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

Chapter 2
Atoms, Molecules and Ions

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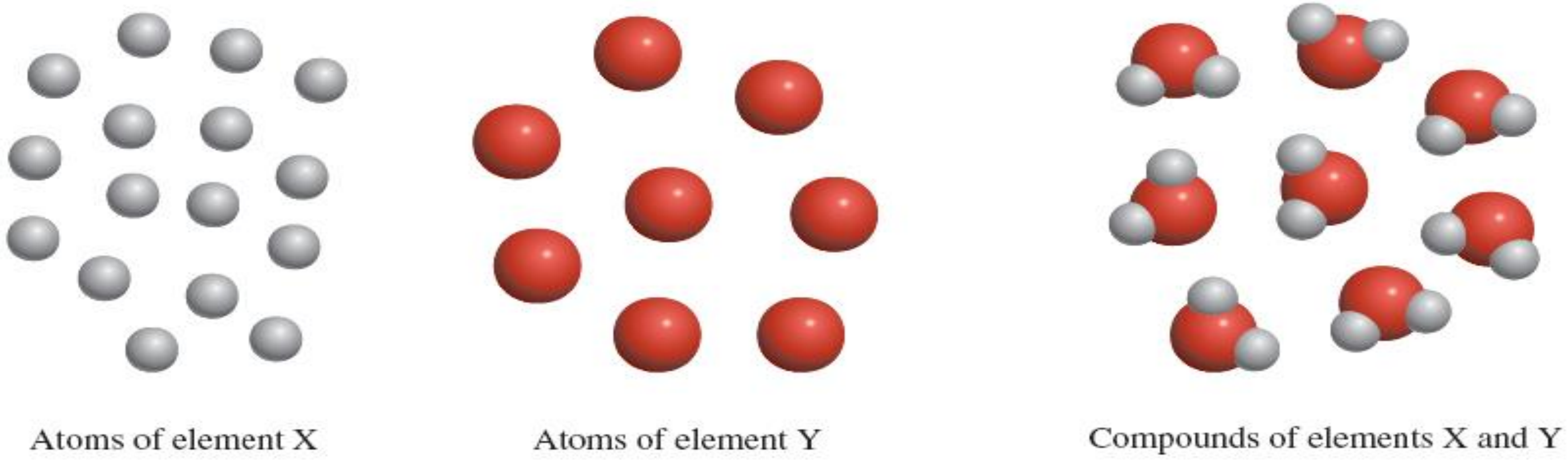
2.2
The Structure of the Atom

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).

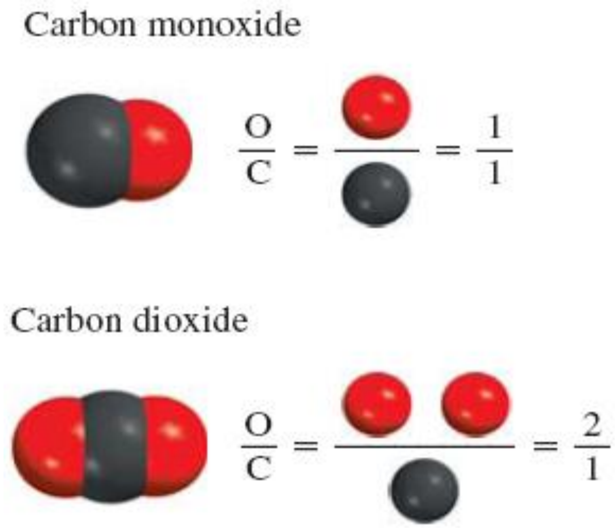


The third hypothesis is an extension of a law published in 1799 by Joseph Proust. Proust's **law of definite proportions** states that different samples of the same compound always contain its constituent elements in the same proportion by mass. Thus, if we were to analyze samples of carbon dioxide gas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

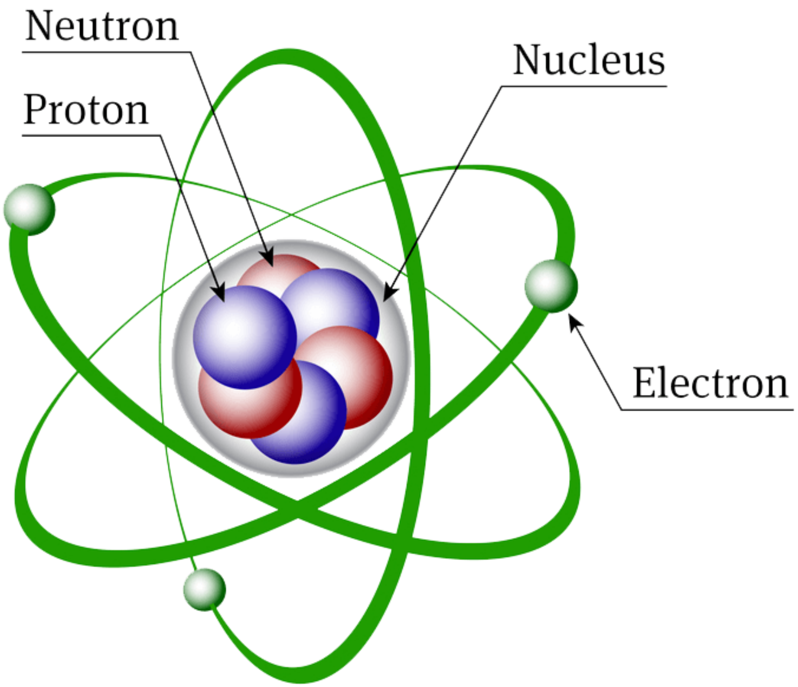


Ratio of oxygen in carbon monoxide to oxygen in carbon dioxide: 1:2

The Structure of the Atom

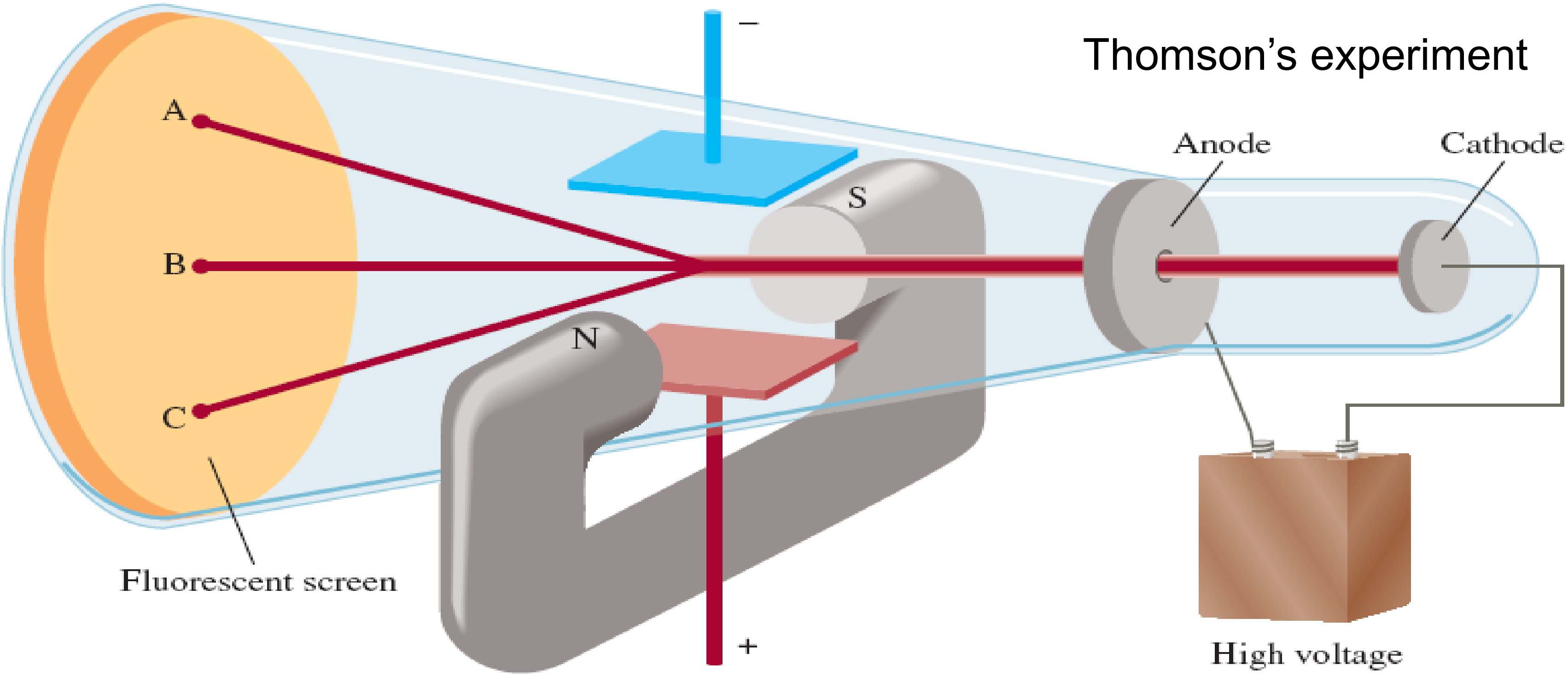
The **atom** is the basic unit of an element that can enter into chemical combination.

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

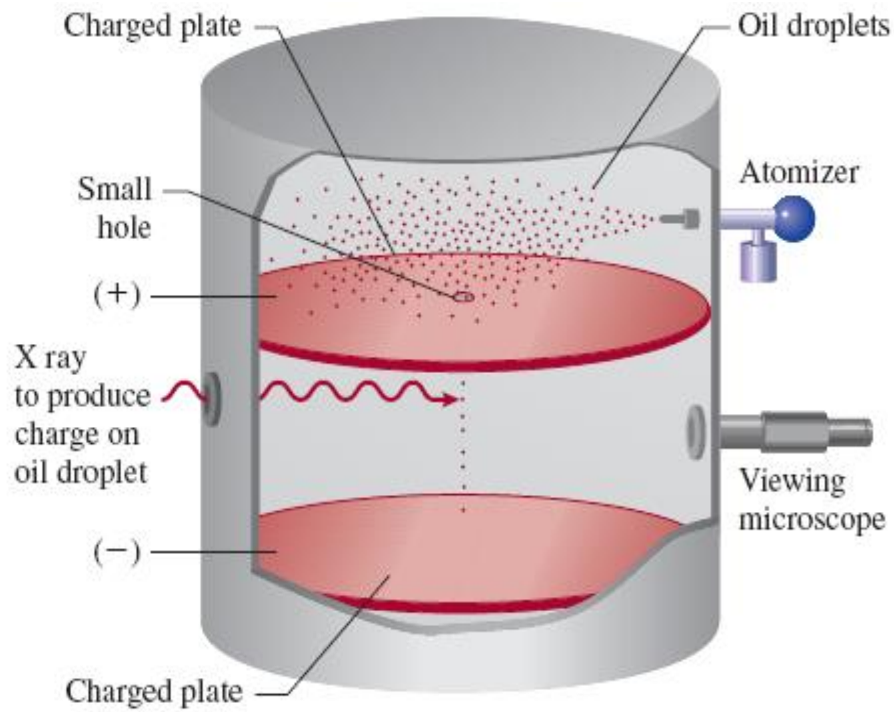
A cathode ray tube with an electric field perpendicular to the direction of the cathode rays and an external magnetic field. The cathode rays will strike the end of the tube at A in the presence of a magnetic field, at C in the presence of an electric field, and at B when there are no external fields 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 **-1.76×10^8 C/g**.

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.



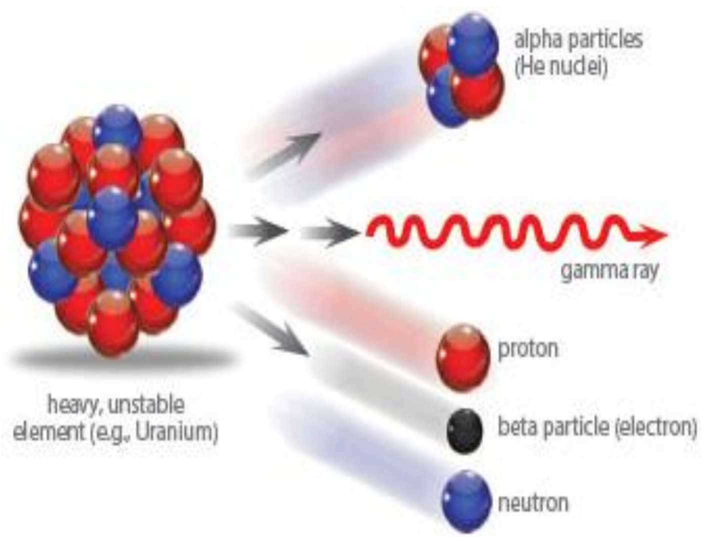
Millikan's experiment

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

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

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.



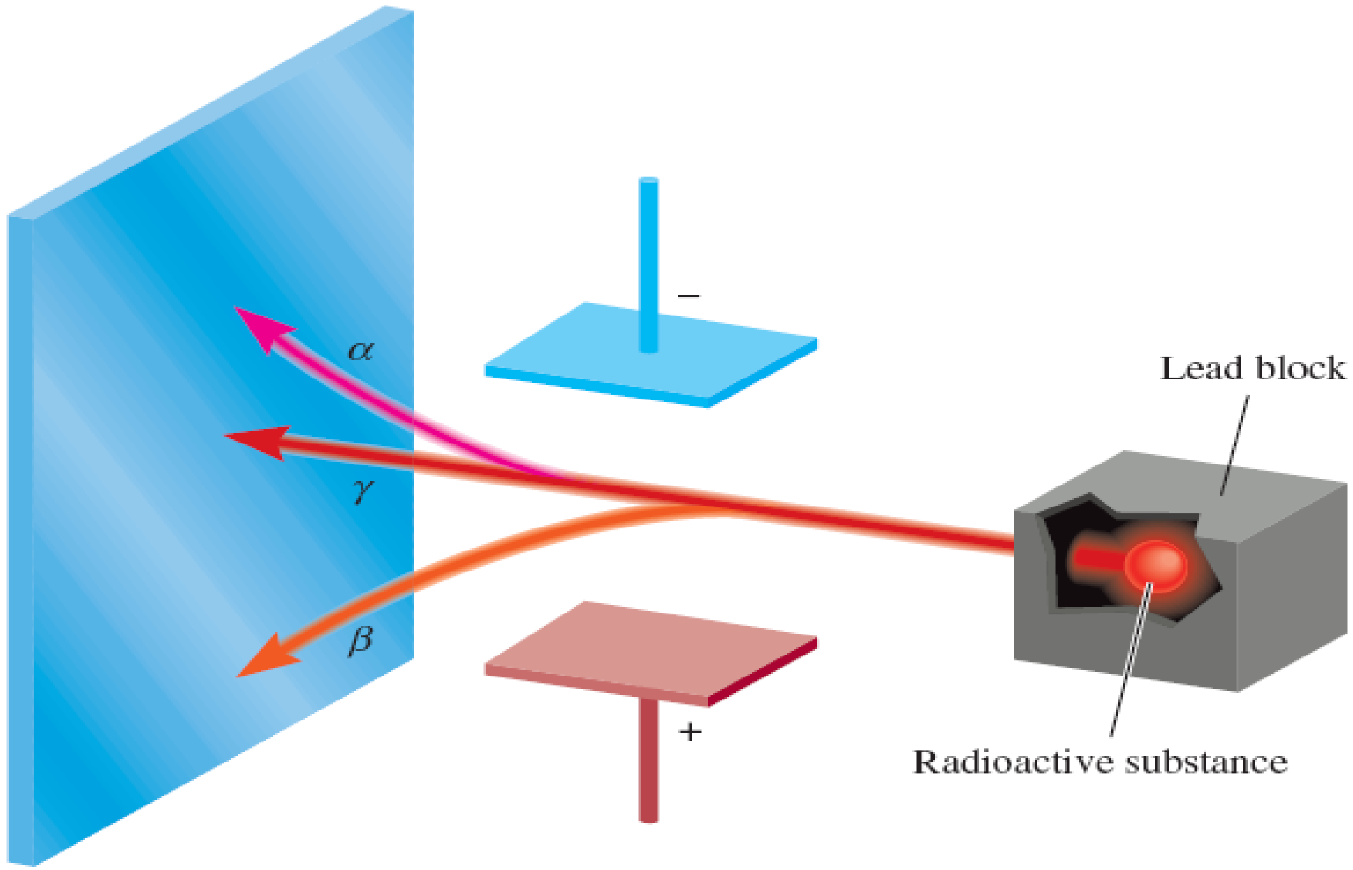
Antoine Becquerel found that exposing thickly wrapped photographic plates to a certain uranium compound caused them to darken, even without the stimulation of cathode rays. Like X-rays, the rays from the uranium compound were highly energetic and could not be deflected by a magnet, but they differed 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.

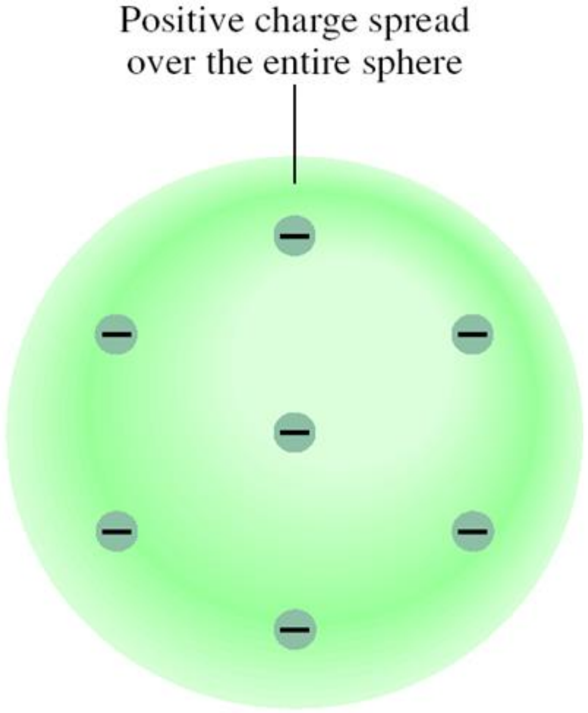
-Gamma (γ) rays, like X-rays, γ rays are high-energy rays and have no charge and are not affected by an external electric field.



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.

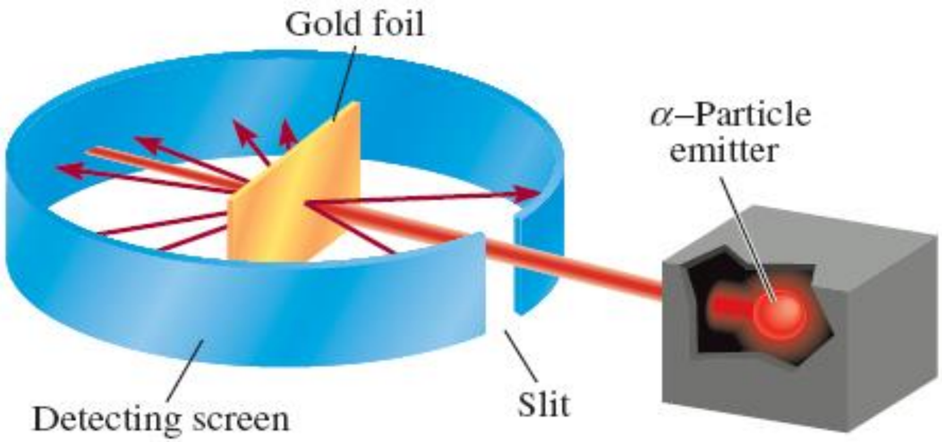
Therefore, Thomson proposed that an atom could be thought of as a uniform, positive sphere of matter in which electrons are embedded like raisins in a cake. This so-called "plum-pudding" model was the accepted theory for a number of years.



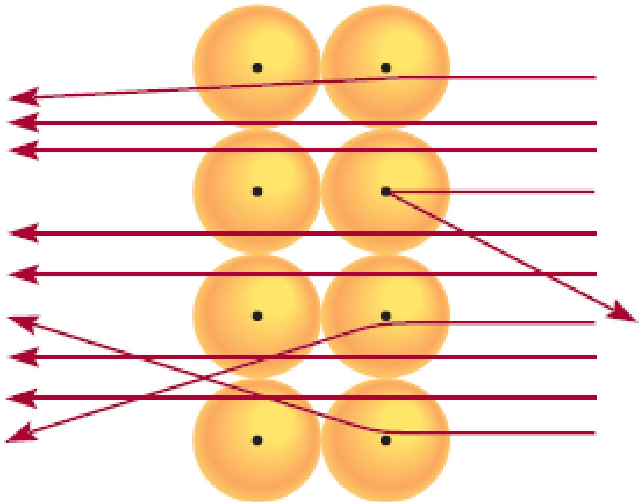
Thomson's model of the atom "plum-pudding" model, after a traditional English dessert containing raisins. The electrons are embedded in a uniform, positively charged sphere.

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

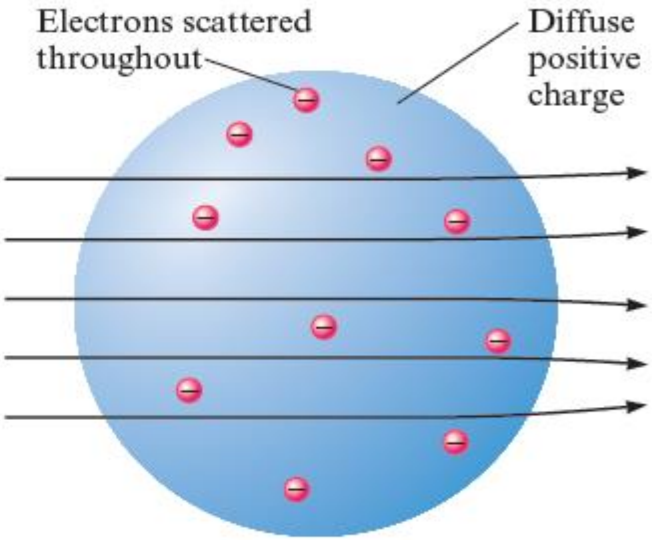


Rutherford's experimental design for measuring the scattering of a particles by a piece of gold foil.

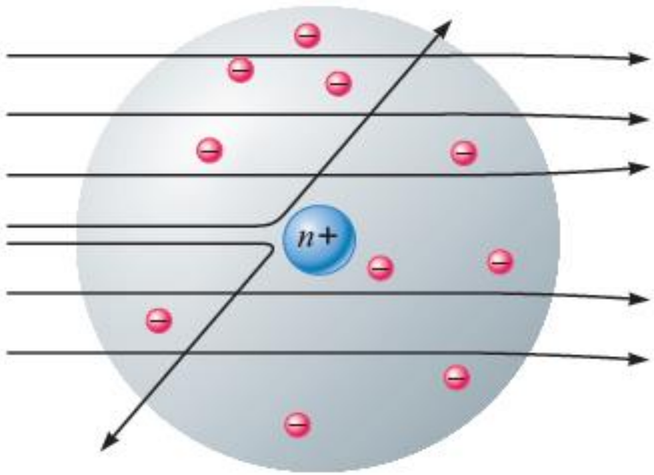


Magnified view of α particles passing through and being deflected by nuclei.

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.

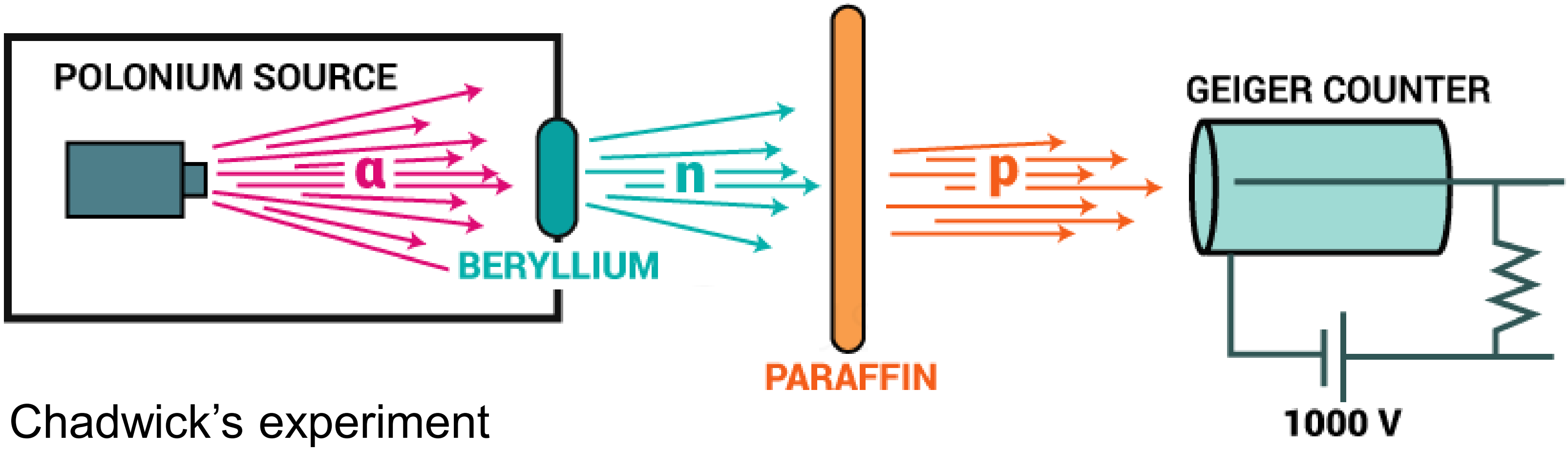
Rutherford's model of the atom

- 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 1.67262×10^{-24} g, about 1840 times the mass of electron.

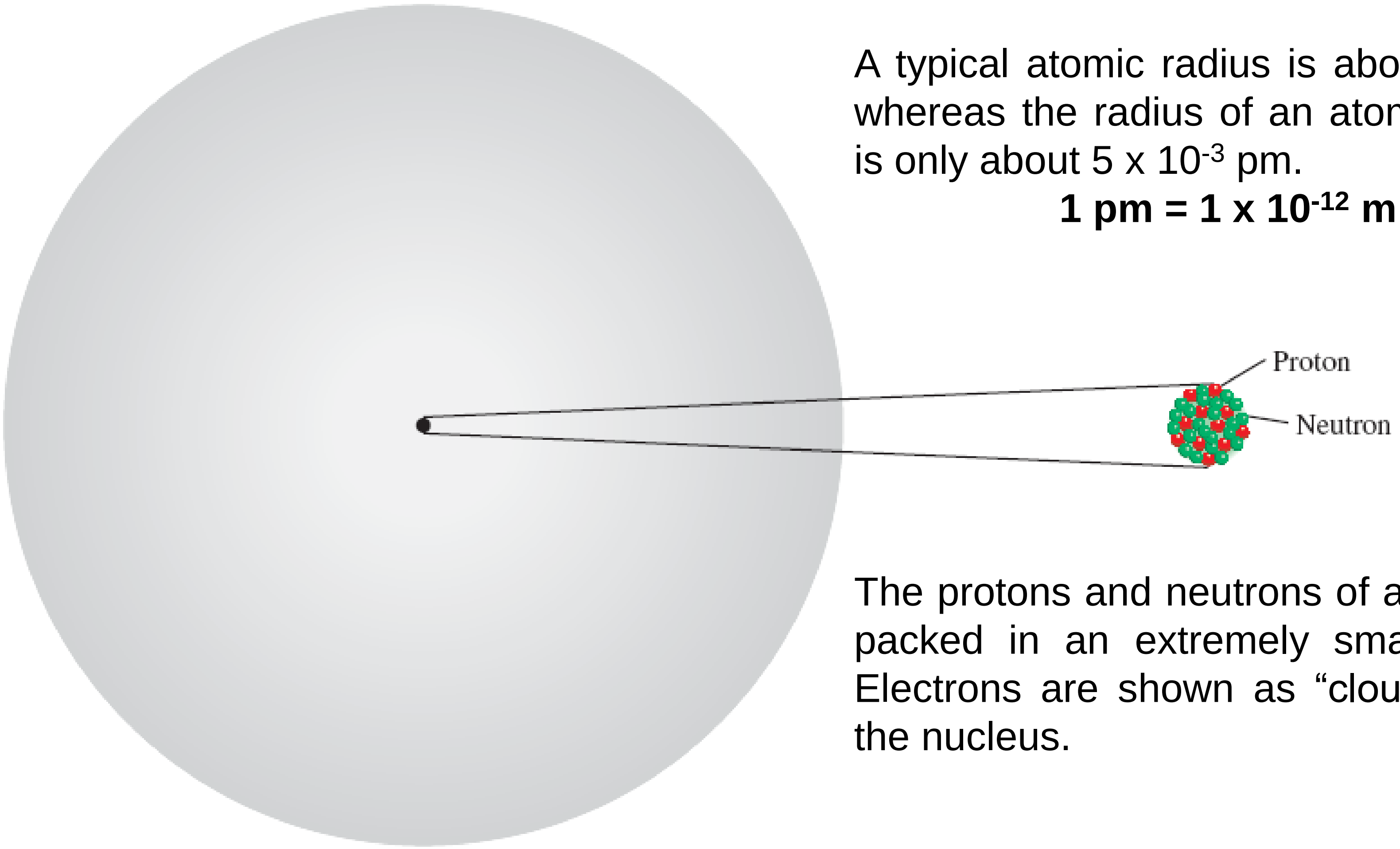
The Neutron

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.



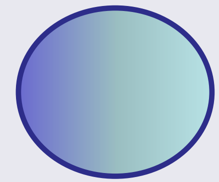
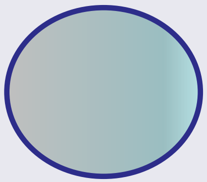
A typical atomic radius is about 100 pm, whereas the radius of an atomic nucleus is only about 5×10^{-3} pm.

1 pm = 1×10^{-12} 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

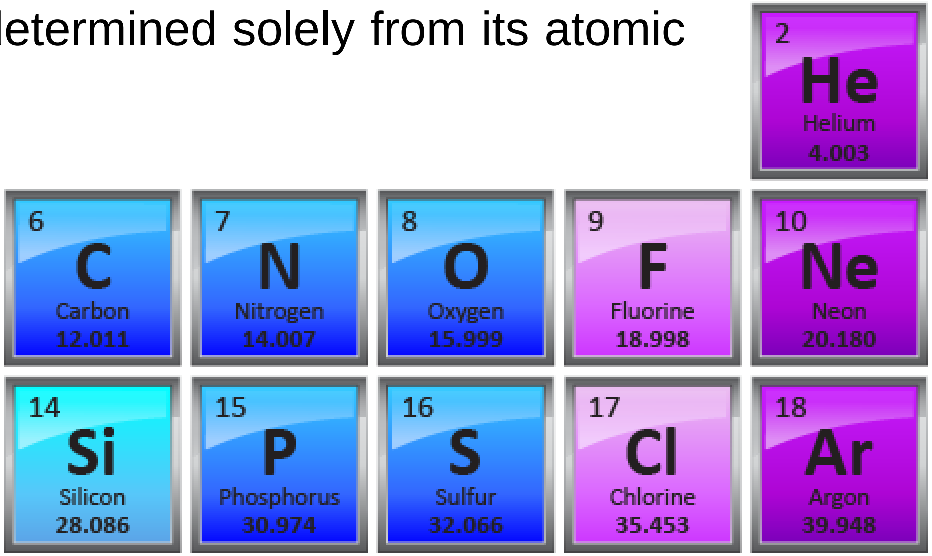
Atomic Number

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

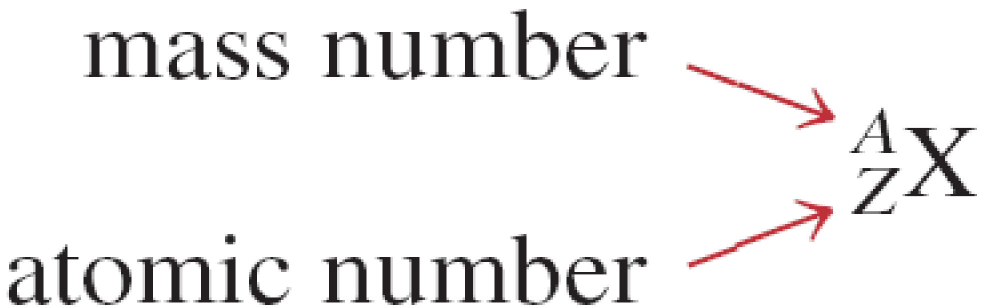
The mass number (A) is the total number of neutrons and protons present in the nucleus of an atom of an element. Except for the most common form of hydrogen, which has one proton and no neutrons, all atomic nuclei contain both protons and neutrons. In general, the mass number is given by

mass number = number of protons + number of neutrons
= atomic number + number of neutrons

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

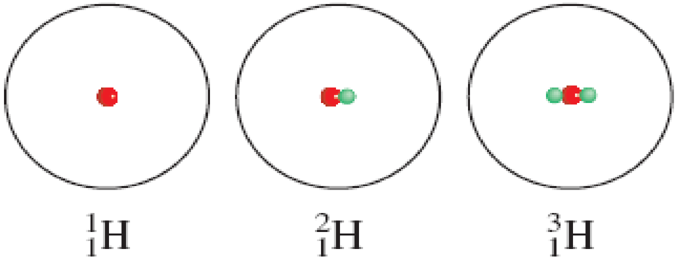
Atoms of a given element do not all have the same mass. Most elements have two or more isotopes, atoms that have the same atomic number but different mass numbers.

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

$$^{1}_{1}\mathrm{H} \qquad ^{2}_{1}\mathrm{H} \qquad ^{3}_{1}\mathrm{H}$$

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

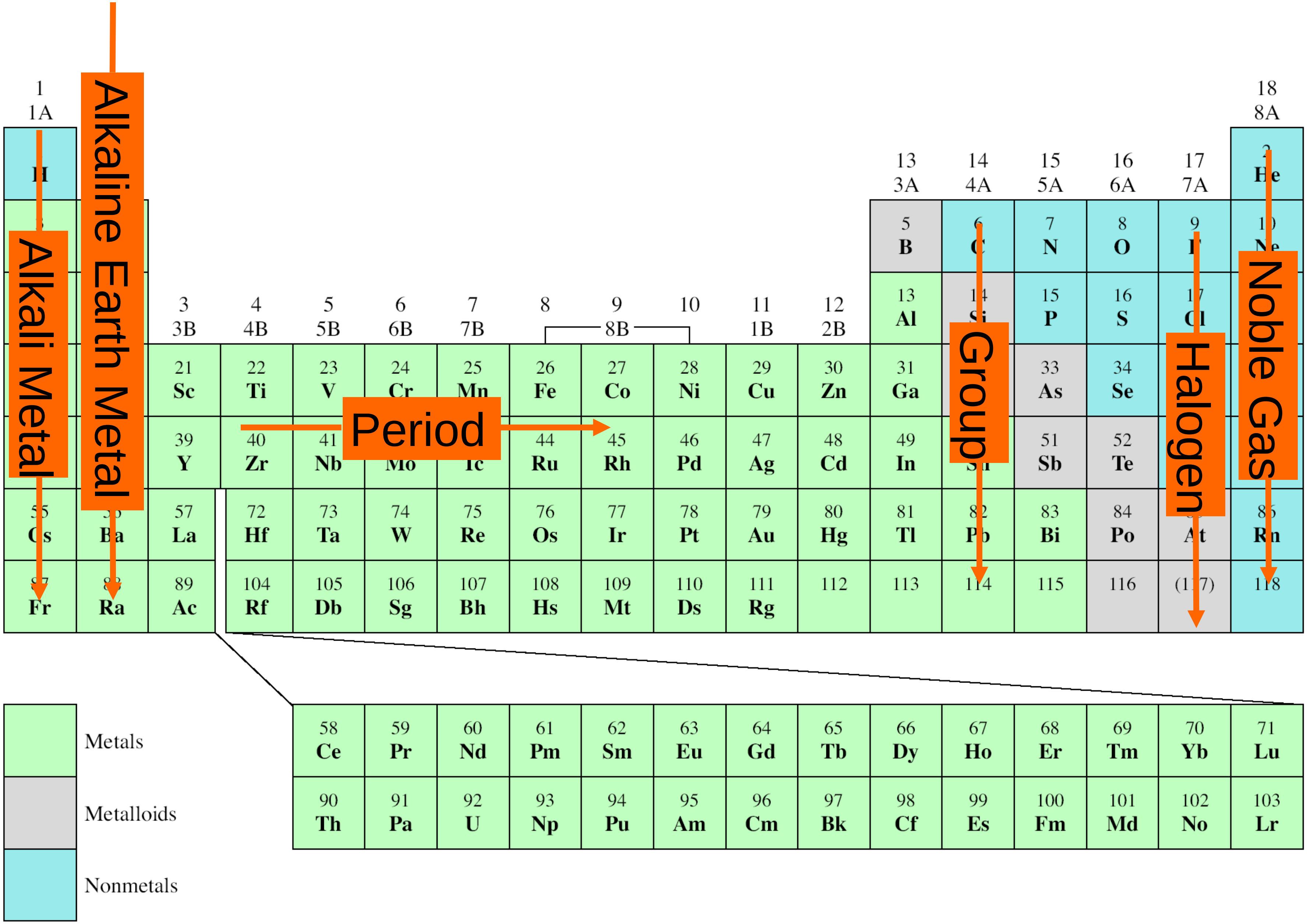
1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 **H**																	2 **He**
3 **Li**	4 **Be**											5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
11 **Na**	12 **Mg**											13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**
19 **K**	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

Atomic #
Symbol
Name
Weight

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

Pnictogens (15) · Chalcogens (16) · Halogens (17)

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H Hydrogen 1.008																	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)

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.


Thank You!