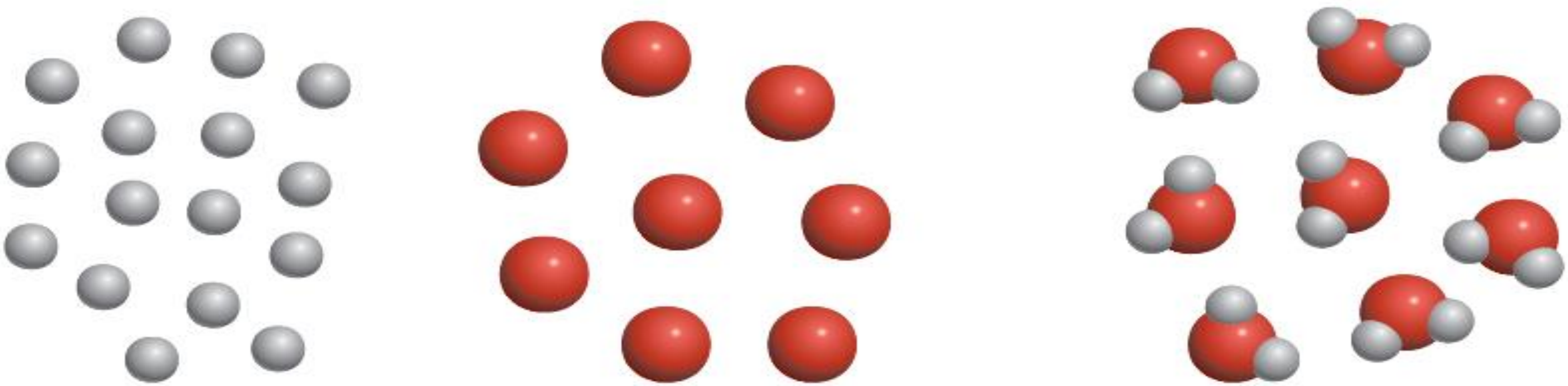
جامعة الملك سعود
King Saud University

2.2
The Structure of the Atom

The Atomic Theory

The modern version of atomic theory was laid by **John Dalton** in 1808, who postulated that elements are composed of extremely small particles, called atoms. The hypotheses about the nature of matter on which Dalton's atomic theory is based can be summarized as follows:

1. Elements are composed of extremely small particles called atoms.
2. All atoms of a given element are identical, having the same size, mass and chemical properties. The atoms of one element are different from the atoms of all other elements.
3. Compounds are composed of atoms of more than one element. To form a certain compound, atoms of the right kinds of elements and specific numbers are needed.
4. A chemical reaction involves only the separation, combination or rearrangement of atoms; it does not result in their creation or destruction (law of conservation of mass, matter can be neither created nor destroyed).



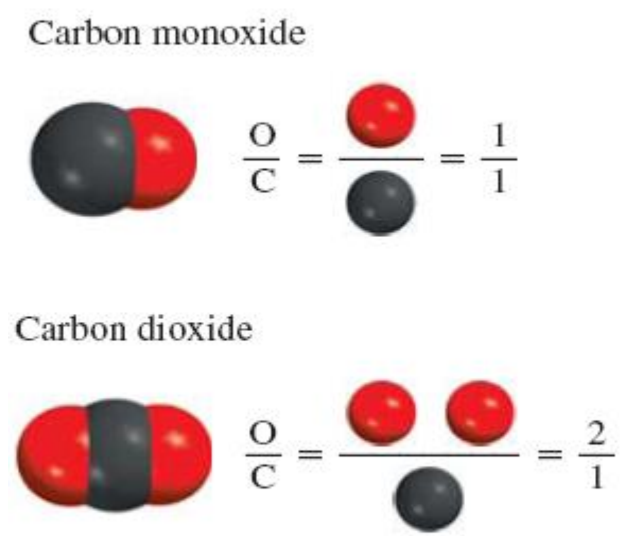
Atoms of element X   Atoms of element Y   Compounds of elements X and Y

The third hypothesis is an extension of a law published in 1799 by **Joseph Proust**. Proust's **law of definite proportions** states that different samples of the same compound always contain its constituent elements in the same proportion by mass. Thus, if we were to analyze samples of carbon dioxide gas obtained from different sources, we would find in each sample the same ratio by mass of carbon to oxygen.

Dalton's third hypothesis supports another important law, the **law of multiple proportions**. According to the law, if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.

Different compounds made up of the same elements differ in the number of atoms of each kind that combine.

e.g., carbon forms two stable compounds with oxygen, namely, carbon monoxide (CO) and carbon dioxide (CO_2). Modern measurement techniques indicate that one atom of carbon combines with one atom of oxygen in CO and with two atoms of oxygen in CO_2. Thus, the ratio of oxygen in CO to oxygen in CO_2 is 1:2.

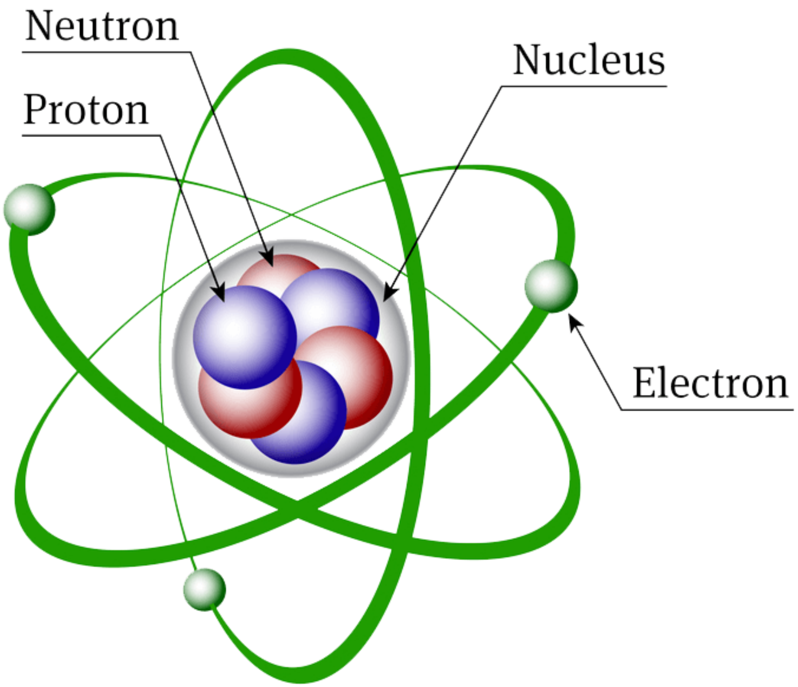


Ratio of oxygen in carbon monoxide to oxygen in carbon dioxide: 1:2

The Structure of the Atom

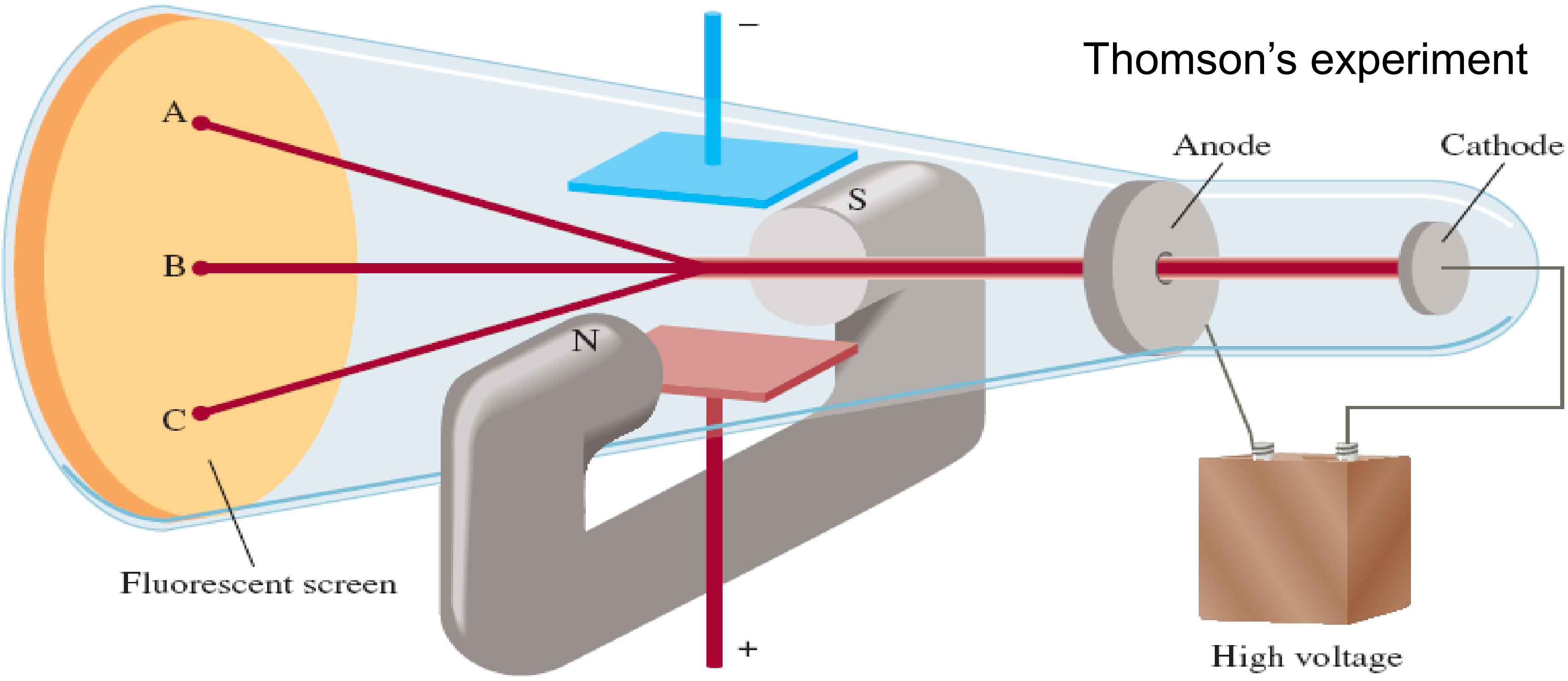
The **atom** is the basic unit of an element that can enter into chemical combination.

Dalton imagined an atom that was both extremely **small** and **indivisible**. However, a series of investigations that began in the 1850s and extended into the twentieth century clearly demonstrated that atoms actually possess internal structure; that is, they are made up of even smaller particles, which are called **subatomic particles**. This research led to the discovery of three such particles; **electrons**, **protons** and **neutrons**.



The Electron

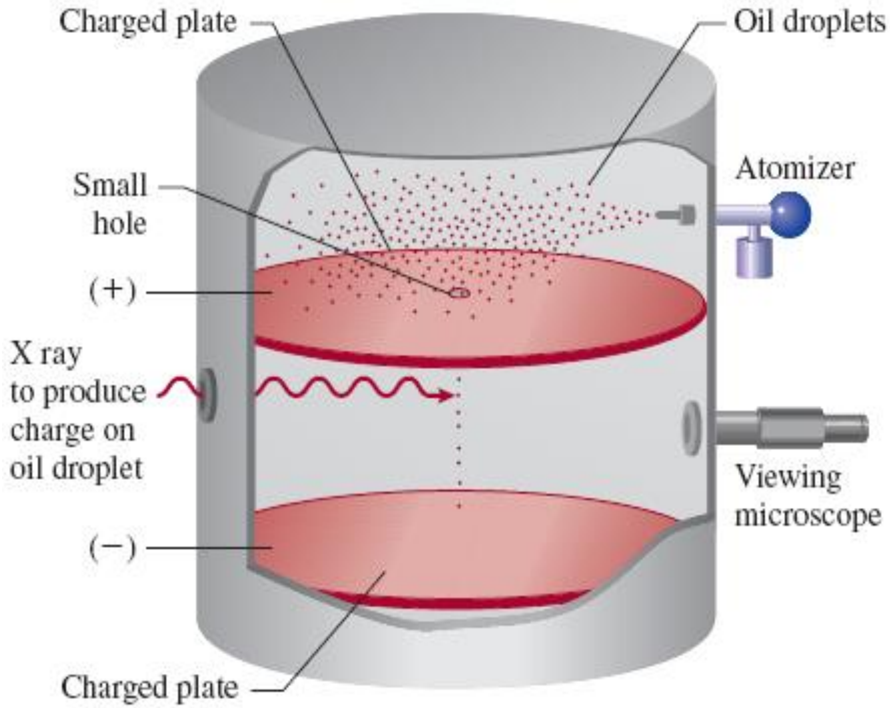
A cathode ray tube with an electric field perpendicular to the direction of the cathode rays and an external magnetic field. The cathode rays will strike the end of the tube at A in the presence of a magnetic field, at C in the presence of an electric field, and at B when there are no external fields present or when the effects of the electric field and magnetic field cancel each other.



Because the cathode ray is attracted by the plate bearing positive charges and repelled by the plate bearing negative charges, it must consist of negatively charged particles (**electrons**).

Joseph Thomson, used a cathode ray tube to determine the ratio of electric charge to the mass of an individual electron (mass/charge of e^-). The number he came up with was **-1.76 x 10⁸ C/g**.

Robert Millikan examined the motion of single tiny drops of oil that picked up static charge from ions in the air. He suspended the charged drops in air by applying an electric field and followed their motions through a microscope.



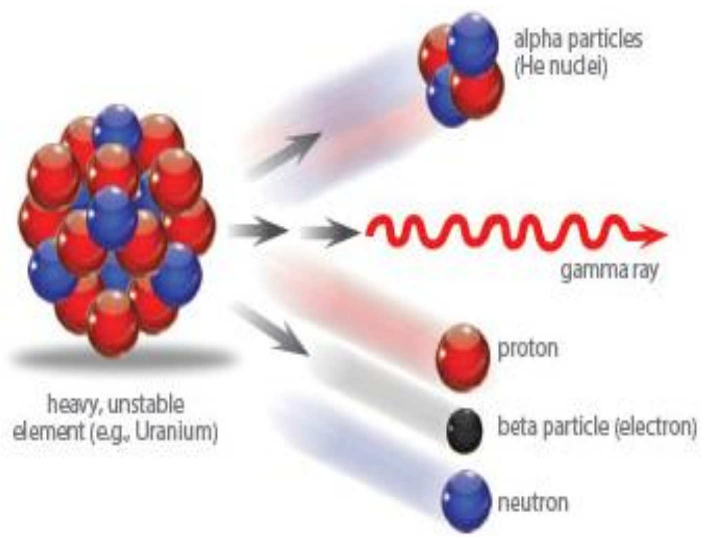
Millikan's experiment

Millikan found the charge of an electron to be -1.6022×10^{-19} C. From these data he calculated the mass of an electron:

$$\begin{aligned} \text{mass of an electron} &= \frac{\text{charge}}{\text{charge/mass}} \\ &= \frac{-1.6022 \times 10^{-19}\,\text{C}}{-1.76 \times 10^{8}\,\text{C/g}} \\ &= 9.10 \times 10^{-28}\ \text{g} \end{aligned}$$

Radioactivity

In 1895, **Wilhelm Röntgen** noticed that cathode rays caused glass and metals to emit very unusual rays. This highly energetic radiation penetrated matter, darkened covered photographic plates, and caused a variety of substances to fluoresce. Because these rays could not be deflected by a magnet, they could not contain charged particles. Röntgen called them **X-rays** because their nature was not known.



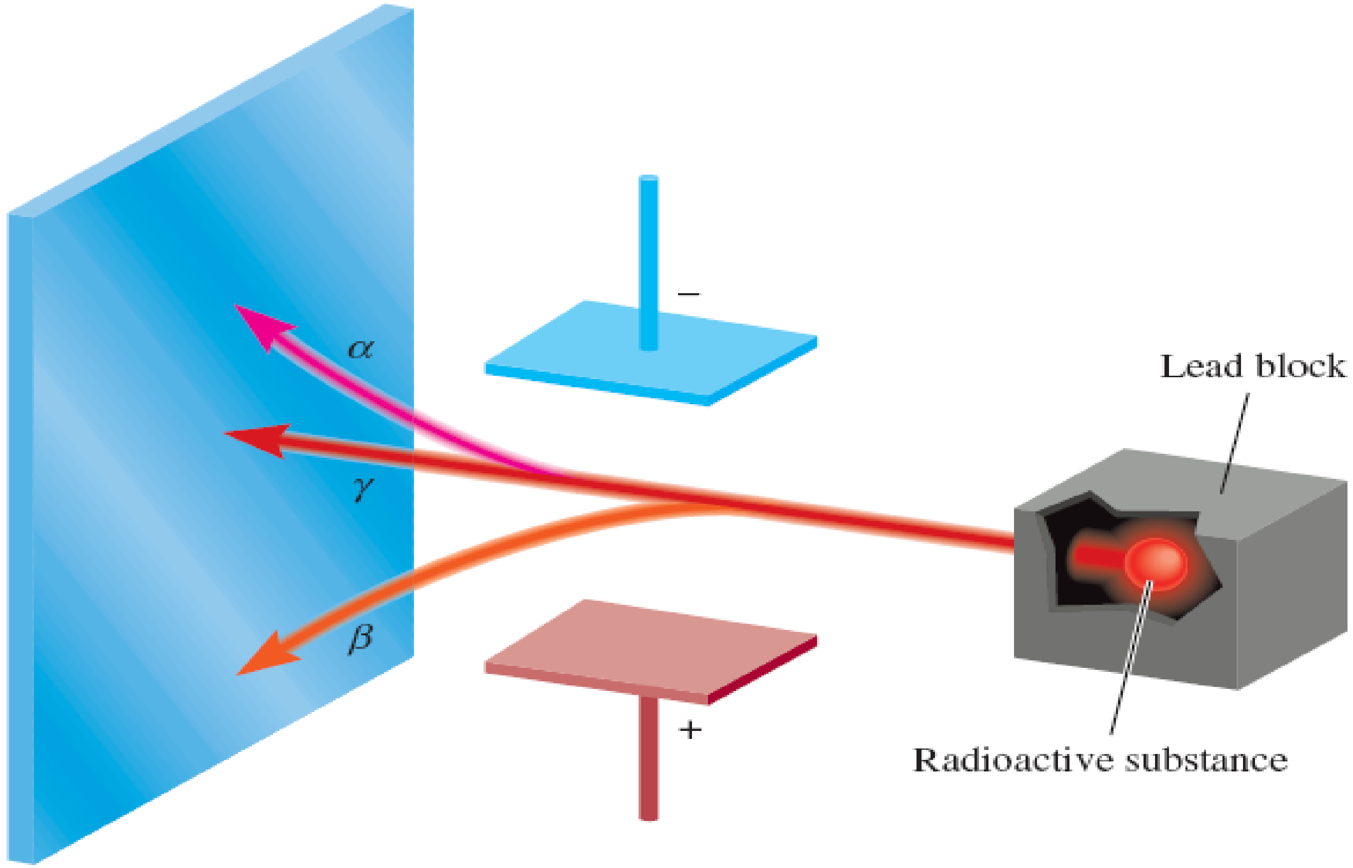
Antoine Becquerel found that exposing thickly wrapped photographic plates to a certain uranium compound caused them to darken, even without the stimulation of cathode rays. Like X-rays, the rays from the uranium compound were highly energetic and could not be deflected by a magnet, but they differed from X-rays because they arose spontaneously. **Marie Curie**, suggested the name radioactivity to describe this spontaneous emission of particles and/or radiation (any element that spontaneously emits radiation is said to be radioactive).

Three types of rays are produced by the decay, or breakdown, of radioactive substances such as uranium.

-Alpha (α) rays consist of positively charged particles, called α particles, and therefore are deflected by the positively charged plate.

-Beta (β) rays, or β **particles**, consist of negatively charged particles (electrons) and are deflected by the negatively charged plate.

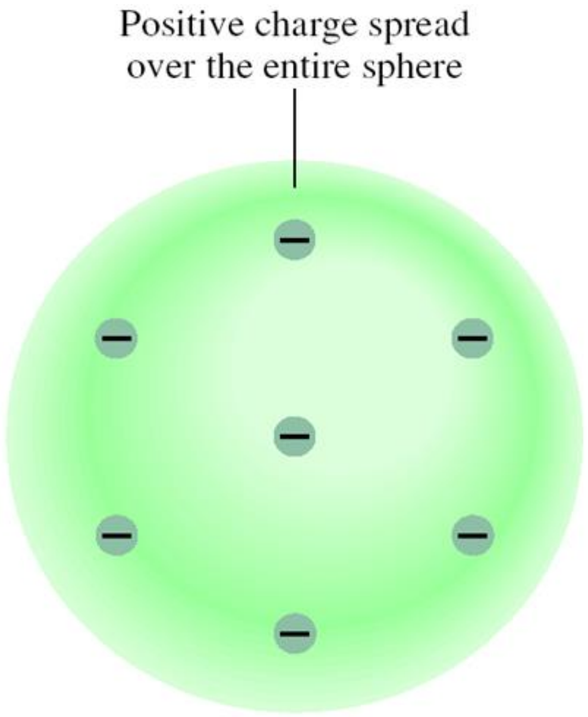
-Gamma (γ) rays, like X-rays, γ rays are high-energy rays and have no charge and are not affected by an external electric field.



The Proton & the Nucleus

By the early 1900s, two features of atoms had become clear: they contain electrons, and they are electrically neutral. To maintain electric neutrality, an atom must contain an equal number of positive and negative charges.

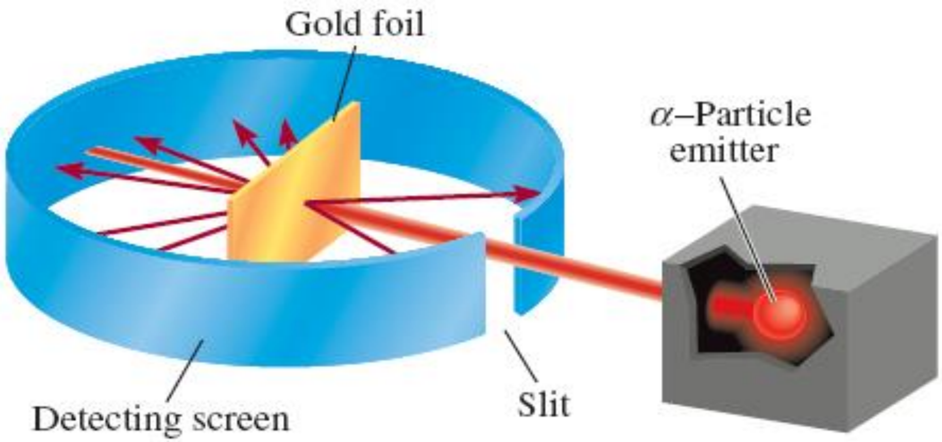
Therefore, Thomson proposed that an atom could be thought of as a uniform, positive sphere of matter in which electrons are embedded like raisins in a cake. This so-called “plum-pudding” model was the accepted theory for a number of years.



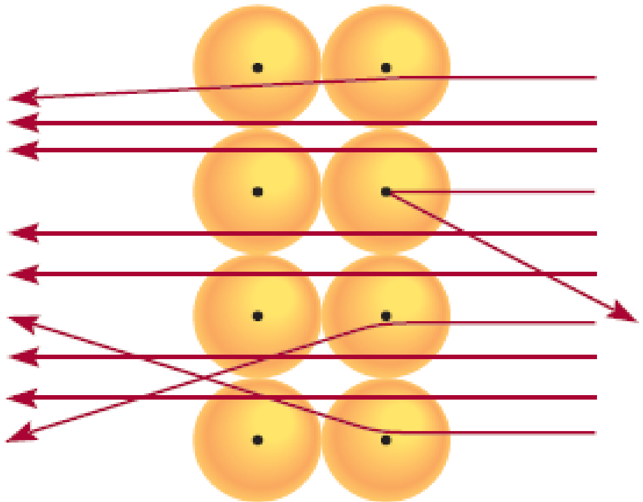
Thomson’s model of the atom “plum-pudding” model, after a traditional English dessert containing raisins. The electrons are embedded in a uniform, positively charged sphere.

In 1910, **Ernest Rutherford** decided to use α particles to probe the structure of atoms. Rutherford carried out a series of experiments using very thin foils of gold and other metals as targets for α particles from a radioactive source.

Rutherford's experiment

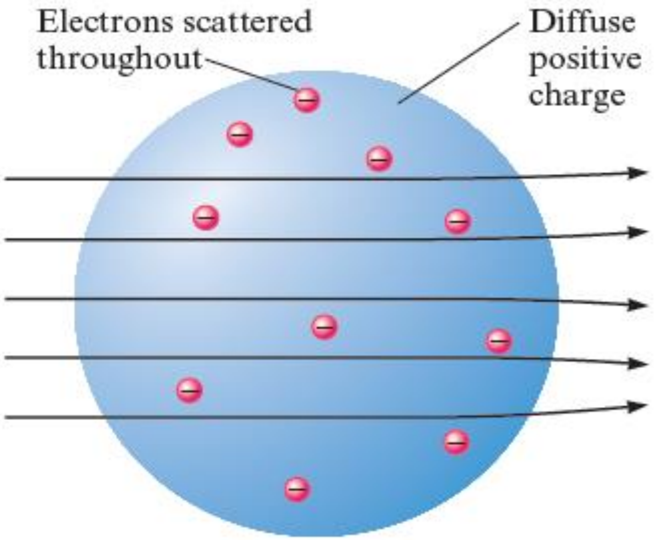


Rutherford's experimental design for measuring the scattering of a particles by a piece of gold foil.

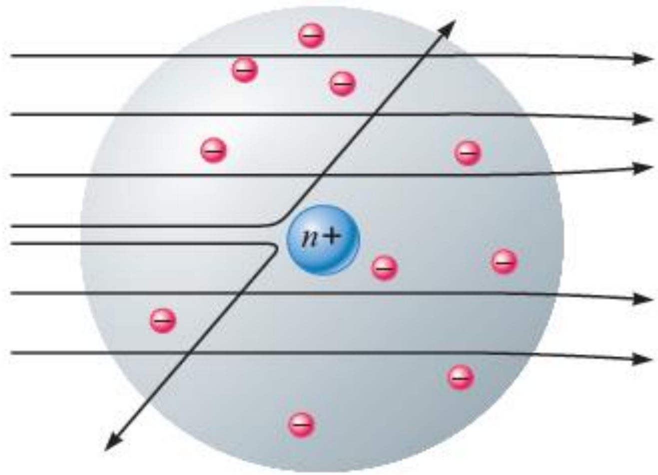


Magnified view of α particles passing through and being deflected by nuclei.

They observed that most α particles passed through the gold foil with little or no deflection. But every now and then an α particle was deflected at wide angle. Occasionally an α particle was turned back. This was a most surprising finding, for in Thomson's model the positive charge of the atom was so diffuse that the positive α particles should have passed through the foil with very little deflection.



The expected results of the metal foil experiment if Thomson's model were correct.



Actual results.

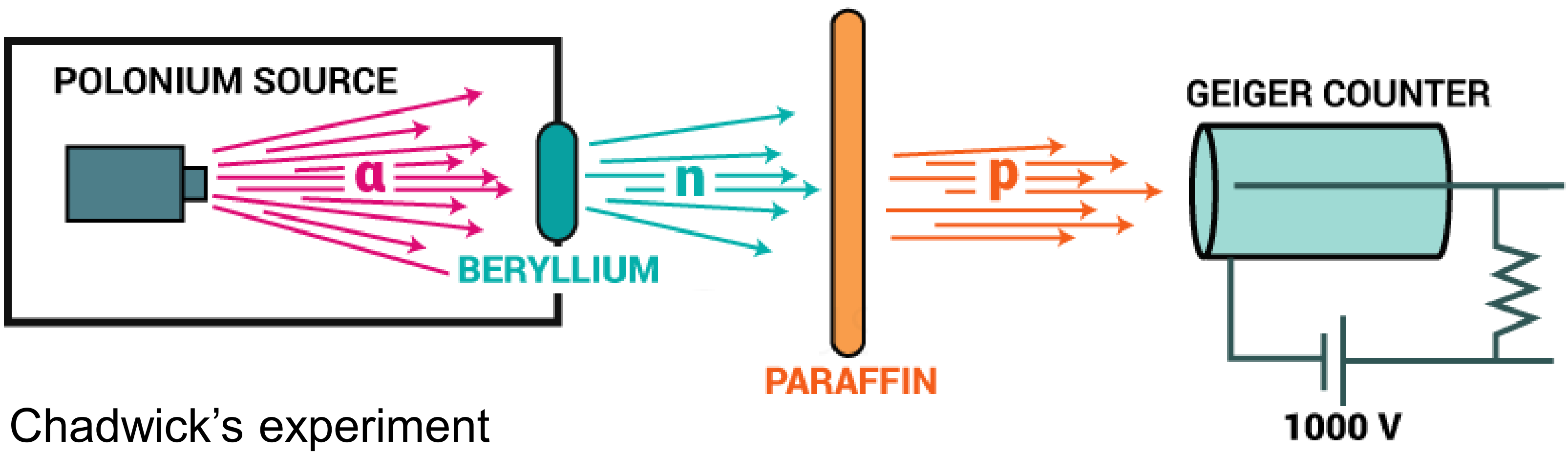
Rutherford's model of the atom

- Most of the atom must be empty space (this explains why the majority of a particles passed through the gold foil with little or no deflection).
- The atom's positive charges, are all concentrated in the **nucleus**, which is a dense central core within the atom.
- Whenever an α particle came close to a nucleus in the scattering experiment, a large repulsive force and therefore a large deflection.
- The positively charged particles in the nucleus are called **protons**.

Each proton carries the same quantity of charge as an electron and has a mass of 1.67262 x 10^{-24} g, about 1840 times the mass of electron.

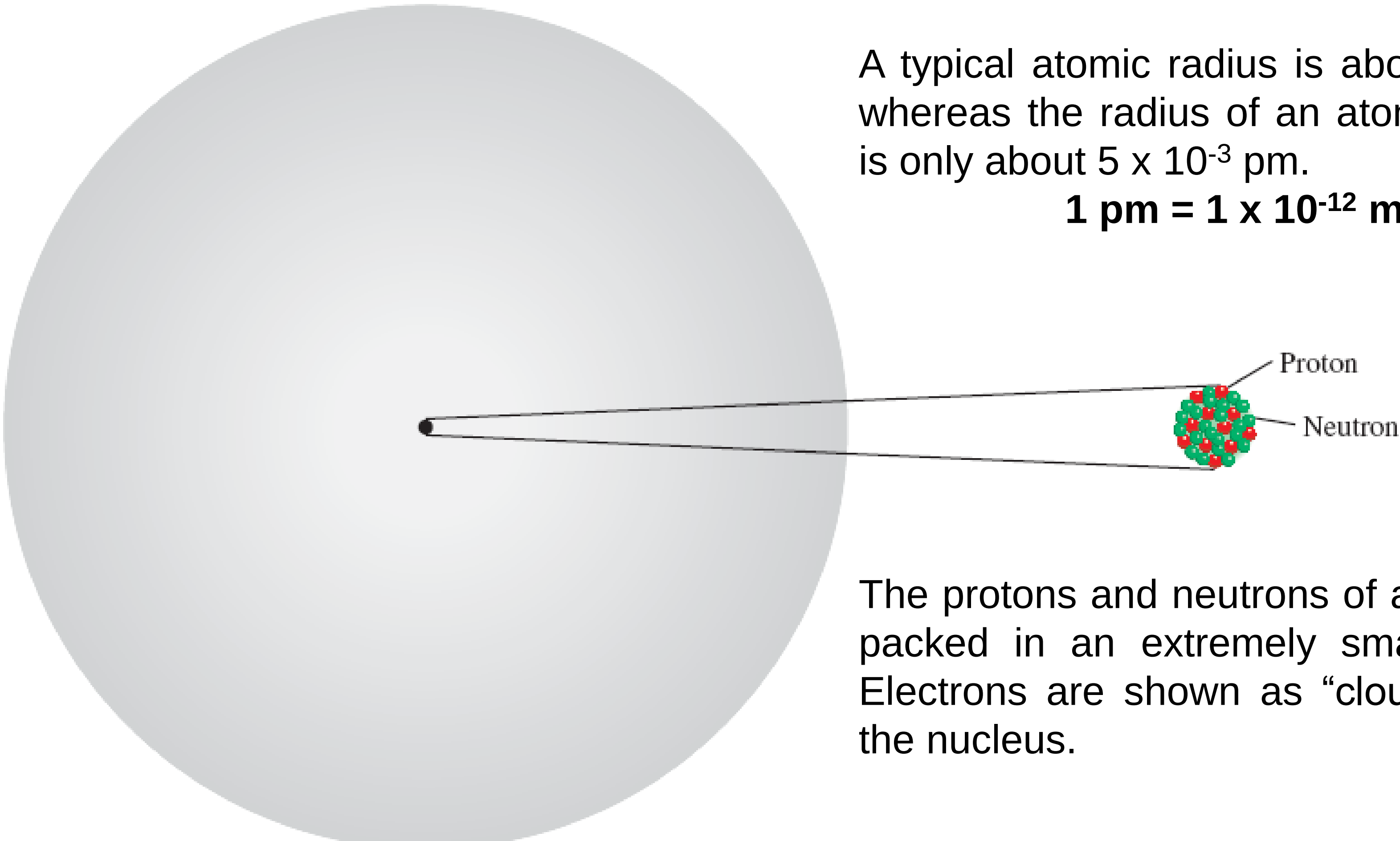
The Neutron

In 1932, **James Chadwick** bombarded a thin sheet of beryllium with α particles, a very high-energy radiation similar to γ rays was emitted by the metal. Later experiments showed that the rays actually consisted of a third type of subatomic particles, which Chadwick named **neutrons**, because they proved to be electrically neutral particles having a mass slightly greater than that of protons.



Chadwick's experiment

The figure shows the location of the elementary particles (protons, neutrons and electrons) in an atom. There are other subatomic particles, but the electron, the proton and the neutron are the three fundamental components of the atom that are important in chemistry.



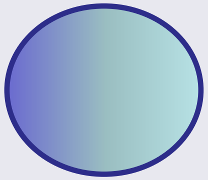
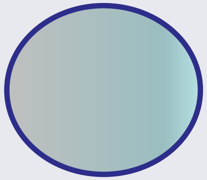
A typical atomic radius is about 100 pm, whereas the radius of an atomic nucleus is only about 5×10^{-3} pm.

$1 \text{ pm} = 1 \times 10^{-12} \text{ m}$

The protons and neutrons of an atom are packed in an extremely small nucleus. Electrons are shown as “clouds” around the nucleus.

mass p ≈ mass n ≈ 1840 x mass e^-

Subatomic particles

Particle	Electron	Proton	Neutron
Symbol	e^- or e	p^+ or p	n^0 or n
Relative size	 size exaggerated		
Actual mass (g)	9.10938×10^{-28}	1.67262×10^{-24}	1.67493×10^{-24}
Mass relative to a proton	1/1836 (0.000545) almost zero	1	1.00138
Mass relative to an electron	1	1836	1839
Charge (Coulomb)	-1.6022×10^{-19}	$+1.6022 \times 10^{-19}$	0
Charge unit (relative charge)	-1	+1	0
Location	Outside nucleus (orbitals)	Inside nucleus	Inside nucleus

2.3
Atomic Number, Mass Number & Isotopes

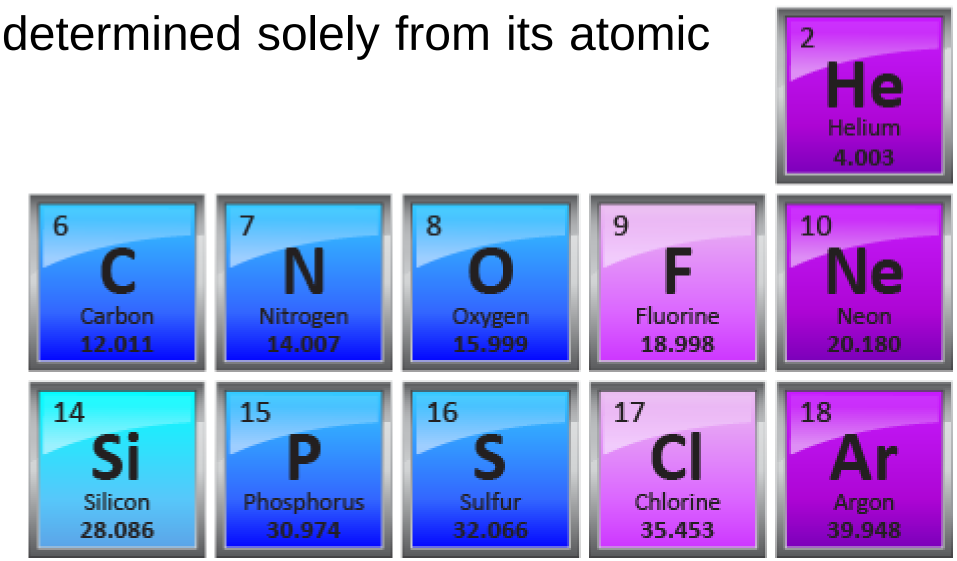
Atomic Number

All atoms can be identified by the number of protons and neutrons they contain.

The atomic number (Z) is the number of protons in the nucleus of each atom of an element. In a neutral atom the number of protons is equal to the number of electrons, so the atomic number also indicates the number of electrons present in the atom.

The chemical identity of an atom can be determined solely from its atomic number.

e.g., the atomic number of fluorine is 9. This means that each fluorine atom has 9 protons and 9 electrons. Or, viewed another way, every atom in the universe that contains 9 protons is correctly named “fluorine”.



Mass Number

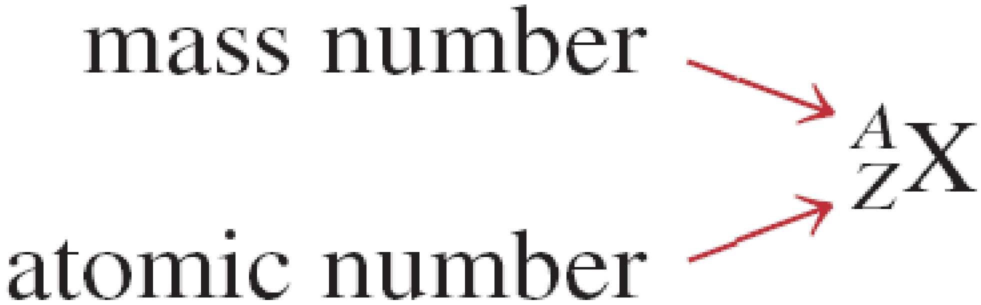
The mass number (A) is the total number of neutrons and protons present in the nucleus of an atom of an element. Except for the most common form of hydrogen, which has one proton and no neutrons, all atomic nuclei contain both protons and neutrons. In general, the mass number is given by

$$\boxed{\begin{aligned}\textbf{mass number} &= \textbf{number of protons + number of neutrons}\\ &= \textbf{atomic number + number of neutrons}\end{aligned}}$$

Protons and neutrons are collectively called *nucleons*.

The number of neutrons in an atom is equal to the difference between the mass number and the atomic number, or (A - Z). e.g., if the mass number of a particular boron atom is 12 and the atomic number is 5 (indicating 5 protons in the nucleus), then the number of neutrons is 12 - 5 = 7.

The accepted way to denote the atomic number and mass number of an atom of an element (X) is as follows:



Note that all three quantities (atomic number, number of neutrons, and mass number) must be positive integers, or whole numbers.

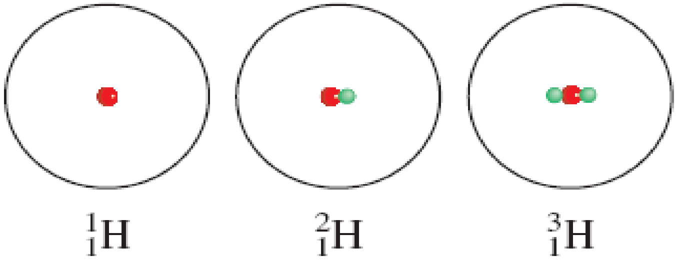
Isotopes

Atoms of a given element do not all have the same mass. Most elements have two or more isotopes, atoms that have the same atomic number but different mass numbers due to the difference in neutron number.

e.g., there are three isotopes of hydrogen. One, simply known as hydrogen, has one proton and no neutrons. The deuterium isotope contains one proton and one neutron, and tritium has one proton and two neutrons.

Thus, for the isotopes of hydrogen, we write

$^{1}_{1}\mathrm{H}$	$^{2}_{1}\mathrm{H}$	$^{3}_{1}\mathrm{H}$
hydrogen	deuterium	tritium



As another example, consider two common isotopes of uranium with mass numbers of 235 and 238, respectively:

$$^{235}_{92}\mathrm{U} \qquad ^{238}_{92}\mathrm{U}$$

EXAMPLE

How many protons, neutrons, and electrons are in ${}^{12}_{6}C$, ${}^{13}_{6}C$, and ${}^{14}_{6}C$

The atomic number of carbon is 6, which means that every carbon atom has 6 protons and 6 electrons, so that the neutron numbers of these isotopes are 6, 7 and 8 respectively.

${}^{12}_{6}C$, 12 – 6 = 6 neutrons

${}^{13}_{6}C$, 13 – 6 = 7 neutrons

${}^{14}_{6}C$, 14 – 6 = 8 neutrons

EXAMPLE

Give the number of protons, neutrons and electrons in each of the following species:

(a) $^{20}_{11}Na$

The atomic number is 11, so there are 11 protons. The mass number is 20, so the number of neutrons is 20 - 11 = 9. The number of electrons is 11 (the same as the number of protons).

(b) $^{22}_{11}Na$

The atomic number is 11. The mass number is 22, so the number of neutrons is 22 - 11 = 11. The number of electrons is 11. Note that the species in (a) and (b) are chemically similar isotopes of sodium.

(c) ^{17}O

The atomic number of O (oxygen) is 8, so there are 8 protons. The mass number is 17, so there are 17 - 8 = 9 neutrons. There are 8 electrons.

(d) carbon-14.

Carbon-14 can also be represented as ^{14}C. The atomic number of carbon is 6, so there are 14 - 6 = 8 neutrons. The number of electrons is 6.

Practice Exercise

How many protons, neutrons and electrons are in the following isotope of copper: ^{63}Cu?

2.4
The Periodic Table

Periodic Table of the Elements

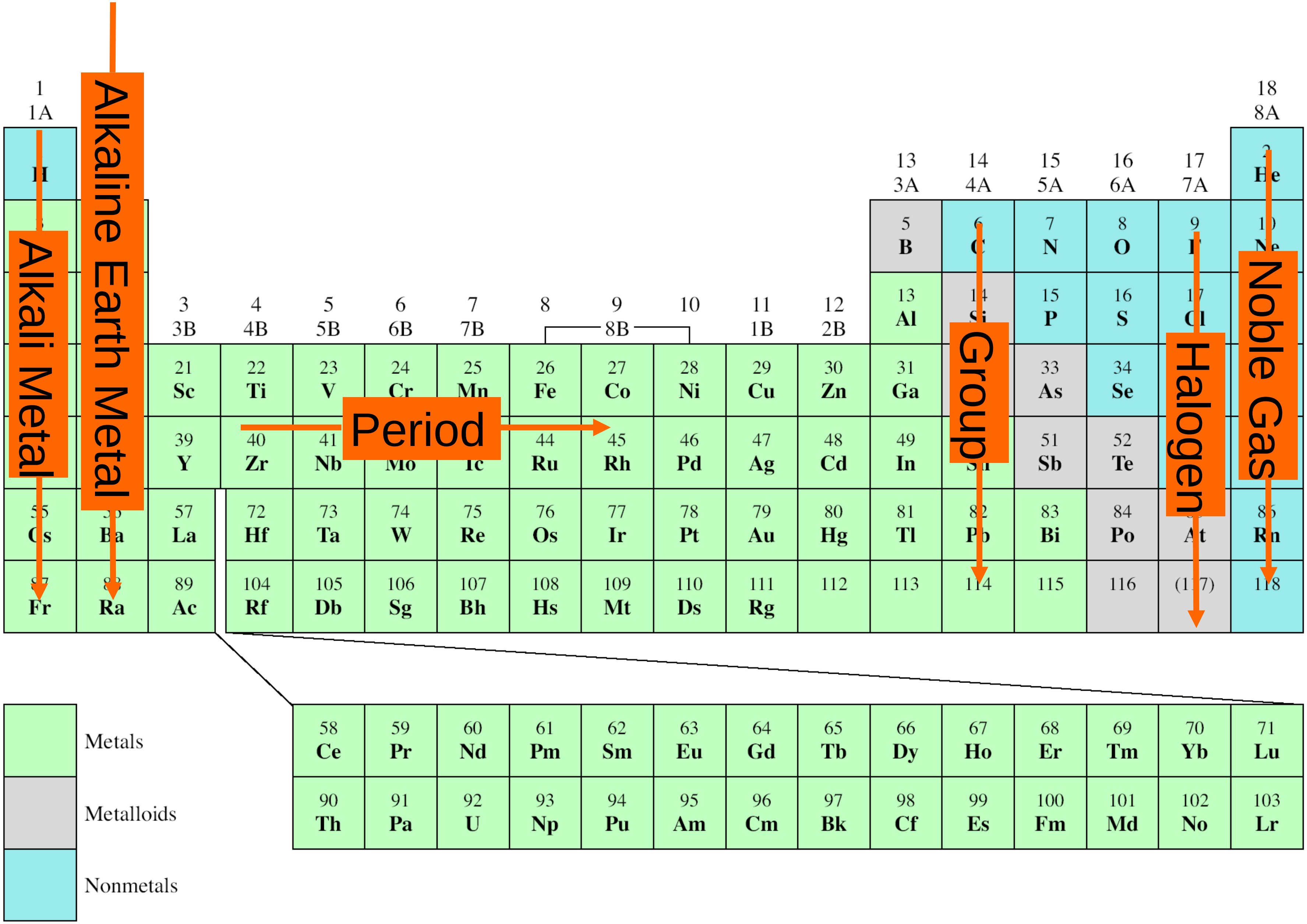
1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	113	114	115	116	(117)	118

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Metals

Metalloids

Nonmetals



Alkali Metal
Alkaline Earth Metal
Period
Group
Halogen
Noble Gas
Metals
Metalloids
Nonmetals

Periodic Table of the Elements

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
															Pnictogens	Chalcogens	Halogens	
1	1 H Hydrogen 1.008	Atomic # Symbol Name Weight																2 He Helium 4.0026
2	3 Li Lithium 6.94	4 Be Beryllium 9.0122											5 B Boron 10.81	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
3	11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminium 26.982	14 Si Silicon 28.085	15 P Phosphorus 30.974	16 S Sulfur 32.06	17 Cl Chlorine 35.45	18 Ar Argon 39.948
4	19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.630	33 As Arsenic 74.922	34 Se Selenium 78.971	35 Br Bromine 79.904	36 Kr Krypton 83.798
5	37 Rb Rubidium 85.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.95	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29
6	55 Cs Caesium 132.91	56 Ba Barium 137.33	57–71	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
7	87 Fr Francium (223)	88 Ra Radium (226)	89–103	104 Rf Rutherfordium (267)	105 Db Dubnium (268)	106 Sg Seaborgium (269)	107 Bh Bohrium (270)	108 Hs Hassium (277)	109 Mt Meitnerium (278)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (282)	112 Cn Copernicium (285)	113 Nh Nihonium (286)	114 Fl Flerovium (289)	115 Mc Moscovium (290)	116 Lv Livermorium (293)	117 Ts Tennessine (294)	118 Og Oganesson (294)

C Solid | Hg Liquid | H Gas | Rf Unknown

Metals: Alkali metals; Alkaline earth metals; Lanthanoids (Lanthanides); Actinoids (Actinides); Transition metals; Post-transition metals

Metalloids

Nonmetals: Other nonmetals; Noble gases

For elements with no stable isotopes, the mass number of the isotope with the longest half-life is in parentheses.

6	57 La Lanthanum 138.91	58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.05	71 Lu Lutetium 174.97
7	89 Ac Actinium (227)	90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (266)

Periodic table is a chart in which elements having similar chemical and physical properties are grouped together.

More than half of the known elements were discovered between 1800 and 1900. To date, 118 elements have been positively identified. Most of them occur naturally on Earth. The others have been created by scientists via nuclear processes.

Elements are arranged by atomic number in horizontal rows called **periods** and in vertical columns known as **groups** or families, according to similarities in their chemical properties.

The elements can be divided into three categories; metals, nonmetals and metalloids. A metal is a good conductor of heat and electricity while a nonmetal is usually a poor conductor of heat and electricity. A metalloid has properties that are intermediate between those of metals and nonmetals.

Elements are often referred to collectively by their periodic table group number (Group 1A, Group 2A, and so on). However, some element groups have been given special names. The Group 1A elements (Li, Na, K, Rb, Cs & Fr) are called **alkali metals**, and the Group 2A elements (Be, Mg, Ca, Sr, Ba & Ra) are called **alkaline earth metals**. Elements in Group 7A (F, Cl, Br, I & At) are known as **halogens**, and elements in Group 8A (He, Ne, Ar, Kr, Xe & Rn) are called **noble gases**, or rare gases.

The 1–18 group designation has been recommended by the IUPAC, but the standard U.S. notation for group numbers (1A–8A and 1B–8B) is most widely used.

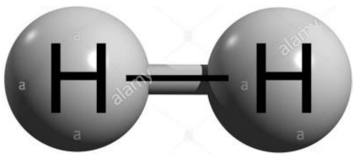
2.5
Molecules and Ions

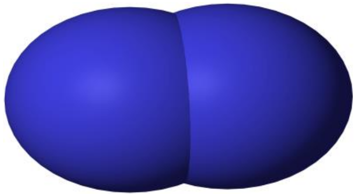
Molecules

A **molecule** is an aggregate of at least two atoms in a definite arrangement held together by chemical forces (also called chemical bonds).

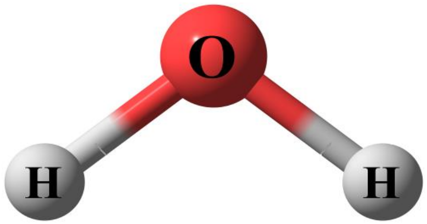
A molecule may contain atoms of the same element or atoms of two or more elements joined in a fixed ratio, in accordance with the law of definite proportions. Thus, a molecule is not necessarily a compound, which, by definition, is made up of two or more elements. Like atoms, molecules are electrically neutral.

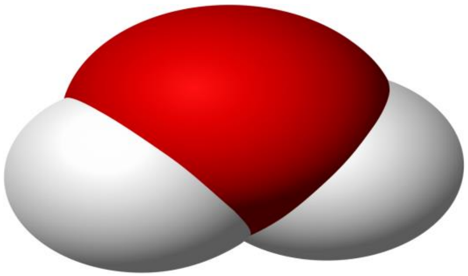
e.g., **Hydrogen gas (H_2)**; is a pure element, but it consists of molecules made up of two H atoms each.

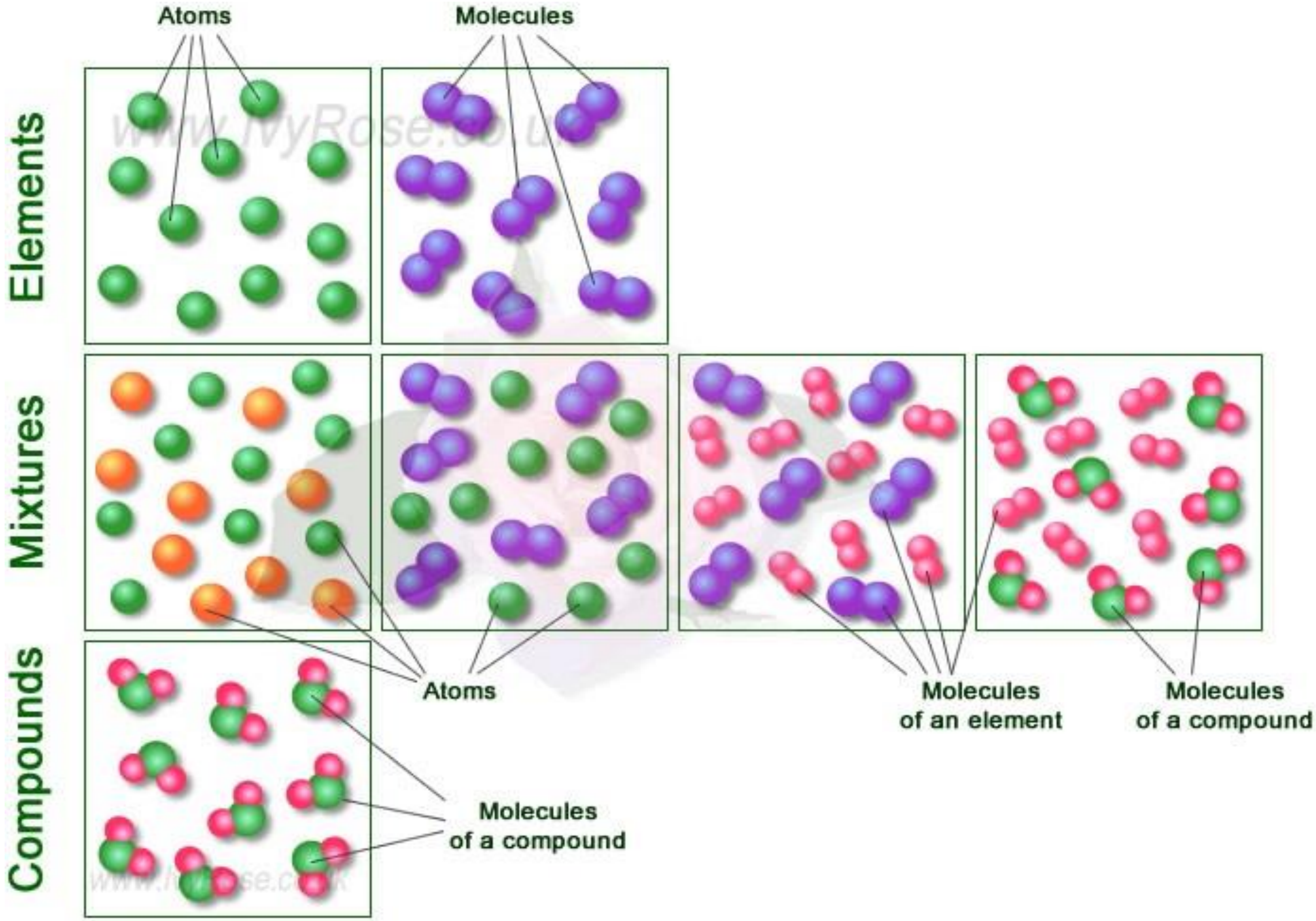




e.g., **Water (H_2O)**; is a molecular compound that contains hydrogen and oxygen in a ratio of two H atoms and one O atom.



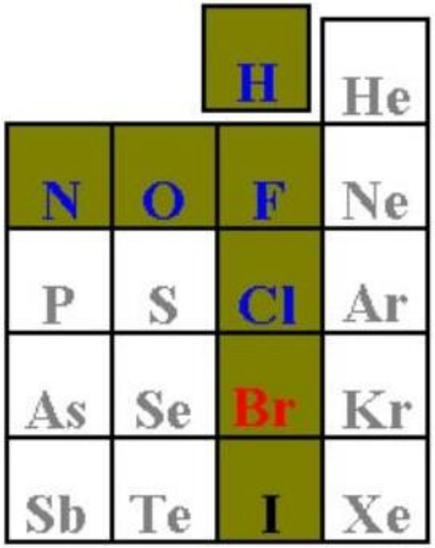




Atoms: represented by single spheres.
Elements: represented by the spheres of the same kind; same size and color.
Molecules: represented by two or more spheres jointed together.
Molecules of elements: represented by two or more spheres of the same kind jointed together.
Molecules of compounds: represented by two or more spheres of the different kind jointed together.

Diatomic molecules

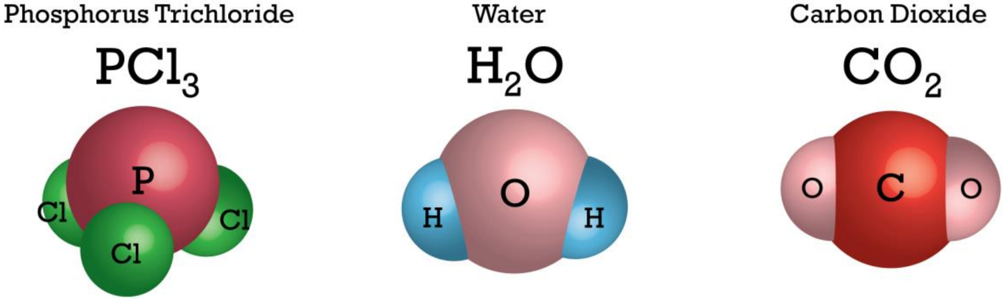
The hydrogen molecule (H_2), is called a **diatomic molecule** because it contains only two atoms. Other elements that normally exist as diatomic molecules are nitrogen (N_2) and oxygen (O_2), as well as the Group 7A elements; fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2). A diatomic molecule can contain atoms of different elements. e.g., hydrogen chloride (HCl) and carbon monoxide (CO).



The seven diatomic molecules

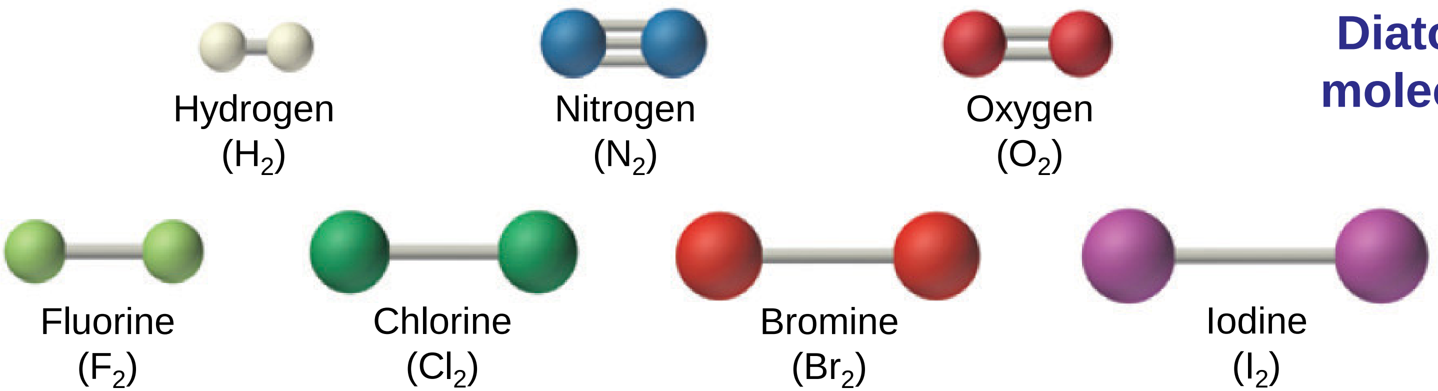
Polyatomic molecules

The vast majority of molecules contain more than two atoms. They can be atoms of the same element, as in ozone (O_3), which is made up of three atoms of oxygen, or they can be combinations of two or more different elements. Molecules containing more than two atoms are called **polyatomic molecules.** Like ozone, water (H_2O) and ammonia (NH_3) are polyatomic molecules.

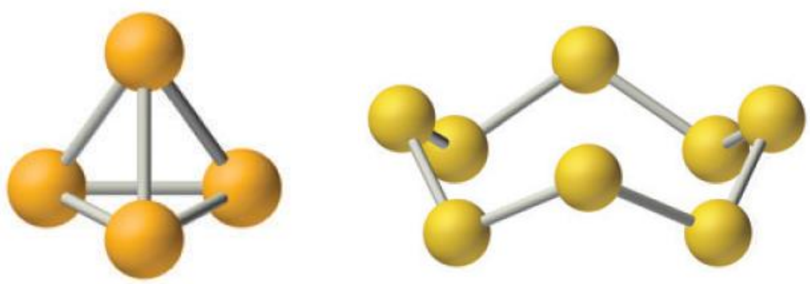


Homonuclear molecules

Molecules composed of only one type of element. Homonuclear molecules may consist of various numbers of atoms. Homonuclear diatomic molecules include H_2, O_2, N_2 and all of the halogens (F_2, Cl_2, Br_2 and I_2). Ozone (O_3) is a triatomic homonuclear molecule. Homonuclear tetratomic molecules include arsenic (As_4) and phosphorus (P_4).



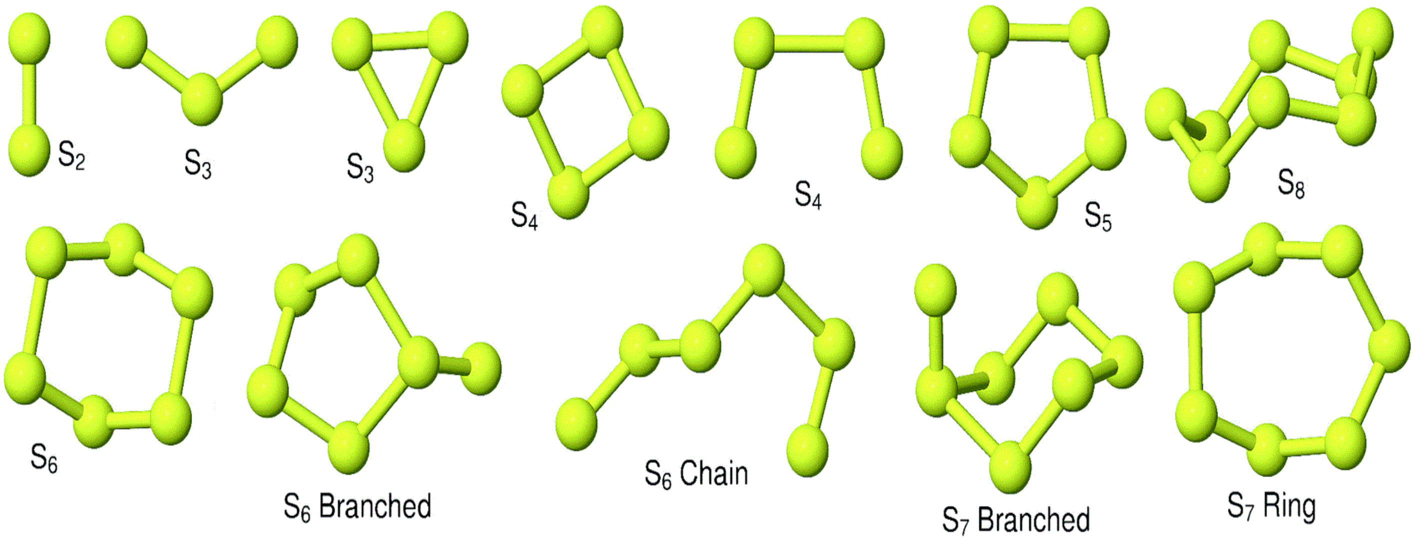
Diatomic molecules

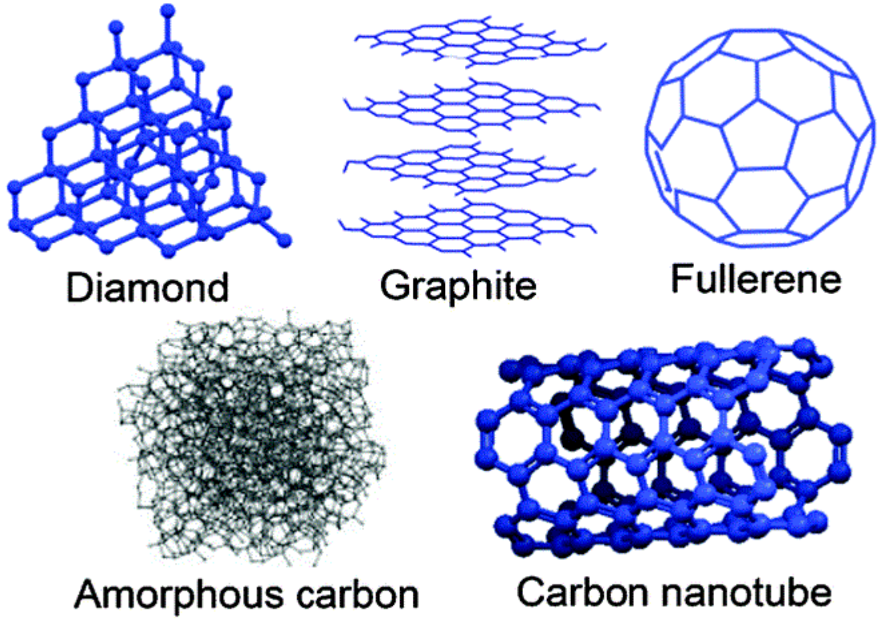


Polyatomic molecules

Allotropes are different chemical forms of the same element (not containing any other element). In that sense, allotropes are all homonuclear.

e.g., Sulfur forms several allotropes containing different numbers of sulfur atoms.





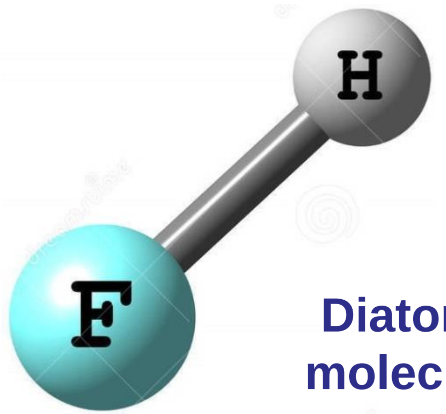
e.g., The element carbon is known to have a number of homonuclear allotrope molecules.

Heteronuclear molecules

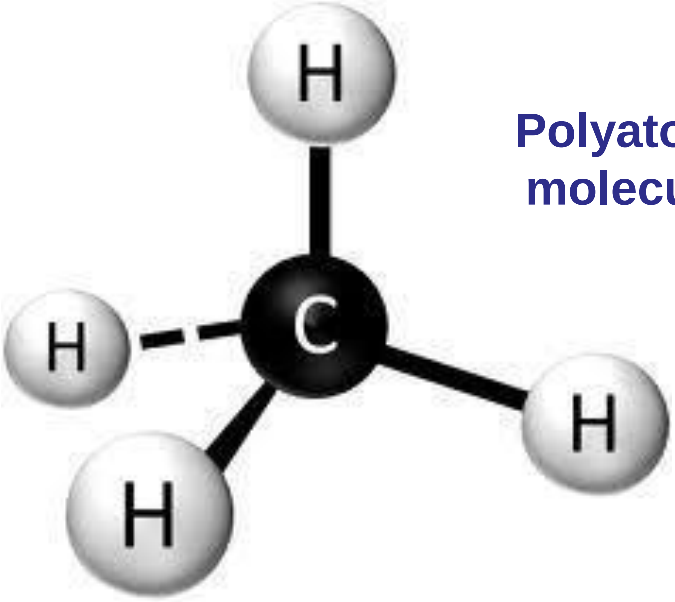
Molecules composed of more than one type of element.

e.g., HCl, HF and CO are diatomic heteronuclear molecules.

e.g., CO_2, CH_4 and NH_3 are polyatomic heteronuclear molecules.



Diatomic molecules



Polyatomic molecules

Ions

An **ion** is an atom or a group of atoms that has a net positive or negative charge.

The number of positively charged protons in the nucleus of an atom remains the same during ordinary chemical changes (chemical reactions), but negatively charged electrons may be lost or gained.

The loss of one or more electrons from a neutral atom results in a **cation,** an ion with a net positive charge. e.g., a sodium atom (Na) can readily lose an electron to become a sodium cation, Na^+:

Na Atom	Na^+ Ion
11 protons	11 protons
11 electrons	10 electrons

Cl Atom	Cl^- Ion
17 protons	17 protons
17 electrons	18 electrons

On the other hand, an **anion** is an ion whose net charge is negative due to an increase in the number of electrons. A chlorine atom (Cl), for instance, can gain an electron to become the chloride ion, Cl^-:

Sodium chloride (NaCl), ordinary table salt, is called an **ionic compound** because it is formed from cations and anions.

An atom can lose or gain more than one electron. Examples of ions formed by the loss or gain of more than one electron are Mg^{2+}, Fe^{3+}, S^{2-} and N^{3-}. These ions are called **Monatomic ions** because they contain only one atom.

With very few exceptions, **metals** tend to form **cations** and **nonmetals** form **anions**.

In addition, two or more atoms can combine to form an ion that has a net positive or net negative charge. **Polyatomic ions** such as OH^- (hydroxide ion), CN^- (cyanide ion), and NH_4^+ (ammonium ion) are ions containing more than one atom.

Common Polyatomic Ions		
Ammonium NH_4^+	Nitrate NO_3^-	Thiocyanate SCN^-
Acetate CH_3COO^- or $C_2H_3O_2^-$	Nitrite NO_2^-	Thiosulfate $S_2O_3^{2-}$
Bromate BrO_3^-	Hydroxide OH^-	Hypochlorite ClO^-
Carbonate CO_3^{2-}	Perchlorate ClO_4^-	Sulfate SO_4^{2-}
Chlorate ClO_3^-	Periodate IO_4^-	Sulfite SO_3^{2-}
Chlorite ClO_2^-	Permanganate MnO_4^-	Iodate IO_3^-
Chromate CrO_4^{2-}	Peroxide O_2^{2-}	Silicate SiO_4^{4-}
Cyanide CN^-	Phosphate PO_4^{3-}	Oxalate $C_2O_4^{2-}$
Dichromate $Cr_2O_7^{2-}$	Phosphite PO_3^{3-}	Hydrogen carbonate HCO_3^-

2.6
Chemical Formulas

Molecular Formulas

A **molecular formula** shows the exact number of atoms of each element in the smallest unit of a substance. The molecular formula tells us the actual number of atoms in a molecule.

e.g., $\mathbf{H_2}$ is the molecular formula for hydrogen, $\mathbf{O_2}$ is oxygen, $\mathbf{O_3}$ is ozone, and $\mathbf{H_2O}$ is water. The subscript numeral indicates the number of atoms of an element present. There is no subscript for **O** in $\mathbf{H_2O}$ because there is only one atom of oxygen in a molecule of water, and so the number "one" is omitted from the formula.

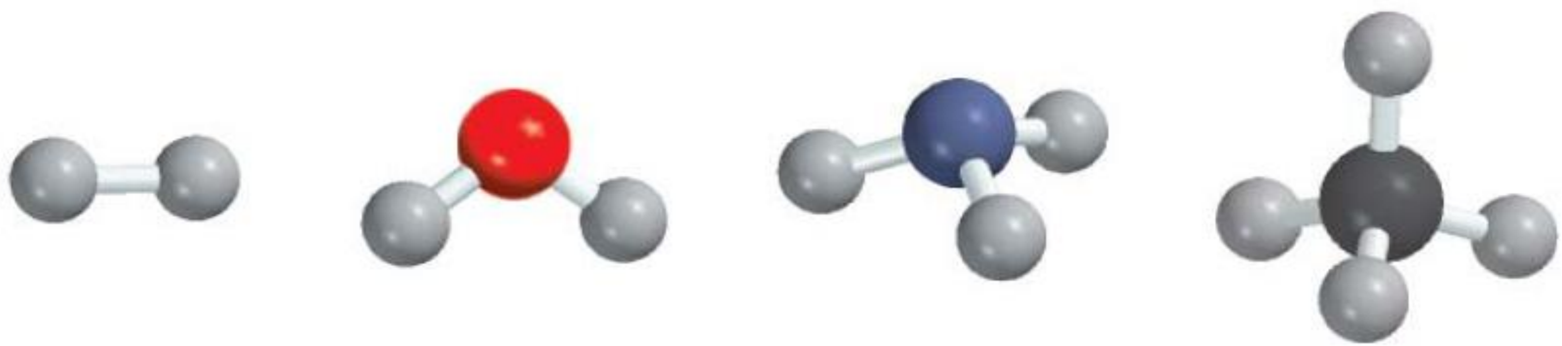
The **structural formula** shows how atoms are bonded to one another in a molecule.

e.g., it is known that each of the two **H** atoms is bonded to an **O** atom in the water molecule. Therefore, the structural formula of water is **H—O—H**. A line connecting the two atomic symbols represents a chemical bond.

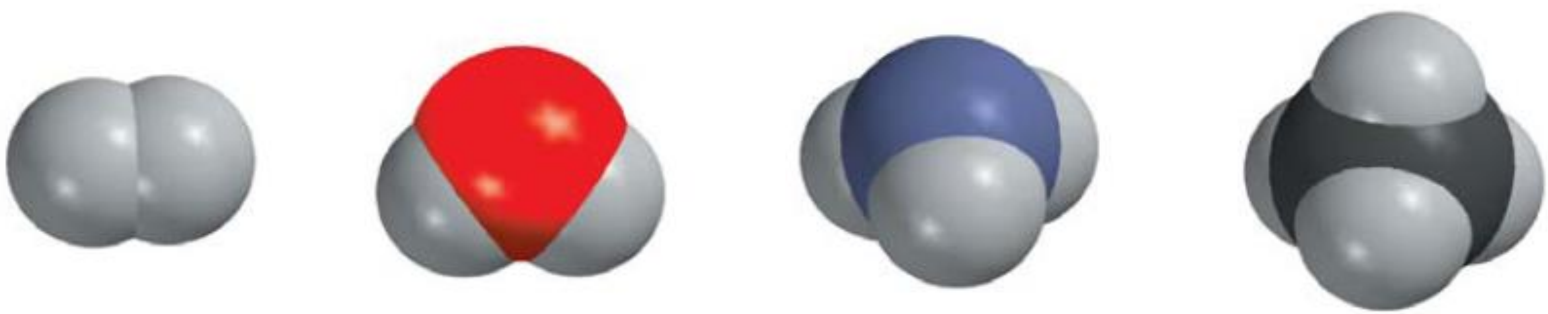
Molecular Models

	Hydrogen	Water	Ammonia	Methane
Molecular formula	H_2	H_2O	NH_3	CH_4
Structural formula	H—H	H—O—H	H—N—H | H	H | H—C—H | H

Ball-and-stick model



Space-filling model



Molecular and structural formulas and molecular models of four common molecules

Empirical Formulas

Empirical formulas are the simplest chemical formulas; they are written by reducing the subscripts in the molecular formulas to the smallest possible whole numbers.

Molecular formulas are the true formulas of molecules. If we know the molecular formula, we also know the empirical formula, but the reverse is not true.

e.g.,

Molecular formula	Empirical formula
H_2O_2	HO
N_2H_4	NH_2
C_6H_{14}	C_3H_7
H_2O	H_2O
CH_4	CH_4

e.g., Glucose ($C_6H_{12}O_6$), ribose ($C_5H_{10}O_5$), acetic acid ($C_2H_4O_2$), and formaldehyde (CH_2O) all have different molecular formulas but the same empirical formula: CH_2O.

EXAMPLE

Write the empirical formulas for the following molecules:

(a) acetylene (C_2H_2), which is used in welding torches,
Solution: CH

(b) glucose ($C_6H_{12}O_6$), a substance known as blood sugar, and
Solution: CH_2O

(c) nitrous oxide (N_2O), a gas that is used as an anesthetic gas "laughing gas" and as an aerosol propellant for whipped creams.
Solution: N_2O (the empirical formula for nitrous oxide is the same as its molecular formula).

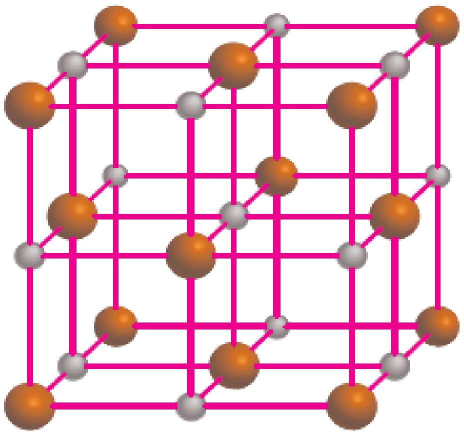
Practice Exercise

Write the empirical formula for caffeine ($C_8H_{10}N_4O_2$), a stimulant found in tea and coffee.

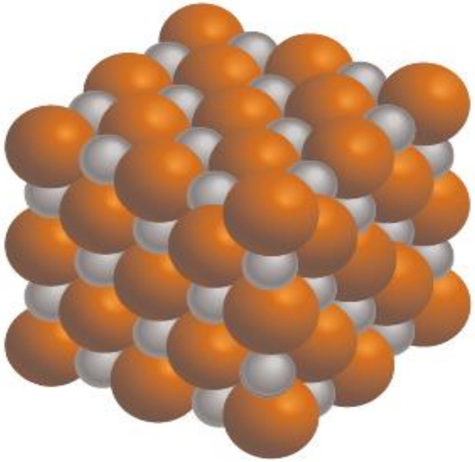
Formula of Ionic Compounds

Ionic compounds are chemical compounds composed of ions held together by electrostatic forces termed ionic bonding. The compound is neutral overall, but consists of positively charged ions called **cations** and negatively charged ions called **anions**.

e.g., a solid sample of sodium chloride (NaCl) consists of equal numbers of Na^+ and Cl^- ions arranged in a three-dimensional network. In such a compound there is a 1:1 ratio of cations to anions so that the compound is electrically neutral.



Structure of solid NaCl



In reality, the cations are in contact with the anions

Crystals of NaCl

The smaller spheres represent Na^+ ions and the larger spheres, Cl^- ions.

The arrangement of cations and anions is such that the compounds are all electrically neutral. For ionic compounds to be electrically neutral, the sum of the charges on the cation and anion in each formula unit must be zero.

Potassium Bromide

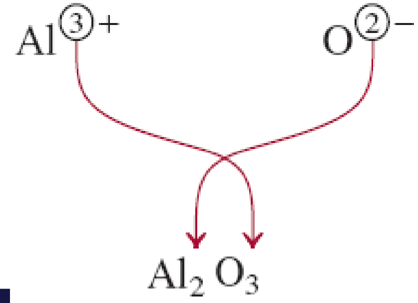
Potassium cation K^+ and bromine anion Br^- combine to form the ionic compound potassium bromide. The sum of the charges is +1 + (-1) = 0, so no subscripts are necessary. The formula is KBr.

Zinc Iodide

Zinc cation Zn^{2+} and iodine anion I^- combine to form zinc iodide. The sum of the charges of one Zn^{2+} ion and one I^- ion is +2 + (-1) = +1. To make the charges add up to zero we multiply the -1 charge of the anion by 2 and add the subscript "2" to the symbol for iodine. Therefore the formula for zinc iodide is ZnI_2.

Aluminum Oxide

The cation is Al^{3+} and the oxygen anion is O^{2-}. The following diagram helps to determine the subscripts for the compound:



The sum of the charges is 2(+3) + 3(-2) = 0.
Thus, the formula for aluminum oxide is Al_2O_3.

EXAMPLES

Calcium Bromide

$1 \times +2 = +2$ $2 \times -1 = -2$

$CaBr_2$

Ca^{2+} Br^-

$+2 + (-2) = 0$

Sodium Carbonate

$2 \times +1 = +2$ $1 \times -2 = -2$

Na_2CO_3

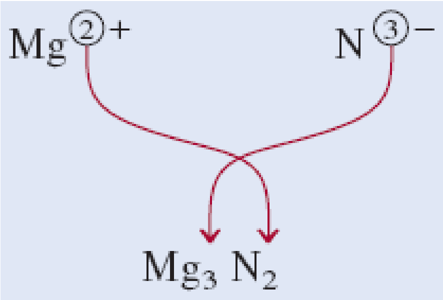
Na^+ CO_3^{2-}

$+2 + (-2) = 0$

EXAMPLE

Write the formula of magnesium nitride, containing the Mg^{2+} and N^{3-} ions.

$$3 \times (+2) + 2 \times (-3) = 0$$



Practice Exercise

Write the formulas of the following ionic compounds:

(a) Chromium sulfate (containing the Cr^{3+} and SO_4^{2-} ions),
Solution: $Cr_2(SO_4)_3$

(b) Titanium oxide (containing the Ti^{4+} and O^{2-} ions).
Solution: Ti_2O_4

2.7
Naming Compounds
(ionic, molecular, acids, bases & hydrates)

Ionic Compounds

Ionic compounds are made up of cations (+ve ions) and anions (-ve ions).

All cations of interest to us are derived from metal atoms (with the important exception of the ammonium ion, NH_4^+).

Metal cations take their names from the elements.

e.g.,

Element	Name	Cation	Name of Cation
Na	sodium	Na^+	sodium ion (or sodium cation)
K	potassium	K^+	potassium ion (or potassium cation)
Mg	magnesium	Mg^{2+}	magnesium ion (or magnesium cation)
Al	aluminum	Al^{3+}	aluminum ion (or aluminum cation)

The anion is named by taking the first part of the element name and adding "-ide."

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^{-})

Many ionic compounds are **binary compounds,** or compounds formed from just two elements. For binary compounds, the first element named is the metal cation, followed by the nonmetallic anion.

	Name
NaCl	sodium chloride
KBr	potassium bromide
ZnI_2	zinc iodide
Al_2O_3	aluminum oxide

Na, K, Zn and Al are **metals**

Cl, Br, I and O are **nonmetals**

The “-ide” ending is also used for certain anion groups containing different elements, such as hydroxide (OH^-) and cyanide (CN^-).

e.g.,

LiOH: lithium hydroxide

KCN: potassium cyanide

These and a number of other such ionic substances are called **ternary compounds,** meaning compounds consisting of three elements.

Certain metals, especially the **transition metals**, can form more than one type of cation.

e.g., iron can form two cations: Fe^{2+} and Fe^{3+}. An older nomenclature system that is still in limited use assigns the ending “-ous” to the cation with fewer positive charges and the ending “-ic” to the cation with more positive charges:

Fe^{2+}: ferrous ion

Fe^{3+}: ferric ion

$FeCl_2$ (left) & $FeCl_3$ (right)

The names of the compounds that these iron ions form with chlorine would thus be

$FeCl_2$: ferrous chloride

$FeCl_3$: ferric chloride

Modern nomenclature system called **Stock system**. In this system, the **Roman numeral** I indicates one positive charge, II means two positive charges, and so on.

e.g., manganese (Mn) atoms can assume several different positive charges:

Mn^{2+}: MnO manganese(II) oxide
Mn^{3+}: Mn_2O_3 manganese(III) oxide
Mn^{4+}: MnO_2 manganese(IV) oxide

Thus the names of the previous iron compounds would be:

$FeCl_2$: ferrous chloride becomes iron(II) chloride
$FeCl_3$: ferric chloride becomes iron(III) chloride

TABLE 2.3 in the textbook

Names and formulas of some common inorganic cations and anions

EXAMPLE

Name the following compounds:

(a) $Cu(NO_3)_2$, copper(II) nitrate.

(b) KH_2PO_4, potassium dihydrogen phosphate.

(c) NH_4ClO_3, ammonium chlorate.

Practice Exercise

Name the following compounds:

(a) PbO, Lead(II) oxide

(b) Li_2SO_3, Lithium sulfite

Molecular Compounds

Unlike ionic compounds, molecular compounds contain discrete molecular units. They are usually composed of nonmetallic elements.

Many molecular compounds are binary compounds.

Naming binary molecular compounds is similar to naming binary ionic compounds. We place the name of the first element in the formula first, and the second element is named by adding -ide to the root of the element name.

e.g.,

	Name
HCl	hydrogen chloride
HBr	hydrogen bromide
SiC	silicon carbide

It is quite common for one pair of elements to form several different compounds. In these cases, confusion in naming the compounds is avoided by the use of Greek prefixes to denote the number of atoms of each element present

e.g.,

CO	carbon monoxide
CO_2	carbon dioxide
SO_2	sulfur dioxide
SO_3	sulfur trioxide
NO_2	nitrogen dioxide
N_2O_4	dinitrogen tetroxide

Greek Prefixes

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

Notes in naming compounds with prefixes:

- The prefix "mono-" may be omitted for the first element.
e.g., PCl_3 is named phosphorus trichloride, not monophosphorus trichloride.

- For oxides, the ending "a" in the prefix is sometimes omitted.
e.g., N_2O_4 may be called dinitrogen tetroxide rather than dinitrogen tetraoxide.

Exceptions to the use of Greek prefixes are molecular compounds containing **hydrogen**. Traditionally, many of these compounds are called either by their common, nonsystematic names or by names that do not specifically indicate the number of H atoms present:

B_2H_6	diborane	CH_4	methane
SiH_4	silane	NH_3	ammonia
PH_3	phosphine	H_2O	water
H_2S	hydrogen sulfide		

EXAMPLE

Name the following molecular compounds:

(a) $SiCl_4$: silicon tetrachloride

(b) P_4O_{10}: tetraphosphorus decoxide

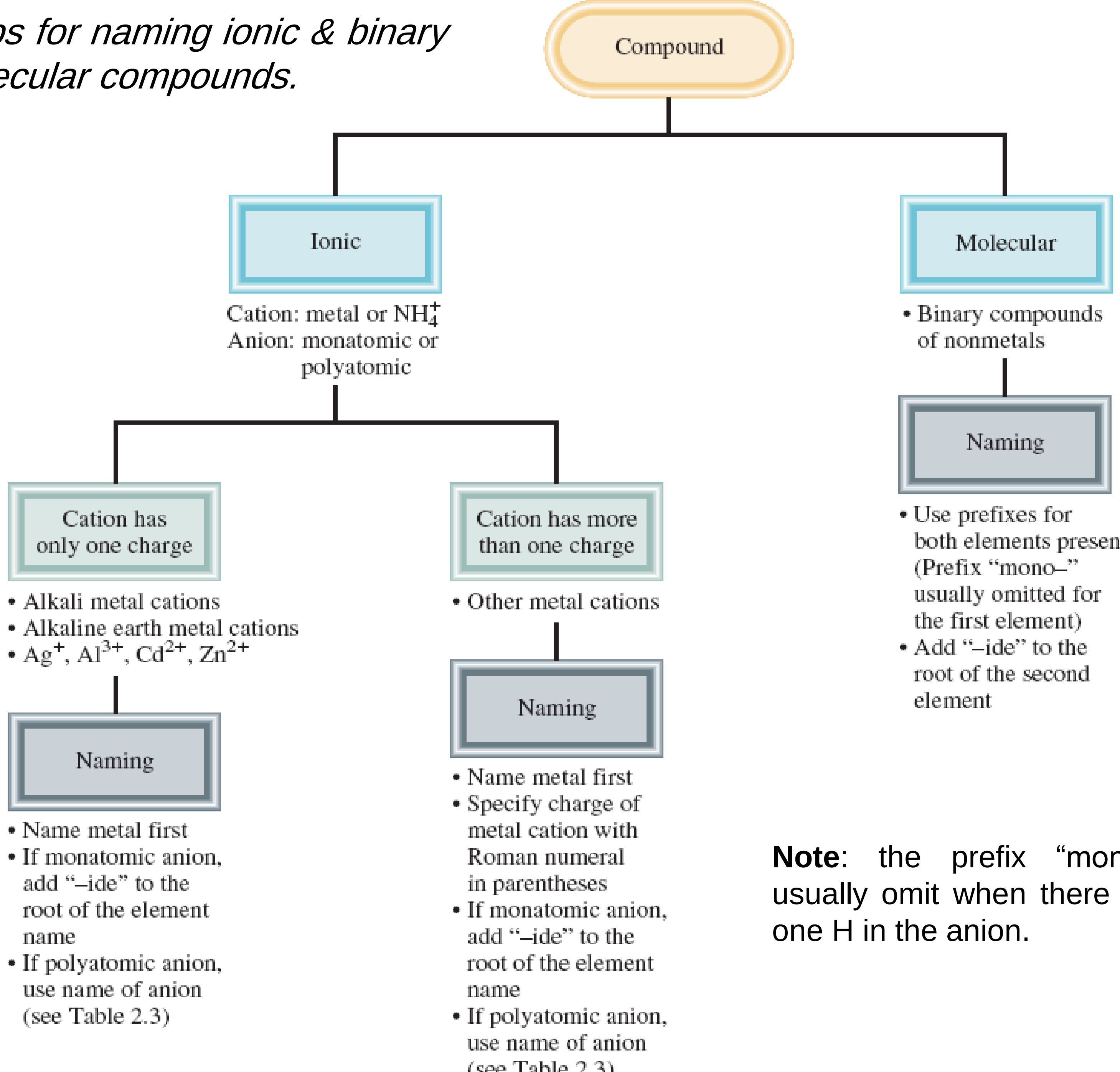
EXAMPLE

Write chemical formulas for the following molecular compounds:

(a) carbon disulfide: CS_2

(b) disilicon hexabromide: Si_2Br_6

Steps for naming ionic & binary molecular compounds.



Note: the prefix "mono-" is usually omit when there is only one H in the anion.

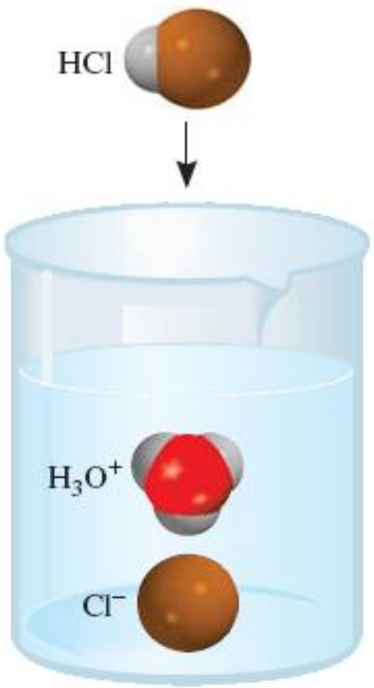
Acids

An acid can be described as a substance that yields hydrogen ions (H^+) when dissolved in water (H^+ is equivalent to one proton).

Formulas for acids contain one or more hydrogen atoms as well as an anionic group.

Anions whose names end in "-ide" form acids with a "hydro-" prefix and an "-ic" ending. In some cases two different names seem to be assigned to the same chemical formula.

HCl: hydrogen chloride, or
HCl: hydrochloric acid



The name assigned to the compound depends on its physical state. In the gaseous or pure liquid state, HCl is a molecular compound called hydrogen chloride. When it is dissolved in water (become solution), the molecules break up into H^+ and Cl^- ions; in this state, the substance is called hydrochloric acid.

Examples of some simple acids

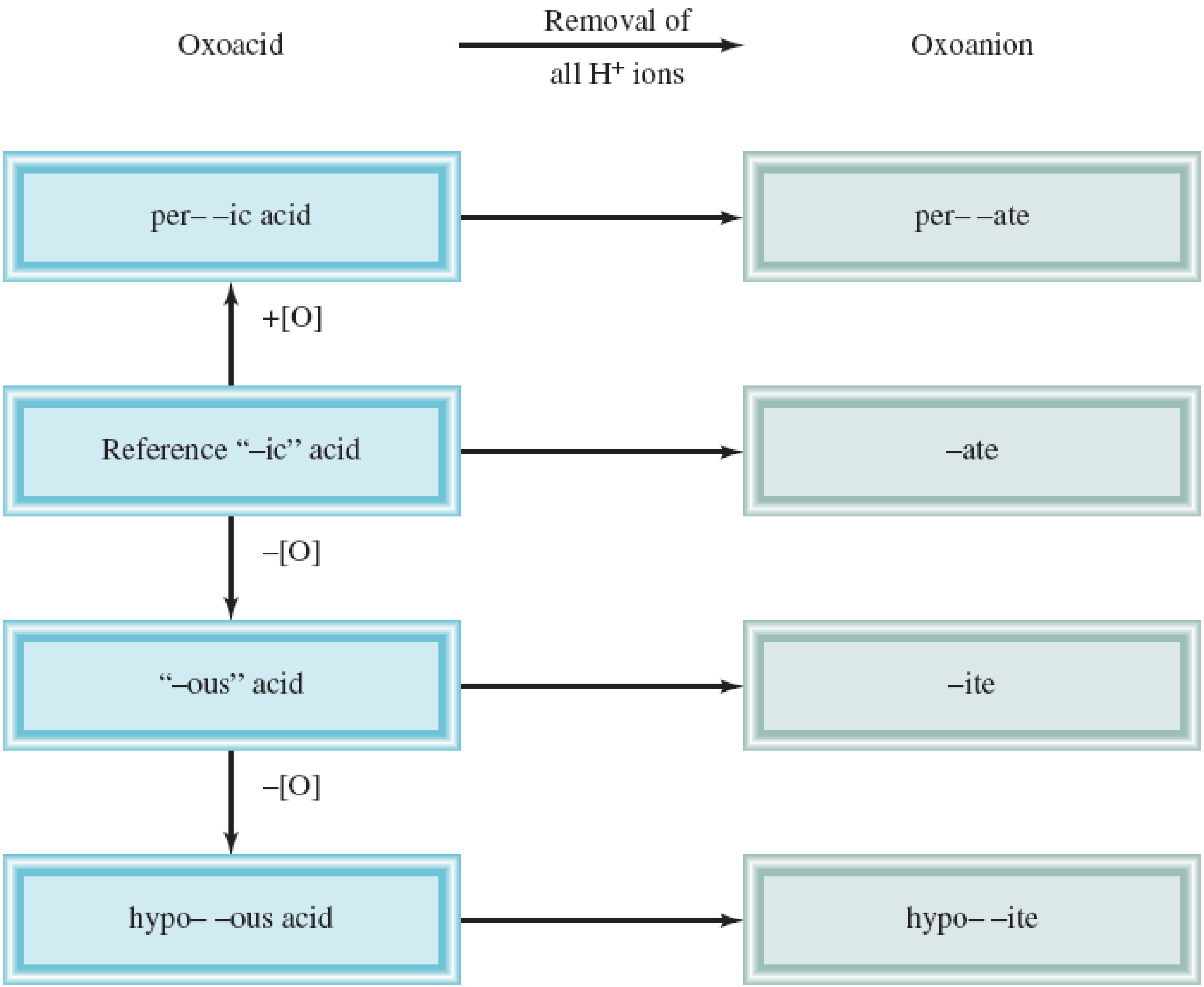
Anion	Corresponding Acid
F^- (fluoride)	HF (hydrofluoric acid)
Cl^- (chloride)	HCl (hydrochloric acid)
Br^- (bromide)	HBr (hydrobromic acid)
I^- (iodide)	HI (hydroiodic acid)
CN^- (cyanide)	HCN (hydrocyanic acid)
S^{2-} (sulfide)	H_2S (hydrosulfuric acid)

Oxoacids are acids that contain hydrogen, oxygen and another element (the central element). The formulas of oxoacids are usually written with the H first, followed by the central element and then O.

Oxoanions are the anions of oxoacids (oxoacid – H).

We use the following five common acids as references in naming oxoacids:

H_2CO_3	carbonic acid
$HClO_3$	chloric acid
HNO_3	nitric acid
H_3PO_4	phosphoric acid
H_2SO_4	sulfuric acid



Naming oxoacids & oxoanions

Examples (oxoacids)	
$HClO_3$ (chloric acid) + O	$HClO_4$ (perchloric acid)
HNO_3 (nitric acid) – O	HNO_2 (nitrous acid)
$HBrO_3$ (bromic acid) – 2 O	HBrO (hypobromous acid)

Examples (oxoanions)	
H_2CO_3 (carbonic acid) – all H	CO_3^{2-} (carbonate)
$HClO_2$ (chlorous acid) – all H's	ClO_2^- (chlorite)

Examples

H_3PO_4: phosphoric acid
$H_2PO_4^-$: dihydrogen phosphate
HPO_4^{2-}: hydrogen phosphate
PO_4^{3-}: phosphate

e.g.,

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

EXAMPLE

Name the following oxoacid and oxoanion:
(a) H_3PO_3
start with the reference acid, phosphoric acid (H_3PO_4)
H_3PO_3 has one fewer O atom, it is called phosphorous acid.

(b) IO_4^-
start with the reference acid, iodic acid (HIO_3)
HIO_4 has one more O atom called periodic acid
Remove H to obtain IO_4^- which called periodate.

Bases

A base can be described as a substance that yields hydroxide ions (OH^-) when dissolved in water.

Examples

NaOH: sodium hydroxide

KOH: potassium hydroxide

$Ba(OH)_2$: barium hydroxide

NH_3: ammonia

$$NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$$

Hydrates

Hydrates: compounds that have a specific number of water molecules attached to them.

The water molecules can be driven off by heating to produce the **anhydrous**; means that the compound no longer has water molecules associated with it.

$CuSO_4 \cdot 5H_2O$ (left) is blue, $CuSO_4$ (right) is white

e.g.,

$CuSO_4 \cdot 5H_2O$	copper(II) sulfate pentahydrate
$BaCl_2 \cdot 2H_2O$	barium chloride dihydrate
$LiCl \cdot H_2O$	lithium chloride monohydrate
$MgSO_4 \cdot 7H_2O$	magnesium sulfate heptahydrate
$Sr(NO_3)_2 \cdot 4H_2O$	strontium nitrate tetrahydrate

Familiar Inorganic Compounds

Common and systematic names of some compounds

Formula	Common Name	Systematic Name
H_2O	Water	Dihydrogen monoxide
NH_3	Ammonia	Trihydrogen nitride
CO_2	Dry ice	Solid carbon dioxide
$NaCl$	Table salt	Sodium chloride
N_2O	Laughing gas	Dinitrogen monoxide
$CaCO_3$	Marble, chalk, limestone	Calcium carbonate
CaO	Quicklime	Calcium oxide
$Ca(OH)_2$	Slaked lime	Calcium hydroxide
$NaHCO_3$	Baking soda	Sodium hydrogen carbonate
$Na_2CO_3 \cdot 10H_2O$	Washing soda	Sodium carbonate decahydrate
$MgSO_4 \cdot 7H_2O$	Epsom salt	Magnesium sulfate heptahydrate
$Mg(OH)_2$	Milk of magnesia	Magnesium hydroxide
$CaSO_4 \cdot 2H_2O$	Gypsum	Calcium sulfate dihydrate


Thank You!