People’s Democratic Republic of Algeria

Ministry of Higher Education and Scientific Research

University of Tlemcen

Faculty of Technology

DEPARTMENT OF BIOMEDICAL ENGINEERING

STRUCTURE OF MATTER COURSE

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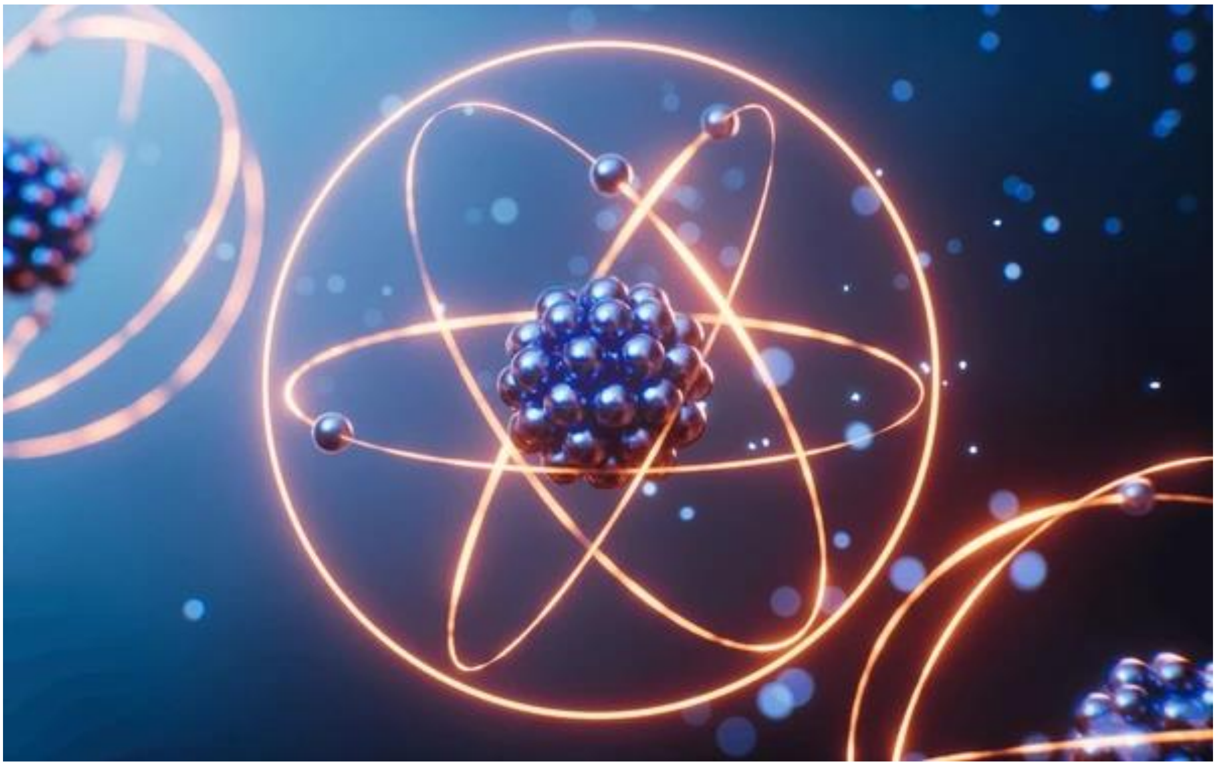


For first-year undergraduate students in Biomedical Engineering,
Medical Sciences, Science and Technology, and Materials Science.

2025 – 2026

Preface

This course handout, titled “Structure of Matter Course”, has been prepared to support first-year undergraduate Biomedical Engineering students in their study of the fundamental concepts of matter. It is intended to serve as a foundational reference throughout their academic studies. It may also be useful for all students enrolled in the common-core Science and Technology programs, whether in Bachelor’s or Engineering tracks, and remains relevant for students in medical and paramedical fields.

Developed in accordance with the new curricular frameworks adopted by these programs, this material provides a clear and structured introduction to the microscopic and atomic nature of matter.

Covering essential concepts such as atomic models, isotopes, radioactivity, chemical species, the periodic table, the periodicity of elemental properties, and fundamental quantitative laws, this course forms an integral part of the fundamental teaching unit. Its objective is to provide students the theoretical tools needed to understand the physical and chemical principles underlying materials and technological applications.

Contents

List of Abbreviations

Abbreviation	Meaning
H, He, C, O…	Chemical symbols of elements
H_2O, CO_2, NaCl…	Chemical formulas
e^-	Electron
p^+	Proton
n^0	Neutron
Na^+, Cl^-	Cation and anion
M^+, X^-	Generic cation / anion
α	Alpha particle (helium nucleus)
β^-	Beta particle (electron emitted in decay)
γ	Gamma ray / photon
Z	Atomic number (number of protons)
A	Mass number (protons + neutrons)
N	Number of neutrons ($N = A - Z$)
c	Speed of light in vacuum ($\approx 3 \times 10^8$ m/s)
e	Elementary charge ($\approx 1.602 \times 10^{-19}$ C)
u or amu	Atomic mass unit
g/mol	Gram per mole (molar mass)
L	Liter
mL	Milliliter
mol	Mole
N_a	Avogadro's number (6.022×10^{23} particles/mol)
atm	Atmosphere (pressure)
Pa	Pascal
°C	Degree Celsius
K	Kelvin
s	Solid
l	Liquid
g	Gas
aq	Aqueous (solution in water)
Δ	Change of
$\rightarrow$	Reaction (produces)
c	Concentration (mol/L)
M	Molarity (mol/L)
P	Pressure
V	Volume
T	Temperature
n	Number of moles

Chapter I. Fundamental Concepts

Chemistry is the science that studies matter, its composition, its structure, its properties, and the transformations it can undergo.

It seeks to understand what the substances around us are made of, how atoms and molecules are organized, and how their interactions explain the phenomena observed in nature and in everyday life (see Figure I.1).

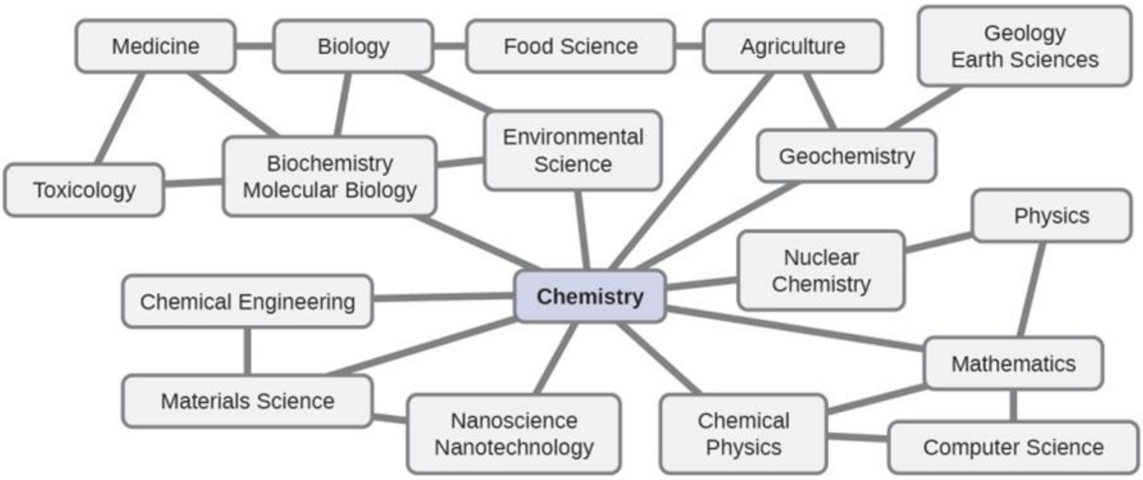


Figure I.1. Chemistry and its interdisciplinary connections

I.1 States of Matter and Their Macroscopic Characteristics

Matter can exist in several macroscopic states: solid, liquid, gas, and sometimes plasma. A solid has its own shape and a fixed volume. Its particles are strongly bound and only vibrate. A liquid has its own volume but no fixed shape: it takes the shape of its container. The particles are close together but can slide past one another. A gas has neither its own shape nor its own volume: it occupies all the space available. Its particles are very far apart and move freely (see Figure I.2). Density is high in solids, medium in liquids, and very low in gases. Compressibility is almost zero for solids, low for liquids, and very high for gases. Fluidity characterizes liquids and gases, but not solids. Changes of state (melting, vaporization, liquefaction, solidification…) modify macroscopic characteristics without changing chemical nature. A plasma is an ionized gas, electrically conductive, found in stars and in certain devices. Macroscopic properties result from the organization and motion of microscopic particles.

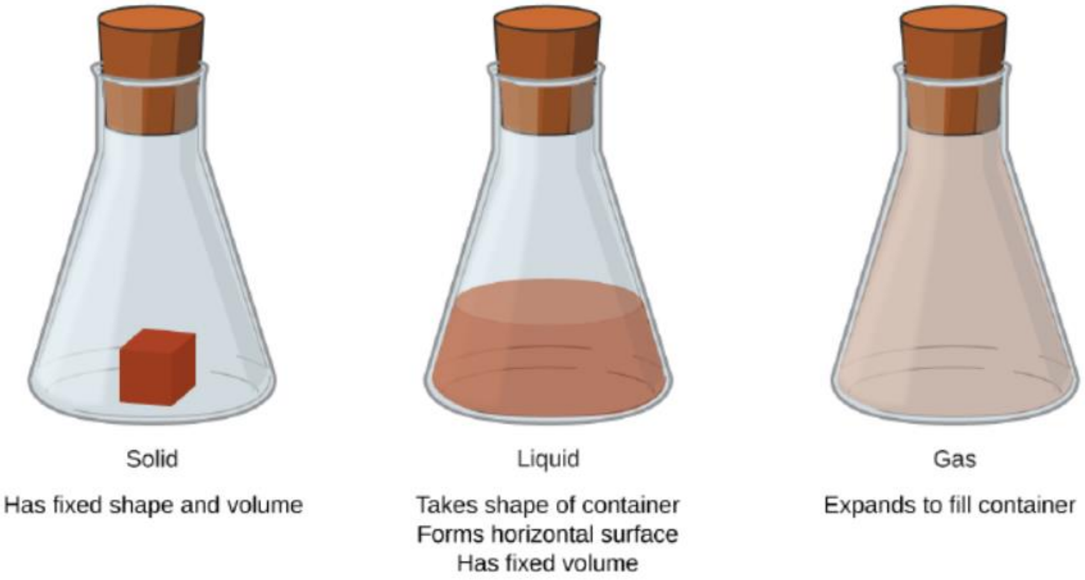


Figure I.2. The three most common states or phases of matter

I.2 Phase Changes of Matter

A change of state is the passage of matter from one physical state to another without changing its chemical nature. Melting: transition from solid to liquid (e.g., ice melting into water). Solidification: transition from liquid to solid (e.g., water turning into ice in a freezer). Vaporization: transition from liquid to gas; it may occur by boiling (boiling water) or evaporation (drying clothes). Liquefaction (or condensation): transition from gas to liquid (e.g., water droplets forming on a cold bottle). Sublimation: direct transition from solid to gas (e.g., snow "disappearing" without melting in dry weather). Solid condensation (or deposition): transition from gas to solid (e.g., frost forming on a window). Each change of state occurs at a specific temperature for pure substances (e.g., water melts at 0 °C). These transformations absorb or release energy (e.g., melting absorbs heat). The mass of matter remains unchanged during a change of state (see Figure I.3).

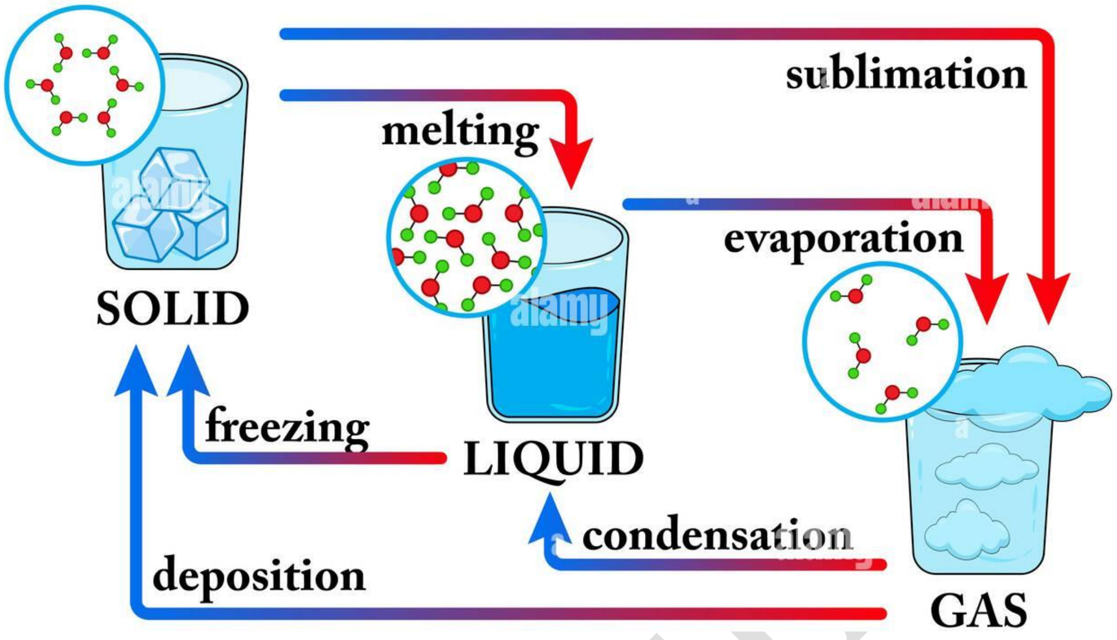


Figure I.3. Phase changes of matter

I.3 Atoms, Molecules, and Ions

An atom is the smallest unit of a chemical element that retains its properties (e.g., an oxygen atom O). An atom is made up of a nucleus (protons + neutrons) and electrons orbiting around it. A molecule is a group of atoms chemically bonded together (e.g., H_2O = 2 hydrogens + 1 oxygen). A molecule may be composed of identical atoms (e.g., O_2) or different atoms (e.g., CO_2). An ion is an atom or a molecule that has gained or lost one or more electrons. An ion that loses electrons becomes positive: it is a cation (e.g., Na^+). An ion that gains electrons becomes negative: it is an anion (e.g., Cl^-). Ions often appear during chemical reactions (e.g., table salt NaCl separates into Na^+ and Cl^- in water). Atoms are neutral, ions are charged, and molecules may be neutral or charged. Atoms, molecules, and ions are the basic units used to describe matter and its chemical transformations (see Figure I.4).

Figure I.5. shows that the elements hydrogen, oxygen, phosphorus, and sulfur from molecules consisting of two or more atoms of the same element. The compounds water, carbon dioxide, and glucose consist of combinations of atoms of different elements.

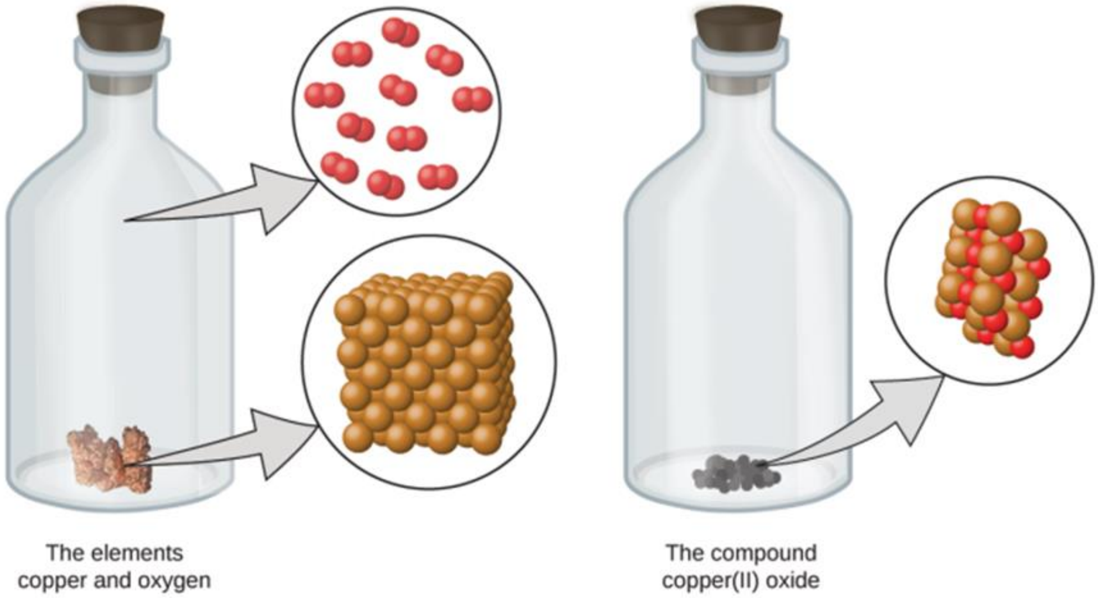


Figure I.4. The elements copper and oxygen and the compound copper(II) oxide

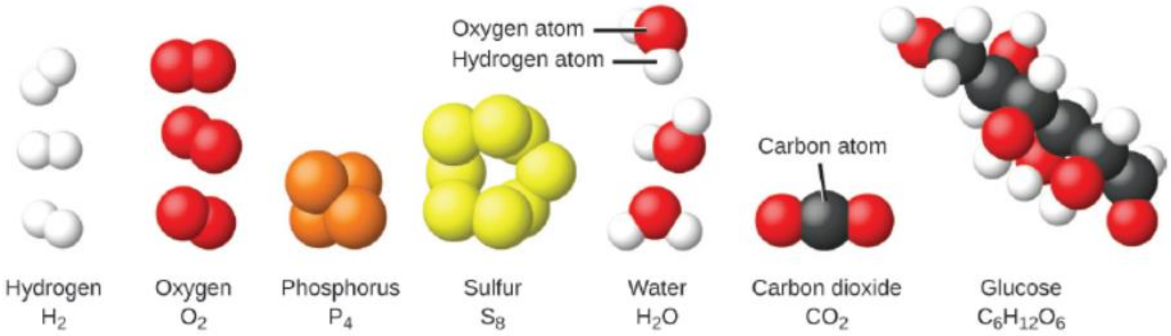


Figure I.5. Elements and compounds

I.4 Mole and Avogadro's Number

A mole is a unit used to count a very large quantity of particles (atoms, molecules, ions…). One mole contains exactly the same number of particles as 12 g of carbon-12. This fixed number is called Avogadro's number. Avogadro's number is equal to 6.022×10^{23} particles per mole. Thus, 1 mole of water contains 6.022 × 10^{23} molecules of H_2O, regardless of the mass amount. Likewise, 1 mole of iron atoms also contains 6.022 × 10^{23} atoms, even though the mass is different. The mole makes it possible to relate mass measured in grams to the number of particles. For example, 18 g of water correspond to 1 mole, because the molar mass of H_2O is 18 g/mol. Similarly, 58.5 g of NaCl correspond to 1 mole of salt (molar mass 58.5 g/mol). Thanks to the mole and Avogadro's number, chemical reactions can be quantified precisely (see figure I.6).

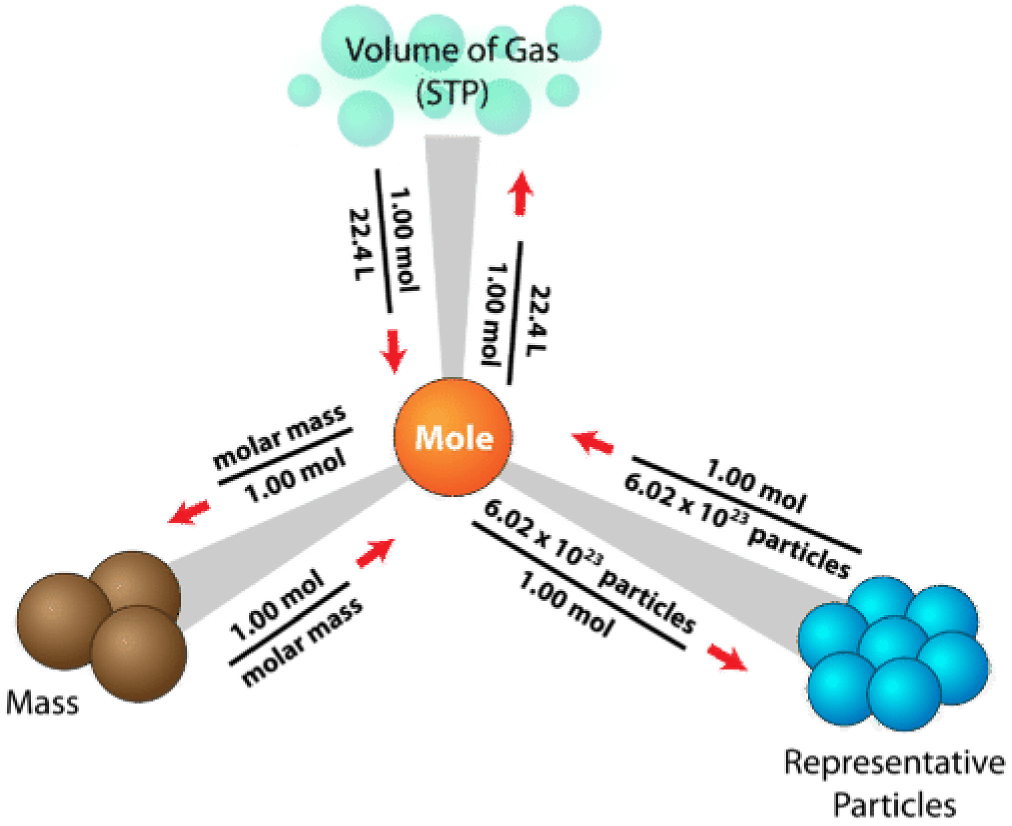


Figure I.6. Mole and Avogadro's number

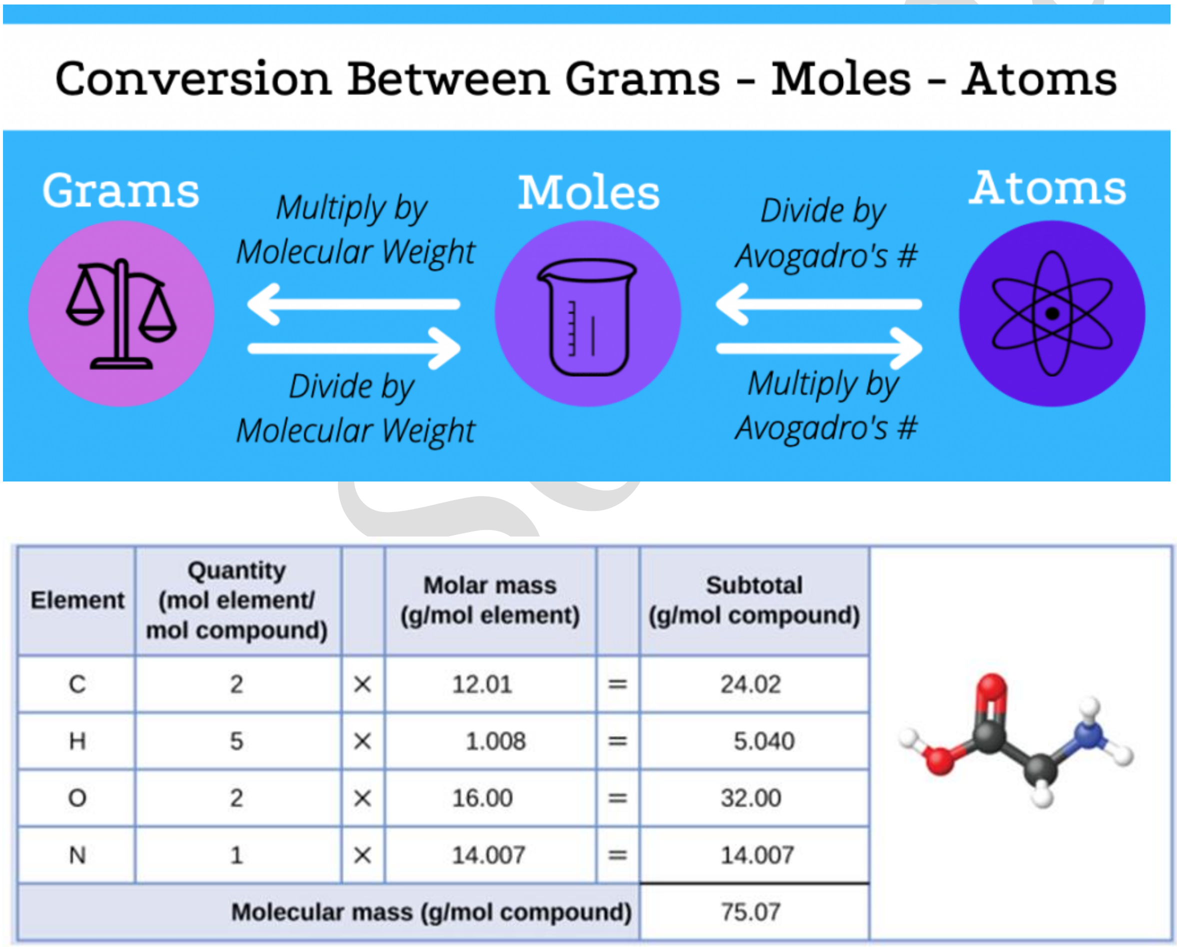
I.5 Atomic Mass Unit, Atomic and Molecular Molar Mass

The atomic mass unit (u or amu) is a unit used to express the mass of atoms, defined as 1/12 of the mass of a carbon-12 atom. For example, a hydrogen atom has a mass of about 1 u, while an oxygen atom has a mass of about 16 u. The atomic molar mass is the mass of one mole of atoms of an element, expressed in g/mol, and its numerical value is the same as the atomic mass in u (e.g., oxygen has an atomic molar mass of 16 g/mol). Thus, 1 mole of carbon atoms weighs 12 g because the atomic mass of carbon is 12 u. The molecular molar mass is the mass of one mole of molecules, also expressed in g/mol, and is obtained by adding the atomic molar masses of the atoms that make up the molecule (see Figure I.7).

For example: the molar mass of H_2O is (2 × 1 g/mol) + (16 g/mol) = 18 g/mol, and

the molar mass of CO_2 is 12 g/mol + (2 × 16 g/mol) = 44 g/mol.

These quantities make it possible to relate measured laboratory masses to the number of particles involved in a chemical reaction.

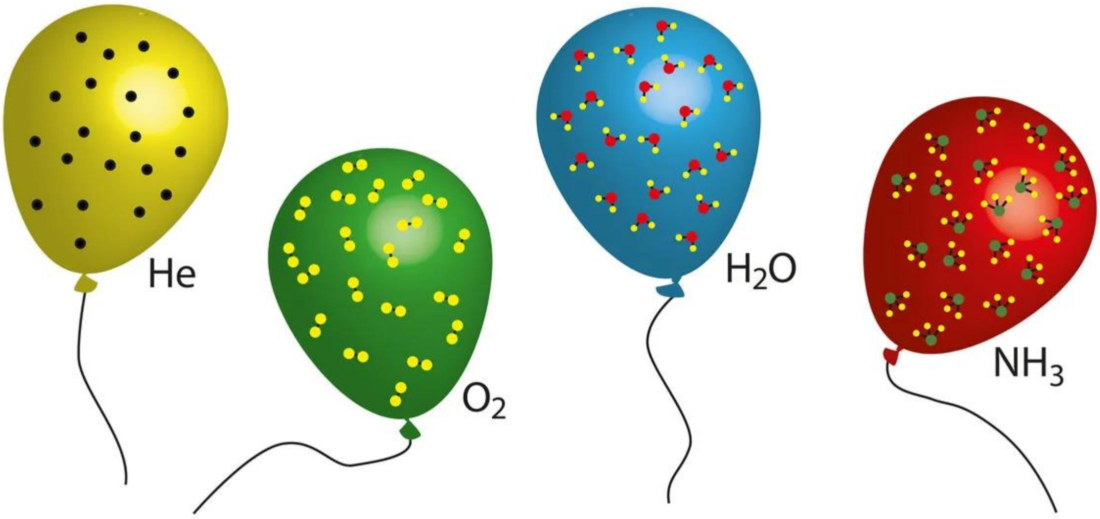


Element	Quantity (mol element/ mol compound)		Molar mass (g/mol element)		Subtotal (g/mol compound)
C	2	×	12.01	=	24.02
H	5	×	1.008	=	5.040
O	2	×	16.00	=	32.00
N	1	×	14.007	=	14.007
Molecular mass (g/mol compound)					75.07

Figure I.7. Molar mass

I.6 Molar Volume

The molar volume is the volume occupied by one mole of a gas under given conditions of temperature and pressure. At standard temperature and pressure (STP: 0 °C and 1 atm), the molar volume of an ideal gas is 22.4 L/mol, meaning that one mole of any gas (O_2, N_2, CO_2, etc.) occupies 22.4 L under these conditions. For example, 1 mole of dioxygen (O_2) has a volume of 22.4 L, while 0.5 mole of CO_2 occupies $0.5 \times 22.4 = 11.2$ L. Molar volume depends on temperature— increasing as temperature increases— and also on pressure, decreasing when pressure increases. At 25 °C and 1 atm, the molar volume is approximately 24.0 L/mol. This concept is particularly useful for relating the number of moles to the measured volume of a gas; thus, measuring a gas volume in liters makes it possible to determine the amount of substance it contains (see Figure I.8).



Avogadro's hypothesis: same volumes of gases present same number of particles

Figure I.8. Molar volume

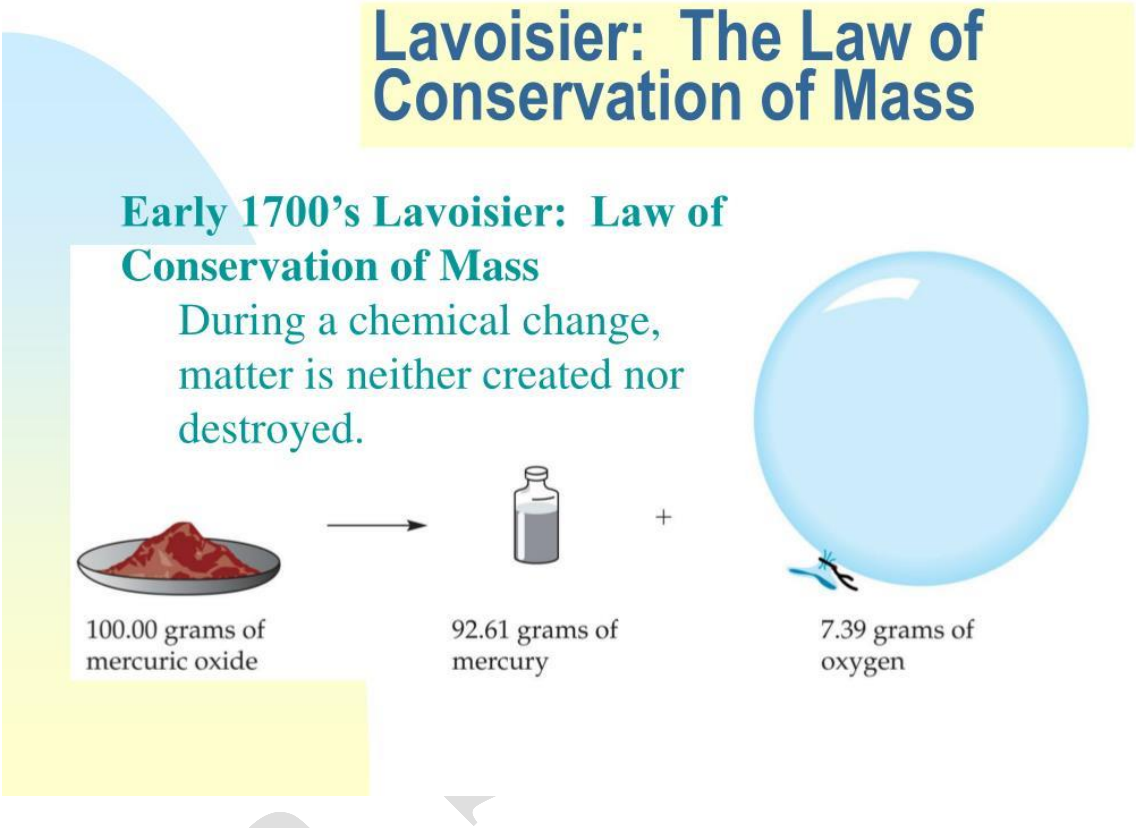
I.7 Law of Mass: Conservation of Mass (Lavoisier)

The law of conservation of mass (Lavoisier, 1789) states that mass is conserved during a chemical reaction (see Figure I.9). This means that the mass of the reactants is equal to the mass of the products formed. In other words: nothing is lost, nothing is created, everything is transformed. This law applies as long as the system is closed (nothing enters, nothing leaves). Example: if copper is heated with sulfur

$$Cu + S \rightarrow CuS$$

the total mass before and after remains the same. Simple example: burning paper in a closed system does not change the total mass (the CO_2 gas remains trapped). If the mass seems to decrease in air, it is because gases escape (the system is not closed). Conservation of mass allows chemical equations to be balanced: the same number of atoms on each side.

During a change of state (ice → water → vapor), the nature changes, but the mass remains identical. Lavoisier’s law is fundamental for all chemical reactions and material balances.



Lavoisier: The Law of Conservation of Mass

Early 1700’s Lavoisier: Law of Conservation of Mass

During a chemical change, matter is neither created nor destroyed.

Figure I.9. Law of conservation of mass (Lavoisier)

I.8 Chemical Reactions

A chemical reaction is a transformation in which reactants are converted into products. It involves a reorganization of atoms: bonds break and new bonds form. During a chemical reaction, atoms are conserved: no atom appears or disappears. Consequently, the total mass remains constant (application of Lavoisier’s law).

Example: $H_2 + O_2 \rightarrow H_2O$

hydrogen and oxygen atoms recombine, but their number remains the same. A chemical equation must be balanced to respect the conservation of atoms.

Balanced example: $2\ H_2 + O_2 \rightarrow 2\ H_2O$ (4 H and 2 O on each side).

Reactants are the substances consumed, products are the substances formed.

A reaction can release energy (exothermic, e.g., combustion) or absorb energy (endothermic).

A chemical reaction changes the nature of substances, unlike a simple change of state (see Figure 1.10).

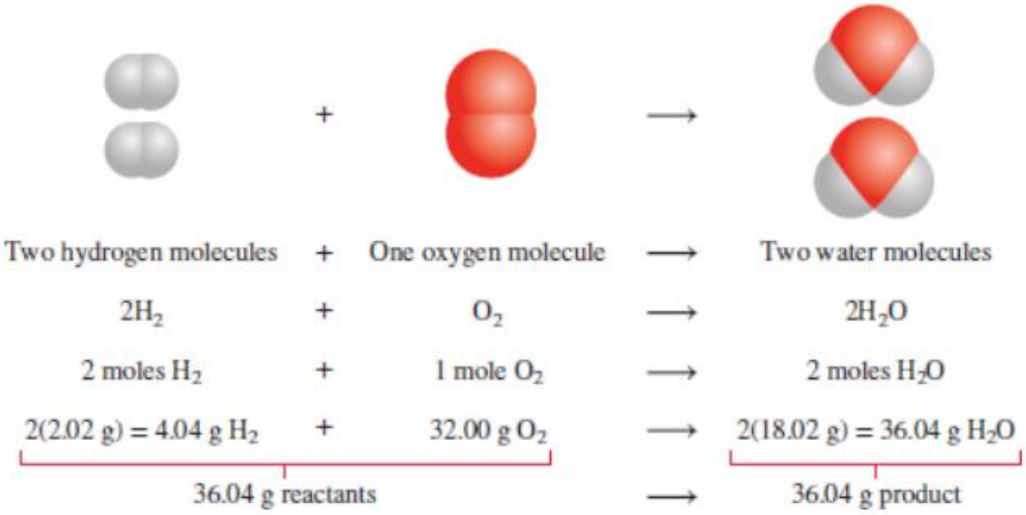


Figure I.10. Three methods for representing hydrogen combustion. According to the law of conservation of mass, the quantity of each type of atom must remain the same on both sides of a chemical equation.

Key points to remember:

- The qualitative aspect of matter concerns the nature of the substances present in a chemical transformation.
- It focuses on what happens to the reactants: which new substances are formed.
- In a chemical reaction, atoms change their associations, but their nature remains the same.
- Example: hydrogen (H_2) and oxygen (O_2) react to form a new substance, water (H_2O).
- Substances are identified qualitatively by their properties: color, odor, solubility, pH, etc.
- Example: iron (gray) reacts with sulfur (yellow) to form iron sulfide (black), a different substance.
- The qualitative aspect shows that the products have new properties, different from those of the reactants.

- It helps determine which chemical species participate in the reaction.
- Unlike the quantitative aspect, it does not focus on masses but on chemical nature.
- The qualitative aspect is essential for recognizing reactions and identifying the new substances formed.
- The quantitative aspect of matter concerns the amounts of reactants and products in a chemical reaction (see Figure I.10).
- It is based on the conservation of mass: total mass of reactants = total mass of products.
- It allows measuring how much matter reacts and how much product is formed.
- Example: if 10 g of hydrogen react with 80 g of oxygen, 90 g of water is obtained.
- The quantitative aspect relies on masses, moles, and stoichiometric relationships (coefficients in the equation).
- Example: $2\ H_2 + O_2 \rightarrow 2\ H_2O$ means that 2 moles of H_2 react with 1 mole of O_2.
- It allows predicting the required amounts: we calculate if a reactant is in excess or limiting.
- Example: to produce 36 g of water, exactly 4 g of H_2 and 32 g of O_2 are needed.
- The quantitative aspect ensures that equations are balanced to respect the conservation of atoms.
- It is essential for measuring, predicting, and controlling the proportions used in chemistry and industry.

Chapter II. Main Constituents of Matter

Understanding the fundamental constituents of matter begins with the discovery that all substances are made of extremely small particles. Through key experiments in the 19th and early 20th centuries, scientists revealed that atoms are not indivisible but contain smaller components; electrons, protons, and neutrons; each with specific properties and roles. From Faraday's electrochemical observations to Thomson's identification of the electron and Rutherford's nuclear model, these discoveries transformed our view of matter and laid the groundwork for modern chemistry and physics. This chapter explores these major milestones and the evidence that established the particulate structure of the atom.

II.1 Faraday's Experiment: Relationship between Matter and Electricity

Faraday demonstrated that electricity can induce chemical transformations through electrolysis. Electrolysis consists of passing an electric current through a solution to decompose a substance. Faraday discovered that the amount of matter released at the electrodes depends directly on the amount of electricity supplied. The stronger the current or the longer it flows, the greater the amount of matter formed. Example: the electrolysis of water produces dioxygen (O_2) and dihydrogen (H_2) when a current is passed. Faraday established that the electric charge carried by 1 mole of electrons is a constant (see eq. [II.1]). He showed that chemical reactions involve electron exchanges, linking matter and electricity.

$$\mathrm{F} = \mathrm{N_A} \times \mathrm{e} \approx 96\,485\ \mathrm{C.mol^{-1}} \qquad [\mathrm{II.1}]$$

Example: depositing metallic copper on an electrode requires a precise amount of electricity. His experiments proved that ions carry charges and move under the effect of a current. Faraday thus laid the foundation of modern electrochemistry, showing that matter obeys measurable and predictable electrical laws.

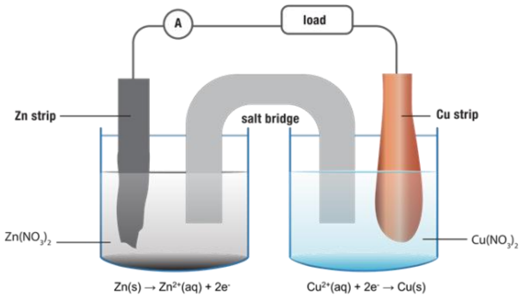


Figure II.1. Faraday's law

II.2 Evidence of the Constituents of the Atom

The atom consists of a central nucleus surrounded by electrons. In 1897, Thomson discovered the electron using a cathode-ray tube, demonstrating that the atom contains negatively charged particles (see Figure II.2). He then proposed the “plum pudding” model (see Figure II.3), in which electrons are embedded within a positively charged sphere. In 1909, Millikan measured the electron’s charge through his oil-drop experiment, providing further confirmation of its existence.

In 1911, Rutherford bombarded a thin gold foil with alpha particles to investigate atomic structure. He observed that most particles passed straight through the foil, while some were strongly deflected. This revealed that the atom’s positive charge and most of its mass are concentrated in a very small nucleus. Rutherford subsequently proposed the nuclear model of the atom, consisting of a positive nucleus with electrons orbiting around it.

Later, in 1932, Chadwick discovered the neutron, a neutral particle also located in the nucleus. Altogether, these experiments showed that the atom is mostly empty space, made up of a dense central nucleus surrounded by light electrons.

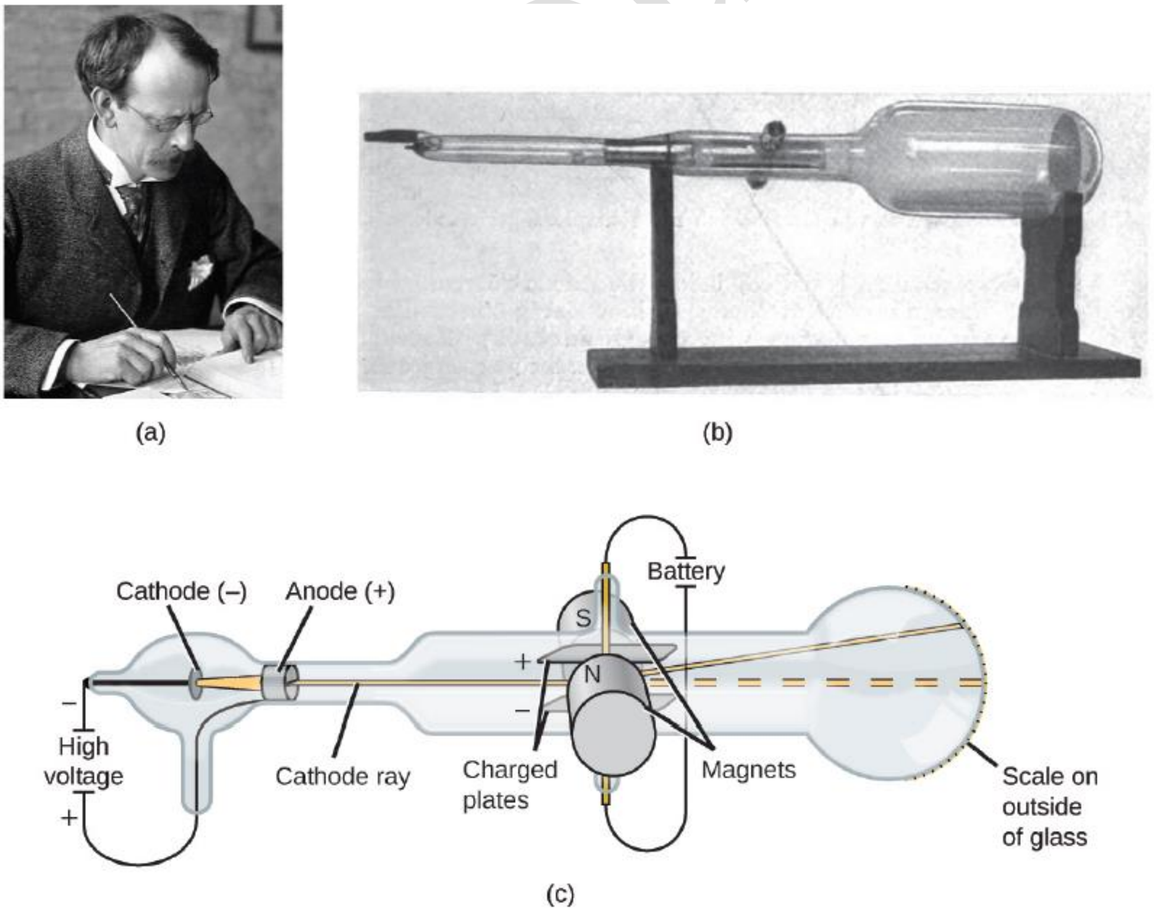


Figure II.2. J. J. Thomson’s experiment

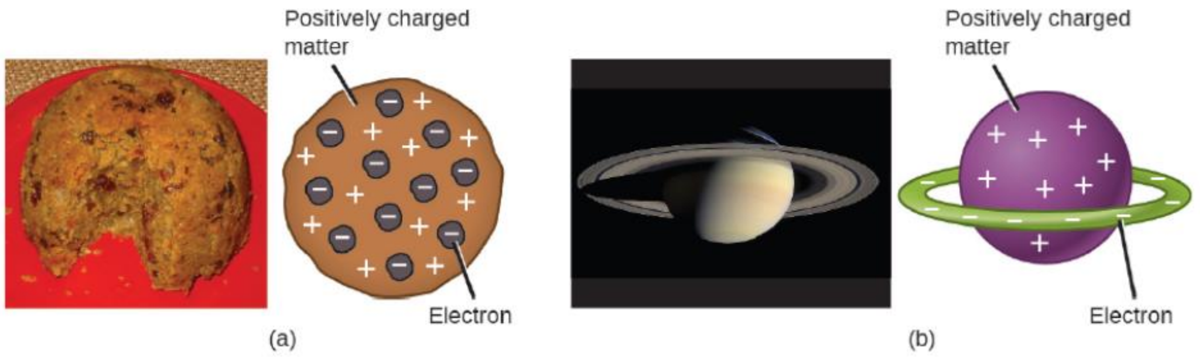


Figure II.3. a) “Plum pudding” model, b) the nuclear model of the atom

II.3 Rutherford’s Planetary Model

Rutherford proposed a model in which the atom contains a very small, dense, and positively charged nucleus at its center. Around this nucleus, negative electrons orbit like planets around the Sun, giving rise to the so-called planetary model. The nucleus holds almost all of the atom’s mass, as it contains both protons and neutrons, while the very light electrons occupy a much larger region of space surrounding it. As a result, the atom is mostly empty space.

This model successfully explains the observations from the gold foil experiment, particularly the strong deflection of some α particles. For example, in an oxygen atom, the eight protons concentrated in the nucleus attract the eight electrons that orbit around it. Rutherford’s model thus introduced the idea of a nuclear structure for the atom. However, it faced a major limitation: according to classical physics, electrons in orbit should continuously lose energy and spiral into the nucleus. This issue ultimately led to Bohr’s model, which introduced quantized energy levels to stabilize electron orbits.

He directed a beam of positively charged alpha particles at a very thin sheet of gold foil, only a few atoms thick (see Figure II.4). A fluorescent screen placed around the foil made it possible to detect where the particles landed. Most of the alpha particles passed straight through the foil, showing that the atom is largely empty space. Some particles, however, were deflected at small angles, indicating that they passed near something positively charged. Very rarely, a few alpha particles bounced back almost completely, a result that was both surprising and crucial. From these observations, Rutherford concluded that almost all the positive charge and nearly all the mass of the atom are concentrated in a tiny central region called the nucleus, while the much lighter electrons occupy the vast surrounding space. This explained why most particles were not deflected and why only a few experienced strong repulsion. The experiment disproved Thomson's "plum pudding" model and led to the development of the nuclear model of the atom, fundamentally transforming our understanding of atomic structure.

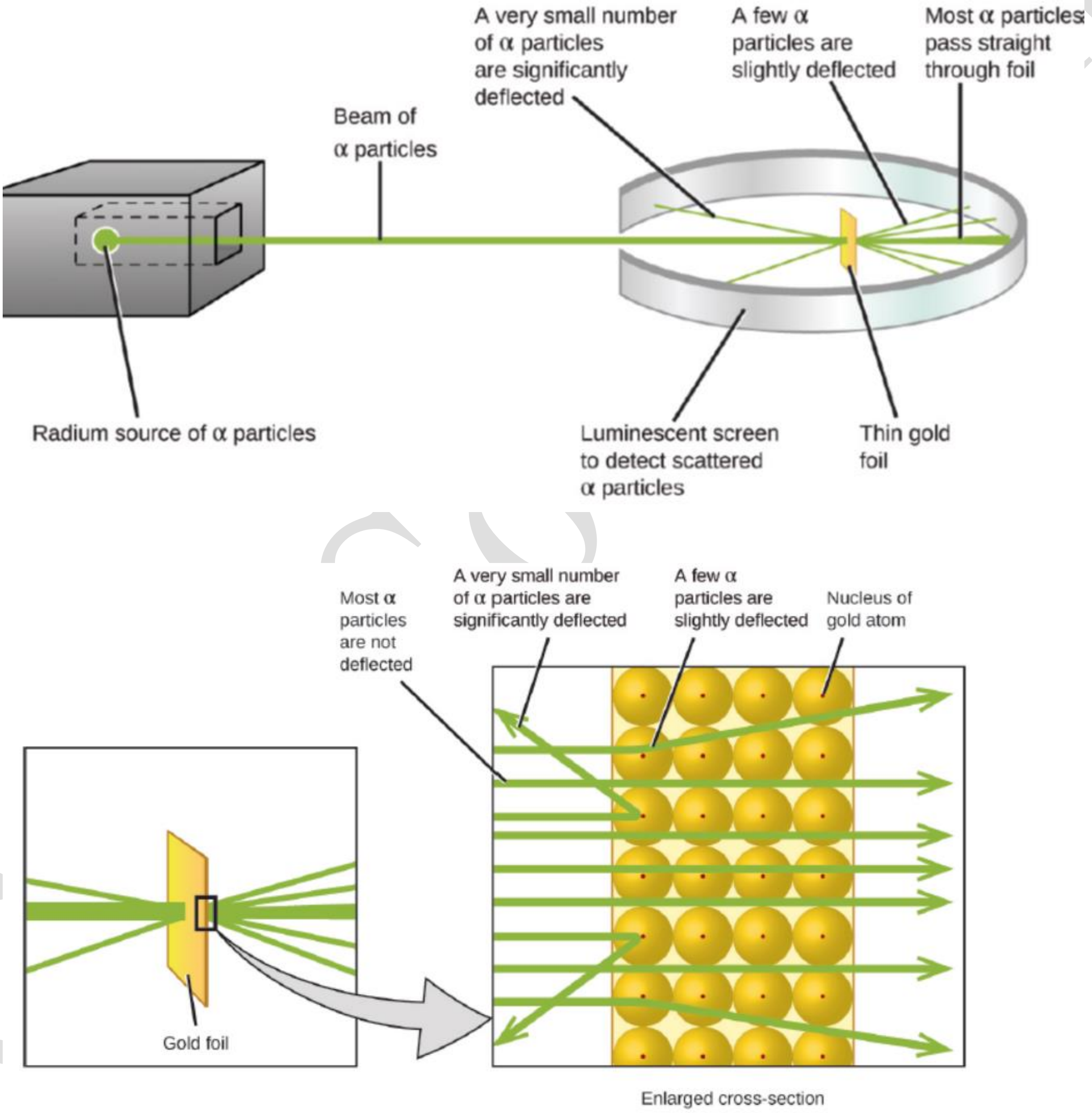


Figure II.4. Rutherford's experiment

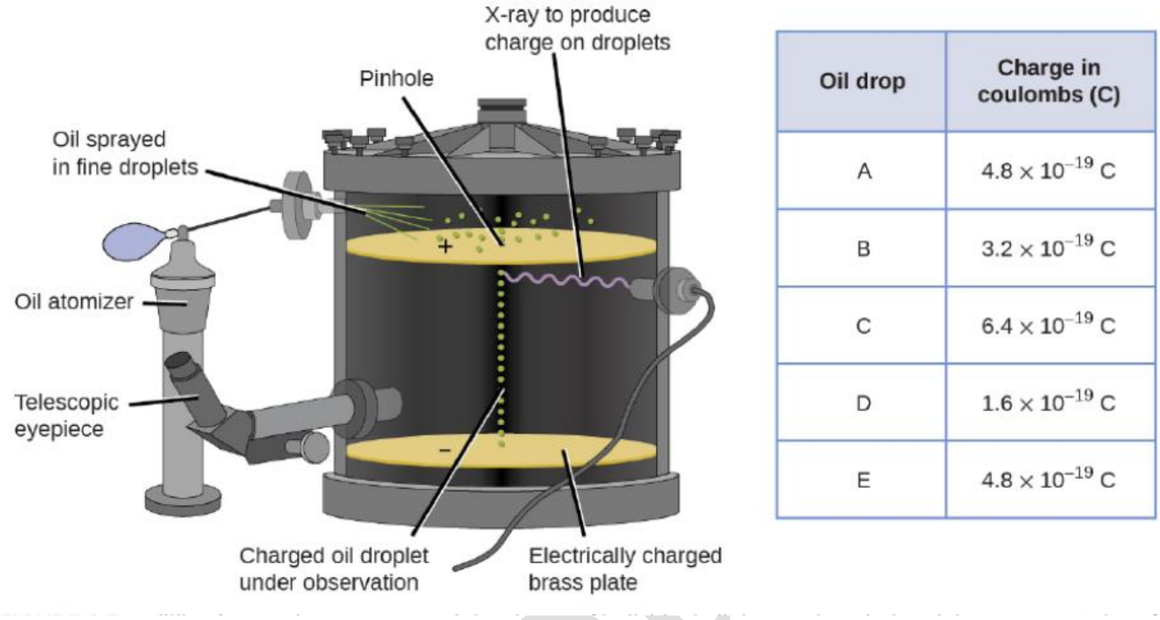
II.4 Millikan’s Experiment

In 1909, Robert Millikan performed his famous oil-drop experiment to determine the electric charge of the electron with remarkable precision (see Figure II.5). The setup consisted of two horizontal metal plates placed one above the other, creating a uniform electric field when a voltage was applied. Above the plates, Millikan introduced tiny droplets of oil using an atomizer, which produced microscopic drops capable of becoming electrically charged through friction. A small hole in the upper plate allowed some of these droplets to pass into the space between the two plates. A strong light source illuminated the chamber, and a microscope was positioned at the side so that Millikan could carefully observe the motion of individual droplets as they fell.

Under normal conditions, the oil droplets would fall due to gravity, but Millikan also used X-rays to ionize the air inside the chamber. This process allowed the droplets to pick up additional electrons, giving them a negative charge. Once charged, the droplets could be influenced not only by gravity but also by the electric force created between the plates. By adjusting the voltage applied to the plates, Millikan could slow down, stop, or even make a droplet rise. When the electric force perfectly balanced the weight of a droplet, the droplet appeared to hover motionless. At that moment, he could calculate the electric charge on the droplet using the known electric field strength and the measured mass of the droplet, which he determined from its rate of fall when no electric field was applied.

Millikan repeated the procedure on many individual droplets, each time observing changes in their motion and determining their charges. He noticed that the charges were always multiples of a smallest, fundamental value. No matter how many droplets he tested, each charge measurement corresponded to an integer multiple of this elementary charge. This was the key observation that allowed him to conclude that electric charge is quantized. The smallest measured value represented the charge of a single electron.

From his precise and repeated measurements, Millikan calculated the electron’s charge to be approximately -1.6×10^{-19} coulomb, a value extremely close to the one accepted today. His experiment therefore provided the first accurate determination of the electron’s charge and confirmed that electrons carry a constant, indivisible quantity of negative charge. The oil-drop experiment also offered strong support for the atomic theory of matter, since it demonstrated that electrical charge behaves as discrete packets rather than as a continuous fluid. Millikan’s work was a major milestone in modern physics and helped refine our understanding of subatomic particles.



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

Figure II.5. Millikan’s experiment. The tabulated data are examples of a few possible values.

II.5 Physical Properties of the Components of the Atom

The atom is made up of three main particles: the proton, the neutron, and the electron. The proton (p^+) carries a positive charge of +1 and has a mass of about 1 u, which is approximately 1.67×10^{-27} kg. The neutron (n^0) is electrically neutral, with a charge of 0, and has a mass nearly identical to that of the proton, also about 1 u. Protons and neutrons together form the nucleus, the dense central region of the atom that contains almost all of its mass. The electron (e^-), on the other hand, carries a negative charge of –1 and has a very small mass, roughly 1/1836 that of the proton. Electrons move around the nucleus in defined energy levels. For example, a hydrogen atom contains 1 proton and 1 electron but no neutrons, whereas an oxygen atom has 8 protons, 8 neutrons, and 8 electrons. In any neutral atom, the number of protons equals the number of electrons, ensuring the balance of charges. These fundamental properties of subatomic particles allow each element to be identified and help explain its chemical behavior.

II.6 Presentation and Characteristics of the Atom

- An atom is the smallest unit of a chemical element that retains its properties.
- The chemical symbol (e.g., H for hydrogen, O for oxygen) represents the atom.
- The atomic number (Z) corresponds to the number of protons in the nucleus.
- In a neutral atom, Z = number of electrons.
- The mass number (A) corresponds to the sum of protons and neutrons in the nucleus.
- The number of neutrons is calculated as: $N = A – Z$.
- Example: oxygen atom: symbol O, $Z = 8$, $A = 16$ → neutrons $N = 16 – 8 = 8$.
- Example: carbon atom: symbol C, $Z = 6$, $A = 12$ → neutrons $N = 12 – 6 = 6$.
- The atomic structure determines the chemical properties of the element and its position in the periodic table.

II.7 Isotopes

1. Isotopy and Relative Abundance of Different Isotopes

- An isotope is an atom of the same element (same Z) with a different number of neutrons.
- Isotopes have the same chemical properties but slightly different physical properties (mass, density).
- Example: carbon has three isotopes: ^{12}C (98.9%), ^{13}C (1.1%), ^{14}C (very low, radioactive).
- Relative abundance is the proportion of each isotope in nature.
- It is used to calculate the average atomic mass of an element.
- Example: for chlorine: ^{35}Cl (75%), ^{37}Cl (25%) → average atomic mass ≈ 35.5 u.

2. Separation of Isotopes

- Isotopes of the same element can be separated due to their mass difference.
- Physical methods: gas diffusion, centrifugation, mass spectrometry (see figure II.6).
- Example: separation of ^{235}U and ^{238}U for nuclear energy.
- In mass spectrometry, ions are deflected according to their mass/charge ratio (m/z).
- Lighter isotopes are deflected more than heavier isotopes.

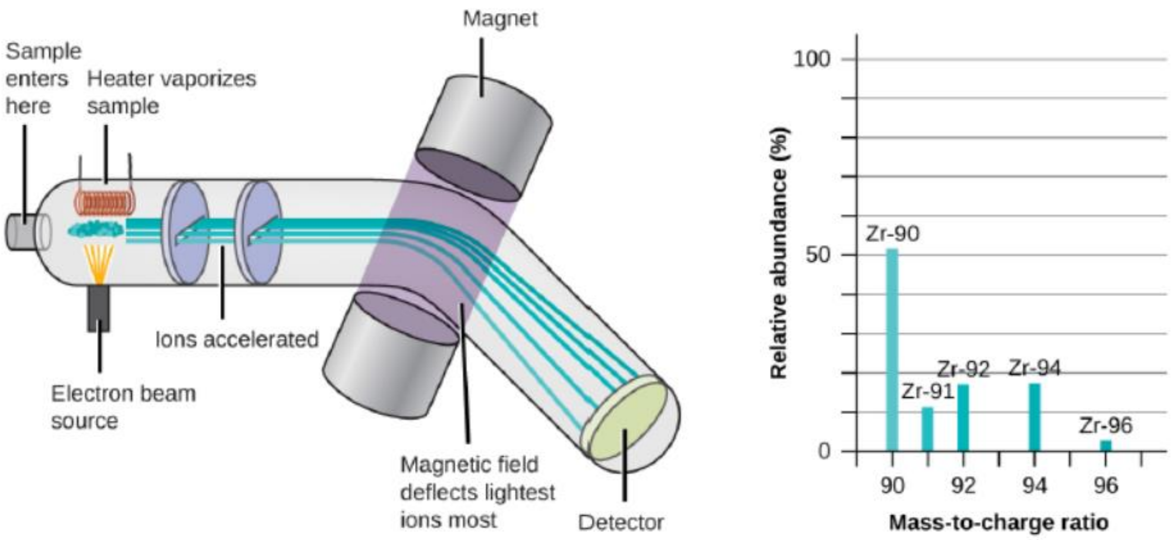


Figure II.6. Mass spectrometer

3. Determination of Atomic Mass and Average Atomic Mass

• The atomic mass of an isotope is measured in atomic mass units (u).

• The average atomic mass of an element is the weighted average according to relative abundance.

• Example: for chlorine

- ^{35}Cl → 34.97 u × 0.75 = 26.23 u
- ^{37}Cl → 36.97 u × 0.25 = 9.24 u
- Average mass ≈ 35.47 u. This average mass corresponds to the value indicated in the periodic table.

4. Mass Spectrometry: Bainbridge Mass Spectrograph

• The Bainbridge spectrograph allows measurement of ion masses and isotope separation.

• Ions are accelerated by an electric field, then deflected by a magnetic field.

• The deflection depends on the m/z ratio: lighter ions are deflected more.

• Isotopes appear separately on a detector, allowing measurement of relative abundance and exact mass.

• Example: isotopes of lithium

^{6}Li and ^{7}Li clearly separated by the spectrograph.

II.8 Nuclear Binding Energy

The nuclear binding energy is the amount of energy required to completely separate a nucleus into its individual nucleons. It reflects the stability of a nucleus: the higher the binding energy per nucleon, the more stable the nucleus. For example, the nucleus of ^{4}He (an alpha particle) has a high binding energy, making it very stable. The cohesive energy refers to the energy that holds the nucleons together within the nucleus. This energy is responsible for the large amounts of energy released during nuclear reactions such as fission and fusion.

II.9 Stability of an Atomic Nucleus

The stability of an atomic nucleus depends largely on its neutron-to-proton ratio (N/Z). Light nuclei tend to be stable when the numbers of neutrons and protons are approximately equal, whereas heavy nuclei require more neutrons than protons to counteract the increasing Coulomb repulsion between protons. Nuclei that fall outside these stability conditions undergo radioactive decay in order to reach a more stable state. For instance, ^{14}C, with an N/Z ratio of 1.33, is radioactive and decays into ^{14}N through β^- emission. Additionally, nuclei with "magic numbers" of nucleons (2, 8, 20, 28, 50, 82, 126) exhibit exceptional stability.

Key points to remember:

- Isotopes explain the differences in mass between atoms of the same element.
- Mass spectrometry makes it possible to measure both mass and isotopic abundance.
- Binding energy provides information about nuclear stability and the release of nuclear energy.
- Understanding these concepts is essential for nuclear chemistry, physics, and many industrial and medical applications.

Chapter III. Radioactivity – Nuclear Reactions

This chapter explores the fundamental concepts of radioactivity and nuclear reactions, which are central to understanding atomic nuclei. It introduces the different types of radioactive decay and the principles governing nuclear transformations. Students will learn how nuclear reactions release energy and how this energy is applied in science, medicine, and industry. Key concepts such as nuclear stability, binding energy, and isotopes are also discussed. This chapter provides a foundation for further study in nuclear chemistry and physics.

III.1 Natural Radioactivity

Natural radioactivity refers to the spontaneous decay of unstable nuclei present in nature. This process is independent of temperature, pressure, or chemical state. Certain heavy nuclei, such as uranium-238 (U-238), thorium-232 (Th-232), and radium-226 (Ra-226), are naturally unstable and transform into more stable forms by emitting radiation. The three main types of natural radiation are alpha (α), beta (β), and gamma (γ) radiation.

Alpha (α) Radiation

Alpha emission involves the expulsion of a helium nucleus ${}^{4}_{2}He$ containing 2 protons and 2 neutrons, carrying a +2 charge and a mass of 4 u. The emission of an alpha particle decreases the atomic number (Z) by 2 and the mass number (A) by 4.

For example: ${}^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He$

Alpha particles have low penetration power and can be stopped by a sheet of paper.

Beta (β) Radiation

Beta radiation occurs in two forms: β^-, where an electron is emitted, and β^+, where a positron is emitted.

In β^- decay, a neutron converts into a proton:

$$n \rightarrow p + e^- + \bar{\nu}_e \qquad [III.1]$$

- A neutron becomes a proton
- Emission of an electron
- Emission of an electron antineutrino

increasing Z by 1 while A remains unchanged. Example: ${}^{14}_{6}C \rightarrow {}^{14}_{7}N + e^-$

In β^+ decay, a proton transforms into a neutron, decreasing Z by 1 with no change in A.

$$p \rightarrow n + e^+ + \nu_e \qquad [III.2]$$

- A proton becomes a neutron
- Emission of a positron
- Emission of an electron neutrino

Beta particles are moderately penetrating and can be stopped by materials such as aluminum.

Gamma (γ) Radiation

Gamma radiation is high-energy electromagnetic radiation consisting of photons (γ) without any mass or charge. It often accompanies α or β decay when the nucleus remains in an excited state, leaving Z and A unchanged. Gamma rays are highly penetrating and require dense materials like lead for shielding.

III.2 Artificial Radioactivity

Artificial radioactivity was discovered by Irène and Frédéric Joliot-Curie in 1934. It involves making a stable nucleus radioactive by bombarding it with particles such as neutrons, protons, deuterons, alpha particles, or gamma photons.

For example: ${}^{27}_{13}Al + {}^{4}_{2}He \rightarrow {}^{30}_{15}P + n$

The resulting nucleus is unstable and decays further, here by β^+ emission. Artificial radioactivity is widely used to produce medical isotopes.

III.3 Nuclear Reactions

A nuclear reaction modifies the structure of the nucleus while conserving Z, A, and total energy. In general, a reaction can be represented as:

target nucleus + projectile → product 1 + product 2 [III.3]

An example of fusion is: ${}^{3}_{1}H + {}^{2}_{1}H \rightarrow {}^{4}_{2}He + n$

Nuclear reactions can be categorized as fusion, fission, or capture. Fission splits a heavy nucleus, such as U-235, while fusion combines two light nuclei, as occurs in the Sun.

III.4 Kinetics of Radioactive Decay

Radioactive decay is a random, spontaneous process. Each nucleus has a constant probability of decaying per unit time, called the decay constant (λ). The number of remaining nuclei follows the exponential law:

$$N(t) = N_0 \, e^{-\lambda t} \qquad [III.4]$$

The half-life ($t_{1/2}$) is the time required for half of the nuclei to decay, calculated as:

$$t_{1/2} = \ln 2 / \lambda \qquad [III.5]$$

For instance, carbon-14 has a half-life of 5730 years. The decay rate, or activity, is the number of transformations per second, measured in becquerels (Bq), where 1 Bq = 1 decay per second.

III.5 Applications of Radioactivity

Radioactivity has numerous applications across medicine, industry, archaeology, energy, and research. In medicine, it is used for imaging (PET scans, scintigraphy) and radiotherapy (cobalt-60). Carbon-14 dating allows archaeologists to determine the age of artifacts. Industrial uses include thickness control and leak detection. Nuclear energy relies on controlled fission in reactors, while scientific research employs isotopic tracing and studies of chemical reactions.

Key points to remember

- Radioactivity is the spontaneous transformation of unstable atomic nuclei into more stable nuclei through the emission of particles or electromagnetic radiation.

- The three main types of radioactive decay are alpha (α), beta (β), and gamma (γ) radiation, each differing in nature, charge, mass, and penetrating power.

- Radioactive decay follows an exponential law described by:

$N(t) = N_0 e^{-\lambda t}$ where λ is the decay constant.

- The half-life ($t_{1/2}$) is the time required for half of the radioactive nuclei in a sample to decay and is given by:

$t_{1/2} = \ln 2 / \lambda$.

- Radioactive decay is a random process at the level of individual nuclei, but it follows predictable statistical behavior for a large number of atoms.

Chapter IV. Electronic Structure of the Atom

IV.I Light: Wave – Particle Duality

Light is an electromagnetic wave (See Figure IV.1) that propagates through a vacuum at a speed of approximately $c = 3 \times 10^8$ m/s. It exhibits wave-like properties (See Figure IV.2) such as interference, diffraction, and polarization. For example, when a beam of light passes through two slits, it produces an interference pattern of alternating bright and dark fringes. At the same time, light is composed of photons, which are massless particles that carry energy given by:

$$E = h\nu \qquad [IV.1]$$

where h is Planck's constant and ν is the frequency of the radiation.

This particle aspect of light explains phenomena such as the photoelectric effect, where a single photon can eject an electron from a metal surface. The dual wave-particle nature of light means that it behaves both as a wave and as a stream of particles depending on the experiment. For instance, wave behavior is observed in the diffraction of a laser beam around an obstacle, while particle behavior is evident when electrons are emitted from a metal plate under illumination. This duality is fundamental in quantum physics for understanding the true nature of light, as the photon embodies both aspects: it propagates as a wave while exchanging energy as a particle.

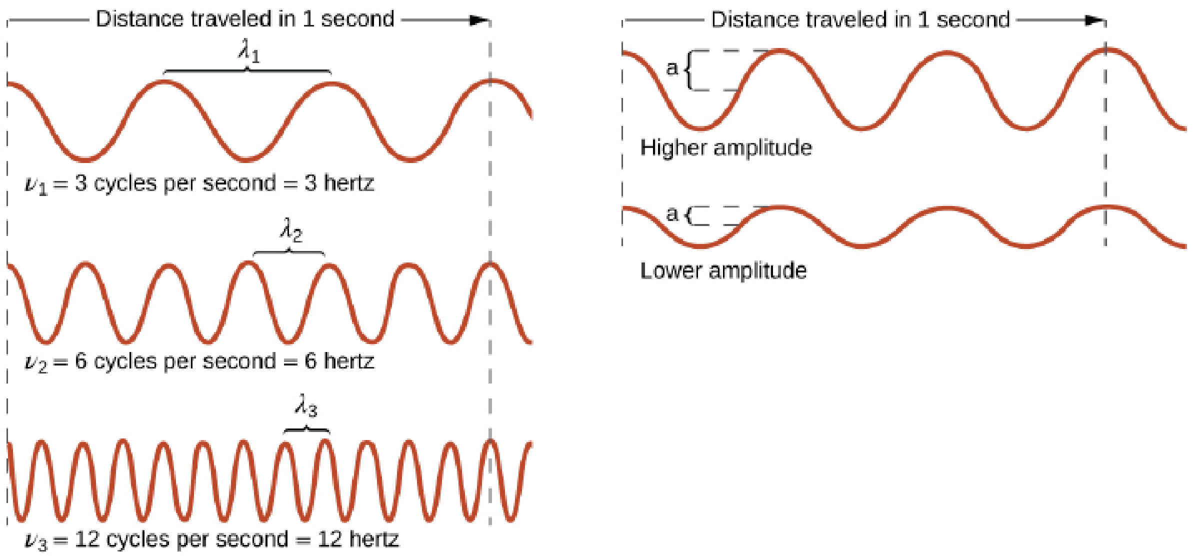


Figure IV.1. The wave

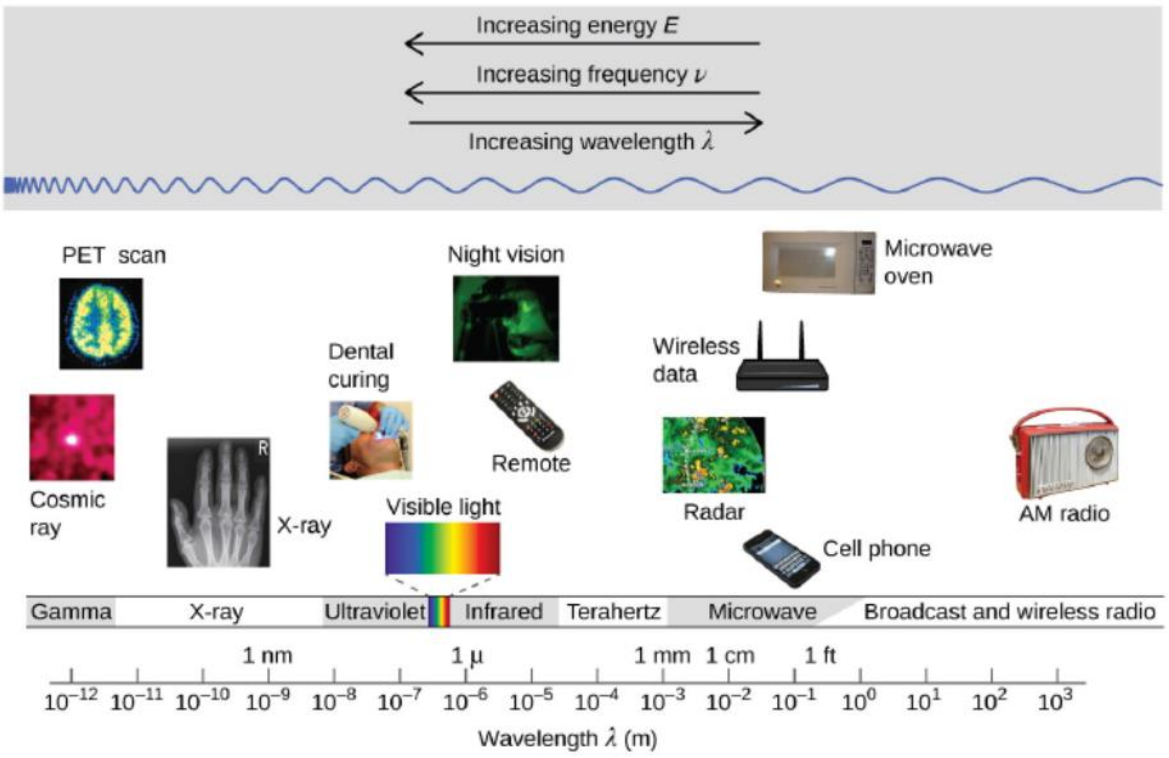


Figure IV.2. Portions of the electromagnetic spectrum

IV.2 Interaction Between Light and Matter: Absorption and Emission

The interaction of light with matter involves two main processes: absorption and emission. During absorption, an atom or molecule captures a photon when the photon’s energy precisely matches the energy difference between two discrete energy levels. In this process, an electron moves from a lower energy level (E_1) to an excited level (E_2). For example, a blue pigment absorbs red and green light, allowing only blue light to be transmitted or reflected. Emission occurs when an excited atom returns to a lower energy level, releasing a photon whose energy corresponds exactly to the difference $E_2 - E_1$. The energy of the absorbed or emitted photon is $h\nu$. Since each atom has discrete energy levels, this explains the observed absorption and emission spectra. A well-known example is hydrogen, which exhibits clearly defined absorption and emission lines corresponding to its quantized energy levels.

IV.3 Planck’s Theory

In 1900, Max Planck proposed that light energy is not continuous but quantized. He suggested that a hot object does not emit energy in a continuous manner, but in discrete packets called quanta. Each quantum carries an energy E (see Eq. [IV.1]).

This hypothesis successfully explained the blackbody radiation spectrum, which classical physics had failed to describe. The higher the frequency of the light, the greater the energy of each quantum; for example, an ultraviolet photon carries more energy than an infrared photon. The intensity of light depends on the number of quanta emitted rather than on the energy of individual quanta. Planck's theory introduced the concept that light possesses a particle-like nature, laying the foundation for quantum physics and explaining phenomena such as the photoelectric effect. For instance, a heated filament emits more visible photons as its temperature increases.

Planck's constant, $h = 6.626 \times 10^{-34}$ J·s, is a fundamental parameter of nature.

IV.4 Photoelectric Effect

The photoelectric effect is the emission of electrons from a metallic surface when it is illuminated by light of sufficiently high energy (see figure IV.3). A photon transfers all its energy ($h\nu$) to an electron in the metal; if this energy exceeds the work function W of the metal, the electron is ejected. The condition for emission is $h\nu \geq W$, where h is Planck's constant and ν is the frequency of light. If the frequency is too low, no electrons are emitted regardless of light intensity. For example, red light cannot eject electrons from sodium, while ultraviolet light can. The kinetic energy Ec of the emitted electrons is given by

$$Ec = h\nu - W \qquad [IV.2]$$

Increasing the light intensity does not increase the energy of the electrons, only the number of electrons emitted. The photoelectric effect demonstrates the particle nature of light, as it behaves as a stream of photons. This phenomenon enabled Einstein to establish the wave-particle duality of light. Applications of the photoelectric effect include photovoltaic cells, optical sensors, and automatic doors.

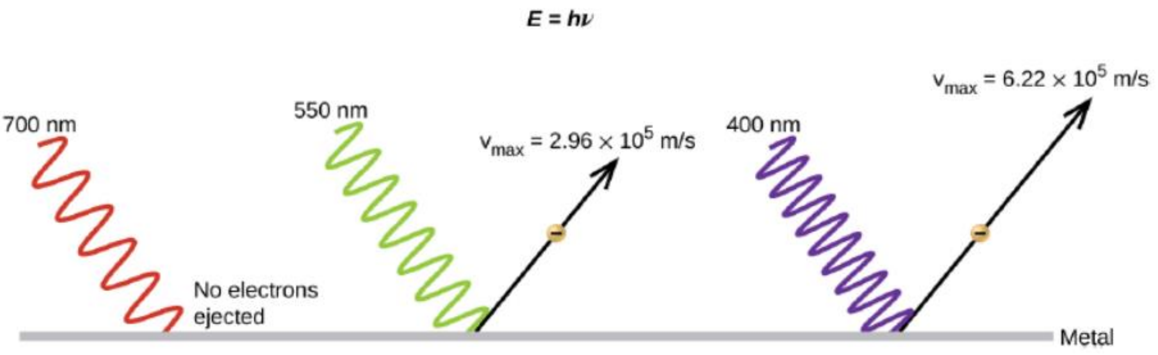


Figure IV.3. Photoelectric effect

IV.5 Bohr's Atomic Model – Hydrogen Atom

In 1913, Niels Bohr proposed that the electron in a hydrogen atom revolves around the nucleus in quantized circular orbits. Only specific orbits are allowed (see Figure IV.4), where the electron's angular momentum is given by

$$n \cdot h/2\pi \qquad [IV.3] \qquad \text{with } n = 1, 2, 3\ldots.$$

Each orbit corresponds to a discrete energy level, meaning that the electron cannot exist between two levels. The state n = 1 represents the ground state, while states with n > 1 are excited states. An electron does not emit energy as long as it remains in an allowed orbit. When it transitions from a higher energy level E_2 to a lower level E_1, it emits a photon with energy

$$h\nu = E_2 - E_1 \qquad [IV.4]$$

Conversely, a photon with exactly this energy can excite the electron to a higher level. For example, transitions in the hydrogen atom produce the lines of the Balmer series, which are visible in the spectrum. Bohr's model successfully explains atomic stability and the structure of hydrogen's emission and absorption spectra. Although it has been partially superseded by modern quantum mechanics, it remains a fundamental model for understanding energy quantization in atoms.

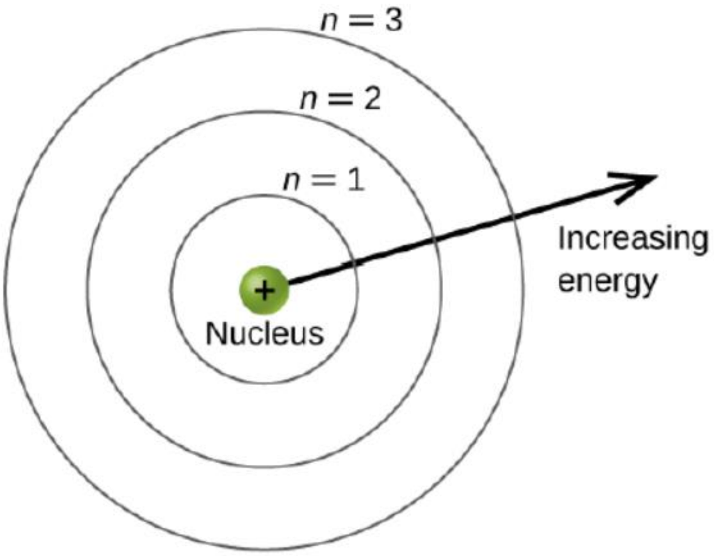


Figure IV.4. Different shells are numbered by principal quantum numbers

IV.6 Hydrogen Atom Spectrum

The hydrogen atom spectrum (see Figure IV.5) consists of lines corresponding to electron transitions between quantized energy levels. When an electron moves from a higher level (n_2) to a lower level (n_1), the atom emits a photon with energy hν (see Eq. [IV.4]). Conversely, the atom absorbs a photon if hν matches exactly the difference between two energy levels. Therefore, the spectral lines are not continuous; they are discrete and precise for each transition. Transitions to $n = 1$ form the Lyman series in the ultraviolet (UV) region, while transitions to $n = 2$ form the Balmer series in the visible region. For instance, the red line of the Balmer series corresponds to the transition $n = 3 \rightarrow n = 2$. Transitions to $n = 3$ form the Paschen series in the infrared region. Each line corresponds exactly to an energy difference predicted by Bohr's model. The hydrogen spectrum provides experimental evidence for energy quantization and is used to determine hydrogen's energy level structure and atomic constants.

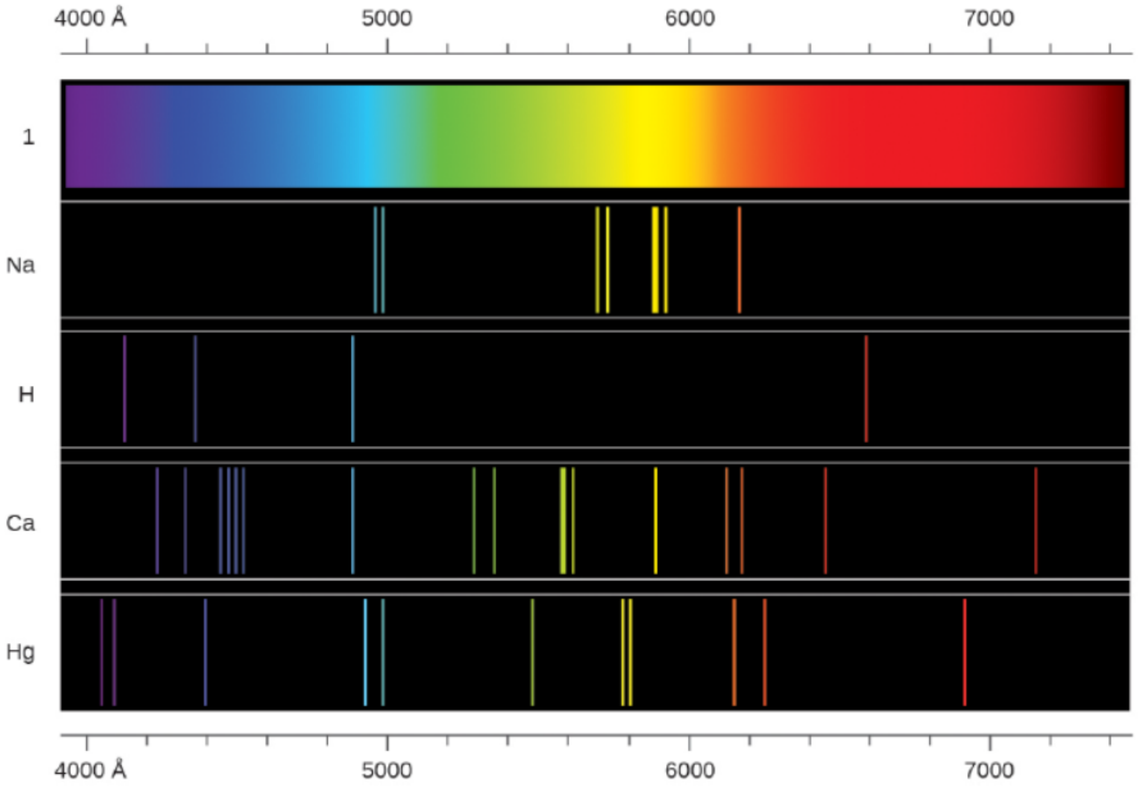


Figure IV.5. Emission spectra

IV.7 Rydberg Equation

The Rydberg equation describes the wavelengths emitted or absorbed by the hydrogen atom when the electron changes energy levels (see Figure IV.7).

It is written as:

$$1/\lambda = R_H (1/n_1^2 - 1/n_2^2) \qquad [IV.5]$$

where λ is the wavelength (m), R_H is the Rydberg constant, with $n_1 < n_2$.

- n_1 corresponds to the electron’s final level, and n_2 to the initial level.
- If the electron moves down (e.g., 3 → 2), the atom emits light.
- If the electron moves up (e.g., 2 → 4), the atom absorbs light.
- The equation allows calculation of the spectral lines for the Lyman (UV), Balmer (visible), Paschen (IR) series, etc.
- Example: for a transition from $n_2 = 3$ to $n_1 = 2$, the formula gives a value in the visible spectrum (Balmer series).
- The Rydberg constant is approximately $R_H = 1{,}097.\ 10^7\ m^{-1}$

This equation confirms Bohr’s model and explains the lines observed in the hydrogen spectrum.

It is fundamental for understanding electronic structure and atomic spectra.

IV.8 Bohr’s Postulates

Bohr proposed that the electron in the hydrogen atom can occupy only certain well-defined energy levels, called quantized orbits.

On these allowed orbits, the electron does not radiate energy, even though it is in circular motion.

- Energy levels are characterized by a quantum number n = 1, 2, 3,…
- The energy of each level is given by:

$$E_n = -13{,}6 / n^2\ (eV) \qquad [IV.6]$$

- The electron can change levels only by absorbing or emitting a photon with energy equal to the difference between the levels:

$$E_{photon} = E_{n2} - E_{n1} \qquad [IV.7]$$

- Absorption: the electron moves to a higher energy level.
- Emission: the electron falls to a lower level and the atom emits light.

- Allowed orbits are such that the electron's angular momentum is quantized:

$$m\,v\,r = n\,h/2\pi \qquad [IV.8]$$

- The radius of the nth orbit r_n of the hydrogen atom is given by:

$$r_n = a_0\,n^2 \qquad [IV.9]$$

Where r_n is the radius of orbit number n, n = 1, 2, 3,… is the principal quantum number, and a_0 is the Bohr radius (constant), $a_0 = 0{,}529.10^{-10}$ m

These postulates explain the spectral lines of the hydrogen atom observed experimentally.

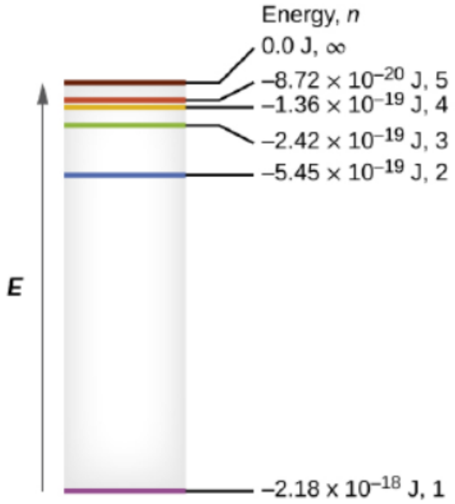


Figure IV.6. Quantum numbers and energy levels in a hydrogen atom

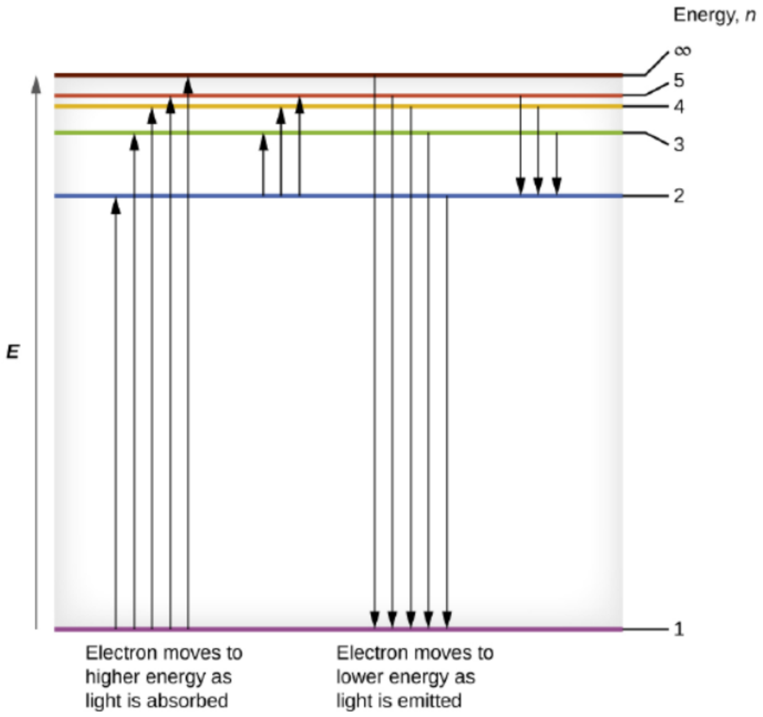


Figure IV.7. The relative energy of orbits in the Bohr model of the hydrogen

IV.9 The hydrogen atom in wave mechanics

The electron is no longer considered as a particle moving on a fixed orbit, but rather as a probability wave described by a wave function ψ. This function ψ is obtained by solving the Schrödinger equation for the hydrogen atom. The solutions of Schrödinger's equation yield quantized energy levels, identical to those predicted by Bohr's model.

Regions where $|\psi|^2$ is high indicate where the electron is most likely to be found (probability). These regions are called atomic orbitals, for example: $1s$, $2s$, $2p$. For instance, the 1s orbital is spherical around the nucleus, while the 2p orbitals have a dumbbell shape (see Figure IV.9). The energy of an orbital depends on the principal quantum number n (1, 2, 3…), as in Bohr's hydrogen atom. Wave mechanics explains phenomena that Bohr's model cannot interpret, such as probability densities. The electron no longer has a defined trajectory: it is described as an electron cloud distributed around the nucleus. This wave description allows us to understand the electronic structure of more complex atoms and chemical bonding.

IV.10 Louis De Broglie Law

In 1924, Louis de Broglie proposed that every material particle also has a wave-like nature. He associated a wavelength with each particle, called the de Broglie wavelength. This wavelength is given by the formula:

$$\lambda = h / p = h / mv \qquad [IV.9]$$

where h is Planck's constant and p is the momentum of the particle of mass m.

De Broglie used the term matter waves to describe the wave characteristics of material particles.

The faster a particle moves (larger p), the shorter its wavelength. For example, an electron in an atom has a wavelength comparable to the size of the atom, leading to significant quantum effects. In contrast, a tennis ball has an enormous momentum, resulting in an extremely small wavelength, making wave-like effects imperceptible. De Broglie's law explains why electrons exhibit interference and diffraction, similar to light waves. This wave-particle duality was experimentally confirmed by electron diffraction experiments performed by Davisson and Germer. It unifies the behavior of light (waves) and matter (particles). De Broglie's principle is fundamental for understanding quantum mechanics and modern atomic models.

IV.11 The Uncertainty Principle

The discovery of the wave properties of matter raised some new and interesting questions. Consider, for example, a ball rolling down a ramp. Using the equations of classical physics, we can calculate, with great accuracy, the ball's position, direction of motion, and speed at any instant. Can we do the same for an electron, which exhibits wave properties? A wave extends in space, and its location is not precisely defined. We might therefore anticipate that it is impossible to determine exactly where an electron is located at a specific instant. The German physicist Werner Heisenberg proposed that the dual nature of matter places a fundamental limitation on how precisely we can know both the location and the momentum of an object at a given instant. The limitation becomes important only when we deal with matter at the subatomic level (that is, with masses as small as that of an electron). Heisenberg's principle is called the uncertainty principle. When applied to the electrons in an atom, this principle states that it is impossible for us to know simultaneously both the exact momentum of an electron and its exact location in space. Heisenberg mathematically related the uncertainty in position, Δx, and the uncertainty in momentum, $\Delta(mv)$, to a quantity involving the Planck constant:

$$\Delta x \,.\, \Delta(mv) \geq h / 4\pi \qquad [IV.10]$$

De Broglie's hypothesis and Heisenberg's uncertainty principle set the stage for a new and more broadly applicable theory of atomic structure. In this approach, any attempt to define precisely the instantaneous location and momentum of the electron is abandoned. The wave nature of the electron is recognized, and its behavior is described in terms appropriate to waves. The result is a model that precisely describes the energy of the electron while describing its location not precisely but rather in terms of probabilities.

We can never know the exact position and momentum of the electron; we can talk about the probability of its being at certain locations in space.

IV.12 Quantum Numbers and Orbitals

Electrons in an atom are described by four quantum numbers, derived from the solution of Schrödinger's equation (see Figure IV.10).

1) Principal quantum number n: n =1, 2, 3, …

defines the electron shell and determines the energy level and size of the orbital.

Example: n = 1 → shell K, n = 2 → shell L.

2) Azimuthal quantum number l: l = 0, 1, 2, …, (n−1)

defines the subshell and determines the shape of the orbital.

Example: l = 0 → s orbital (spherical), l = 1 → p orbital (dumbbell), l = 2 → d orbital (cloverleaf).

3) Magnetic quantum number ml: ml = − l, …, 0, …,+ l

defines the number of orbitals and their spatial orientation.

Example: for l = 1, ml = −1, 0, +1 → three orbitals p_x, p_y, p_z.

4) Spin quantum number m_s:

describes the spin of the electron, with two possible values: +1/2 or −1/2.

Two electrons in the same orbital must have opposite spins (Pauli exclusion principle).

- Together, these four quantum numbers define the complete state of each electron in an atom.
- Example: an electron in a 2p orbital is described by: n = 2, l = 1, ml = 0 (or -1 or 1), and m_s = +1/2 (or −1/2).
- They allow determination of electron configurations and chemical properties of atoms.

- An atomic orbital is described by three quantum numbers: n, l, and ml.

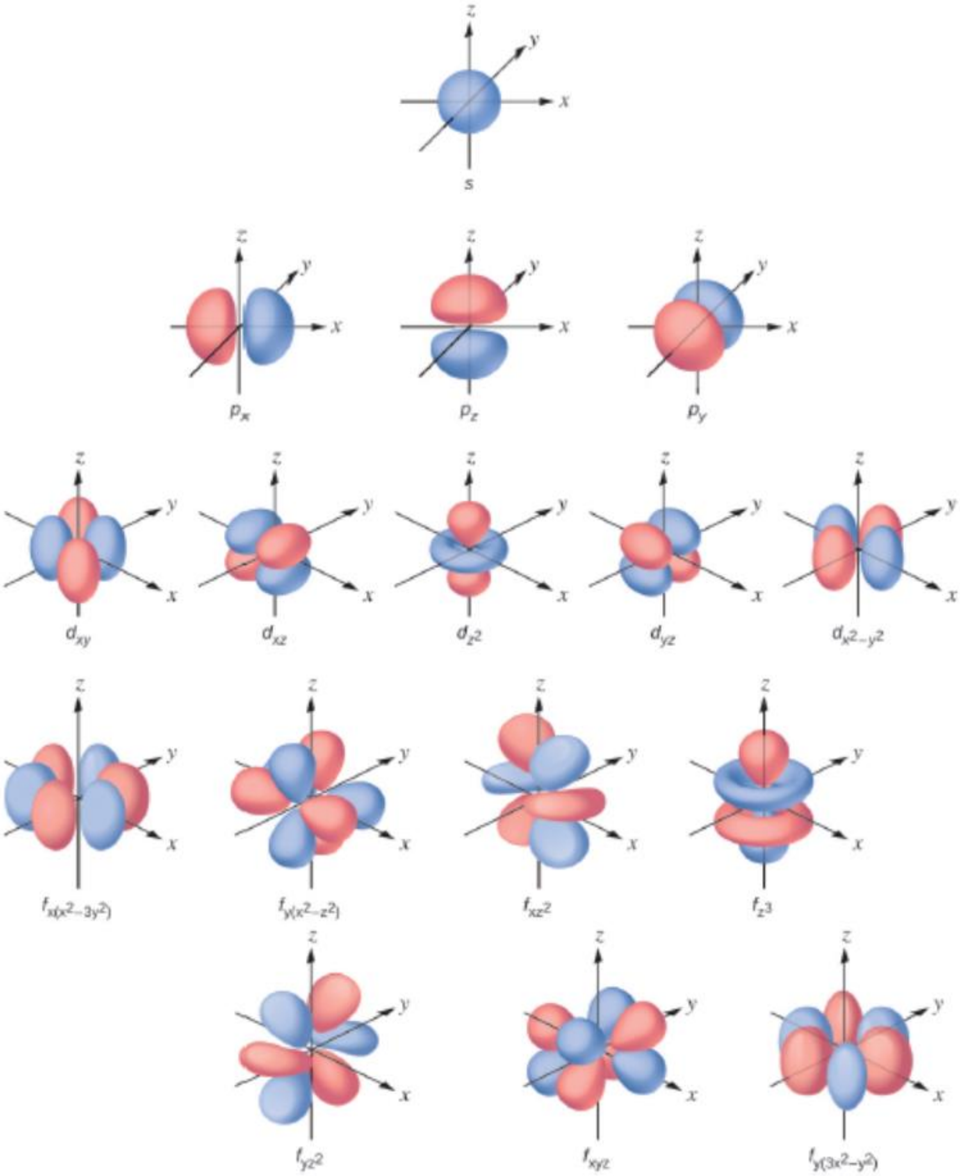
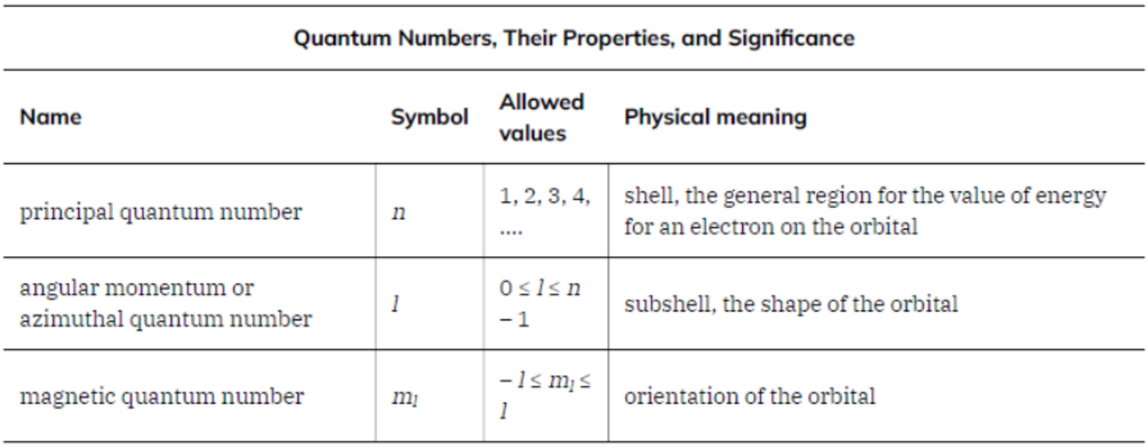
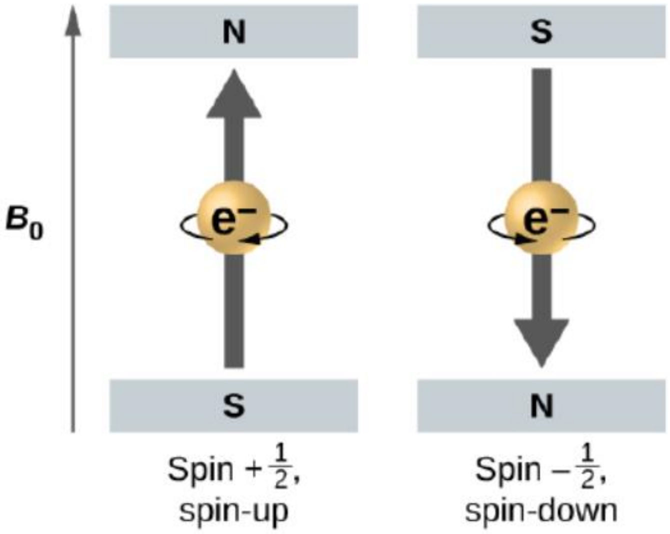


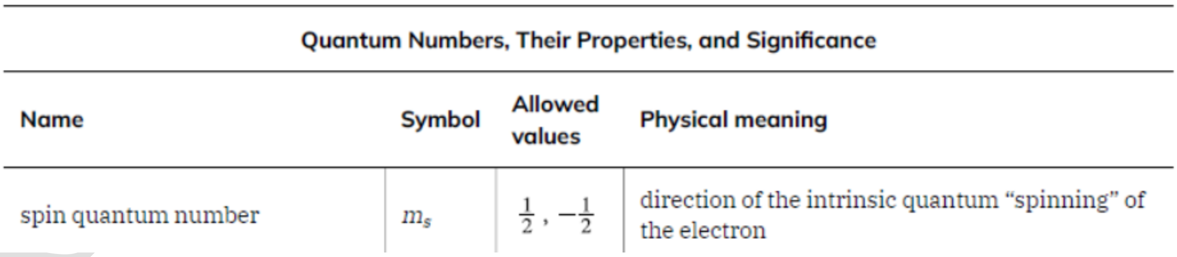
Figure IV.9. Shapes of s, p, d and f orbitals



Quantum Numbers, Their Properties, and Significance

Name	Symbol	Allowed values	Physical meaning
principal quantum number	n	1, 2, 3, 4, ….	shell, the general region for the value of energy for an electron on the orbital
angular momentum or azimuthal quantum number	l	$0 \leq l \leq n - 1$	subshell, the shape of the orbital
magnetic quantum number	m_l	$-l \leq m_l \leq l$	orientation of the orbital





Quantum Numbers, Their Properties, and Significance

Name	Symbol	Allowed values	Physical meaning
spin quantum number	m_s	$\frac{1}{2}, -\frac{1}{2}$	direction of the intrinsic quantum "spinning" of the electron

Figure IV.10. Quantum numbers

IV.13 Multi-electron Atoms

The atomic model based on quantum mechanics (quantum model) allows us to explain the properties and behavior of all atoms.

1) Orbital Energy

In the hydrogen atom, where there are no inter-electronic repulsions, all orbitals within the same shell have the same energy.
Example: the 2s orbital and the 2p orbital have the same energy. Likewise, the 3s, 3p, and 3d orbitals share the same energy.

In multi-electron atoms, two opposing effects influence orbital energy:

- The nucleus of a multi-electron atom has a higher nuclear charge Z than hydrogen (Z_H), which tends to stabilize the orbitals, i.e., lower their energy as electrons are more strongly attracted.
- Electrons repel each other, which tends to increase orbital energy.

The combination of these two factors determines the electronic structures and properties of elements.
In multi-electron atoms, due to inter-electronic repulsion, the orbital energies vary in the following order:

$$E_{ns} < E_{np} < E_{nd} < E_{nf}$$

Example: for the n = 3: $E_{3s} < E_{3p} < E_{3d}$ (see Figure IV.11).

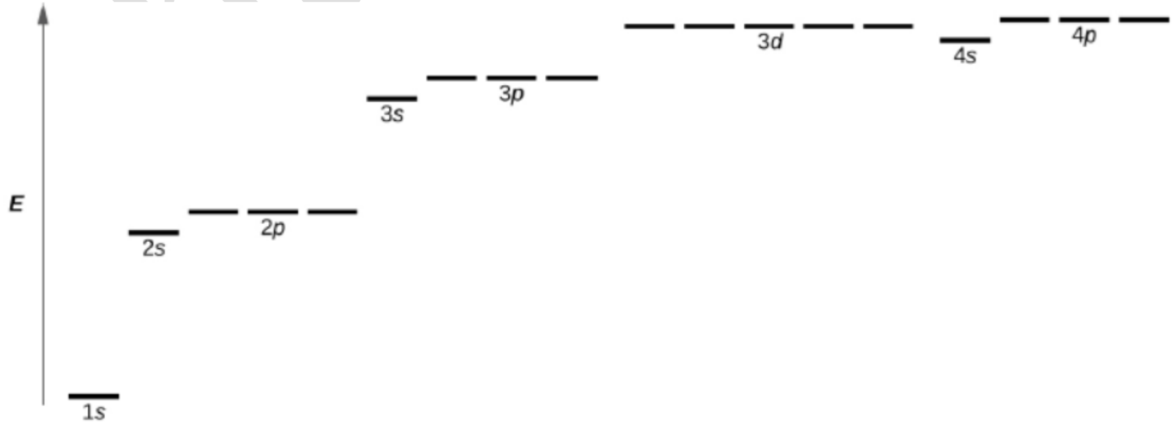


Figure IV.11. The energies of electron orbitals in a multi-electron atom

2) Stability Principle

In the ground state, an atom exists in its lowest energy configuration, corresponding to the most stable state.

Electrons first occupy the lowest available energy levels in the order: 1s, 2s, 2p, 3s… according to Klechkowski’s rule (see Figure IV.12).

Orbitals within the same subshell always have the same energy.
Example: the three 2p orbitals have the same energy.

Klechkowski’s Rule

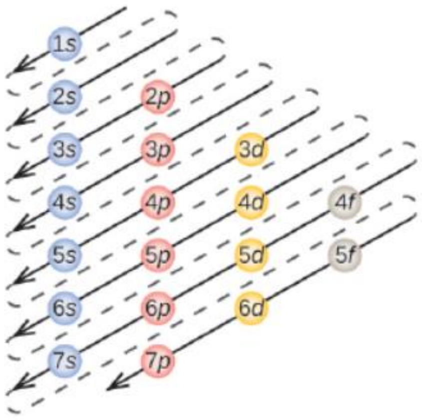


Figure IV.12. The energy order for atomic orbitals

3) Pauli Exclusion Principle

No two electrons in the same atom can have identical values for all four quantum numbers n, l, ml, and m_s. Electrons occupying the same orbital share the same n, l, and ml values, so they must have opposite m_s values. Consequently, each orbital can accommodate a maximum of two electrons with opposite spins.

Example: an orbital can contain:

- 0 electrons → empty orbital
- 1 electron → unpaired or single electron
- 2 electrons → spin $m_s = +1/2$ and $m_s = -1/2$, forming an electron pair; the orbital is saturated.

4) Hund's Rule

When electrons occupy orbitals of the same energy level, they fill them successively with parallel spins, maximizing the number of unpaired electrons.

An atomic orbital is represented as a square called a "box", with arrows indicating electrons; the direction of the arrow represents the spin.

5) Electronic Configuration of Atoms

The electronic configuration of an atom lists all occupied orbitals and the number of electrons they contain.

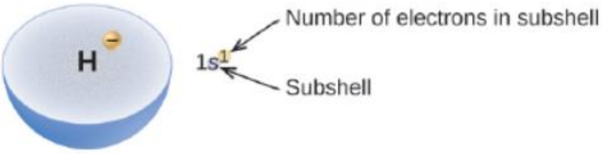
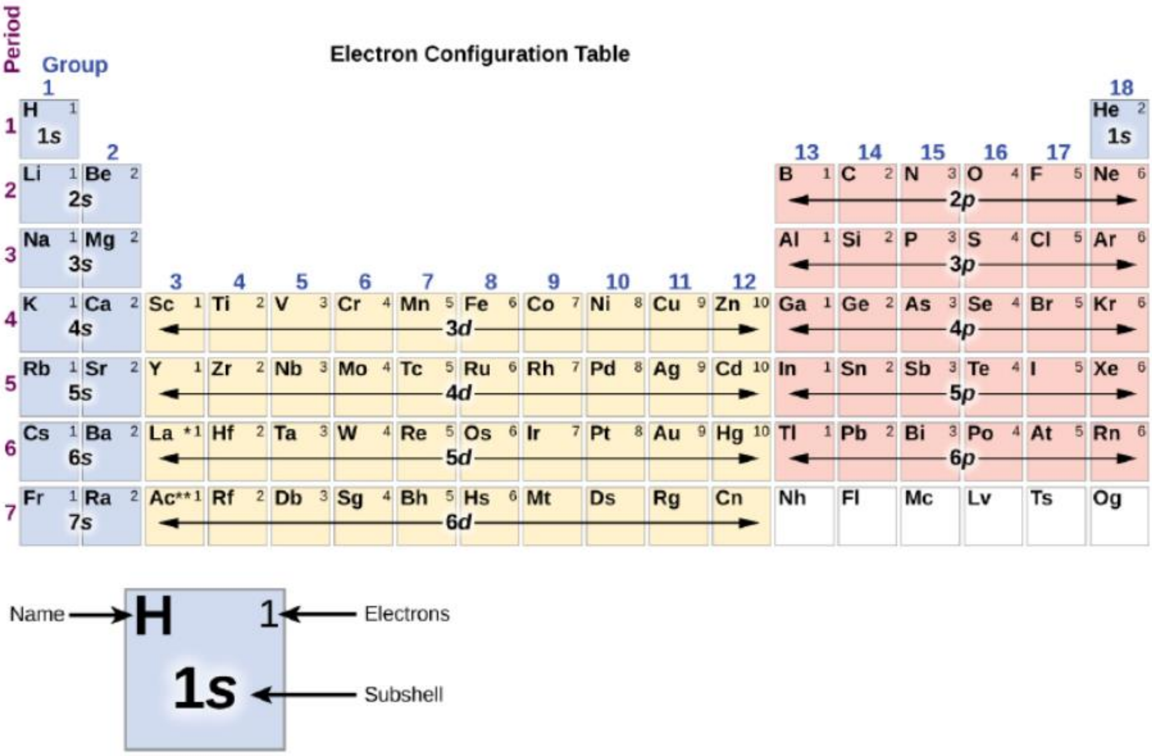


Figure IV.13. The diagram of an electron configuration specifies the subshell (n and l value, with letter symbol) and superscript number of electrons.

Electrons in the outermost shell (highest n) are called valence electrons (or peripheral electrons).

In general, only valence electrons participate in chemical reactions because core electrons, located in inner shells, are too tightly bound.



Electron Configuration Table

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 1 1s																	He 2 1s
2	Li 1 2s	Be 2											B 1 2p	C 2	N 3	O 4	F 5	Ne 6
3	Na 1 3s	Mg 2											Al 1 3p	Si 2	P 3	S 4	Cl 5	Ar 6
4	K 1 4s	Ca 2	Sc 1 3d	Ti 2	V 3	Cr 4	Mn 5	Fe 6	Co 7	Ni 8	Cu 9	Zn 10	Ga 1 4p	Ge 2	As 3	Se 4	Br 5	Kr 6
5	Rb 1 5s	Sr 2	Y 1 4d	Zr 2	Nb 3	Mo 4	Tc 5	Ru 6	Rh 7	Pd 8	Ag 9	Cd 10	In 1 5p	Sn 2	Sb 3	Te 4	I 5	Xe 6
6	Cs 1 6s	Ba 2	La *1 5d	Hf 2	Ta 3	W 4	Re 5	Os 6	Ir 7	Pt 8	Au 9	Hg 10	Tl 1 6p	Pb 2	Bi 3	Po 4	At 5	Rn 6
7	Fr 1 7s	Ra 2	Ac** 1 6d	Rf 2	Db 3	Sg 4	Bh 5	Hs 6	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

Figure IV.14. Electrons configurations for the valance subshells of atoms

Key points to remember:

- Electrons in atoms occupy discrete energy levels described by quantum numbers (n, l, ml, m_s), which define their energy, shape, orientation, and spin.
- Atomic orbitals represent the regions in space where the probability of finding an electron is highest. The main types are s, p, d, and f orbitals.
- The Pauli Exclusion Principle states that no two electrons in an atom can have the same set of four quantum numbers.
- According to the stability principle and Klechkowski's rule, electrons occupy atomic orbitals in order of increasing energy.
- Hund's Rule states that electrons occupy degenerate orbitals singly with parallel spins before pairing occurs.
- The electronic configuration of an atom determines its chemical properties and explains the organization of elements in the periodic table.

Chapter V. Periodic Classification of the Elements

V.1 The Periodic Table of the Elements

Mendeleev's periodic table (1871) arranged chemical elements in order of increasing atomic mass. The modern periodic table, however, is organized by atomic number (Z) and electronic structure. It currently includes 118 known elements.

The table is arranged in 18 vertical columns, called groups, and 7 horizontal rows, called periods.

Using Klechkowski's rule, we can determine, for each period, the number of elements it contains as well as the outer electronic configuration of the atoms in that period.

Elements in the same column have the same outer electronic configuration and therefore exhibit similar chemical properties, which explains the term "periodic classification".

- **Groups:**
  See periodic table on Figure V.1.
  Example: Group IA (column 1)
  The Roman numeral indicates the total number of valence electrons for elements in that group.
  The letters A and B indicate the type of valence electrons involved in the elements of a group.

- Groups A (IA to VIIIA) consist of elements whose valence electrons occupy s and/or p orbitals.

- Elements in group VIIIA (group 18) have an outer electronic configuration of type $ns^2\ np^6$ (except helium, which has $1s^2$). This configuration gives them great stability; these are the noble gases. They have a complete outer shell, which explains their very low reactivity.

- Groups B (IB to VIIIB) consist of elements whose valence electrons occupy s and d orbitals → these are the transition metals. Their electronic configuration ends with $ns^2\ (n\text{-}1)\ d^{1\text{-}10}$.

- Group VIIIB consists of three neighboring columns, called “triads”. Elements in this group exhibit similar chemical behavior.
- The lanthanides (4f) and actinides (5f) (f-block) fill internal f orbitals.

Note: In modern IUPAC notation, columns are numbered 1 to 18.

IUPAC: International Union of Pure and Applied Chemistry

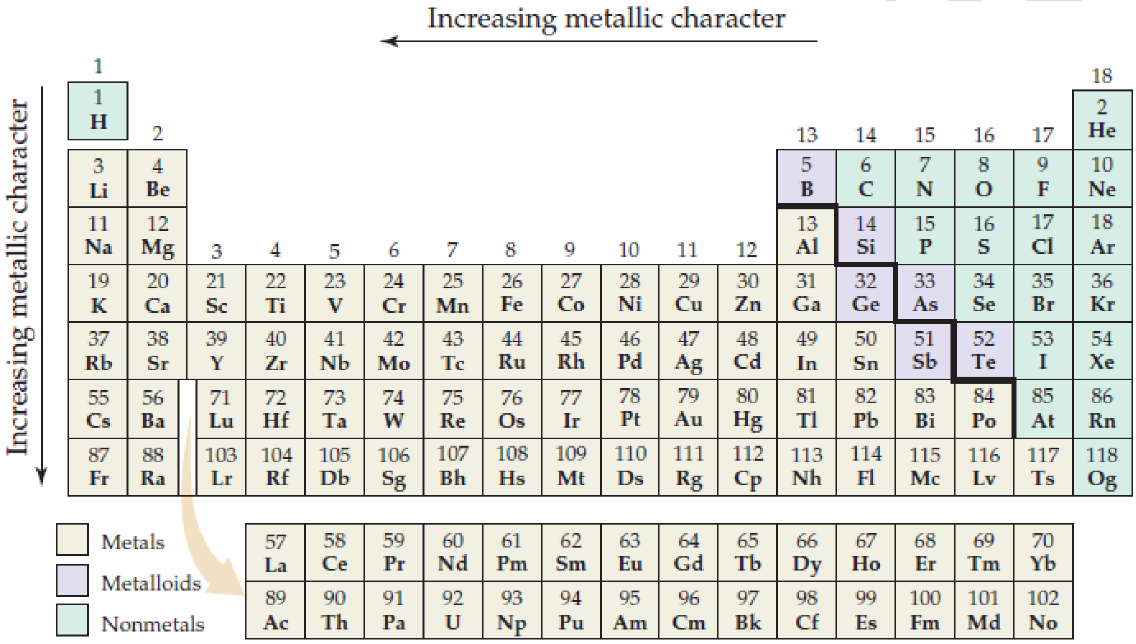


Figure V.1. Periodic Table

V.2 Periodic Properties

The properties of elements vary periodically because they depend on electronic configuration.

1. Atomic radius: the distance between the nucleus and the boundary of the electron cloud (sizes of atoms).

- Across a period (→), the radius decreases because the nuclear charge increases (e.g., Li > Be > B) (see Figure V.2).
- Down a group (↓), the radius increases because an additional electron shell is added (e.g., F < Cl < Br).

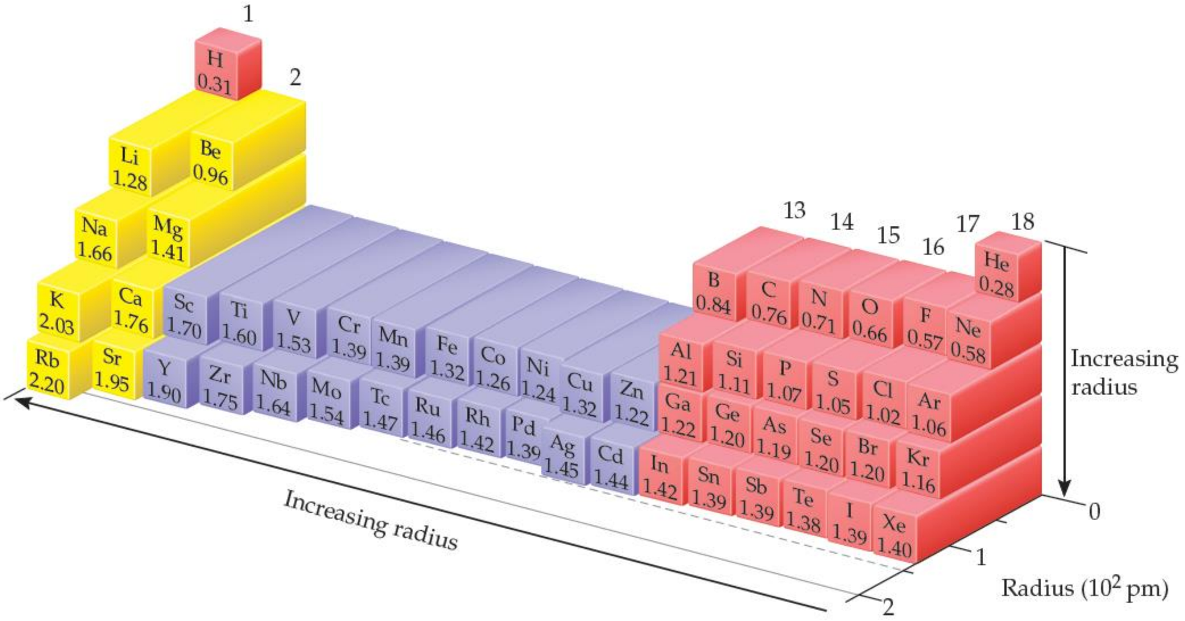


Figure V.2. Trends in bonding atomic radii for Periods 1 through 5

2. First Ionization Energy: In general, the first ionization energy, I_1, is the energy needed to remove the first electron from a neutral atom.

- It increases across a period (→) as electrons are more strongly attracted to the nucleus.
- It decreases down a group (↓) as electrons are farther from the nucleus (see Figure V.3).

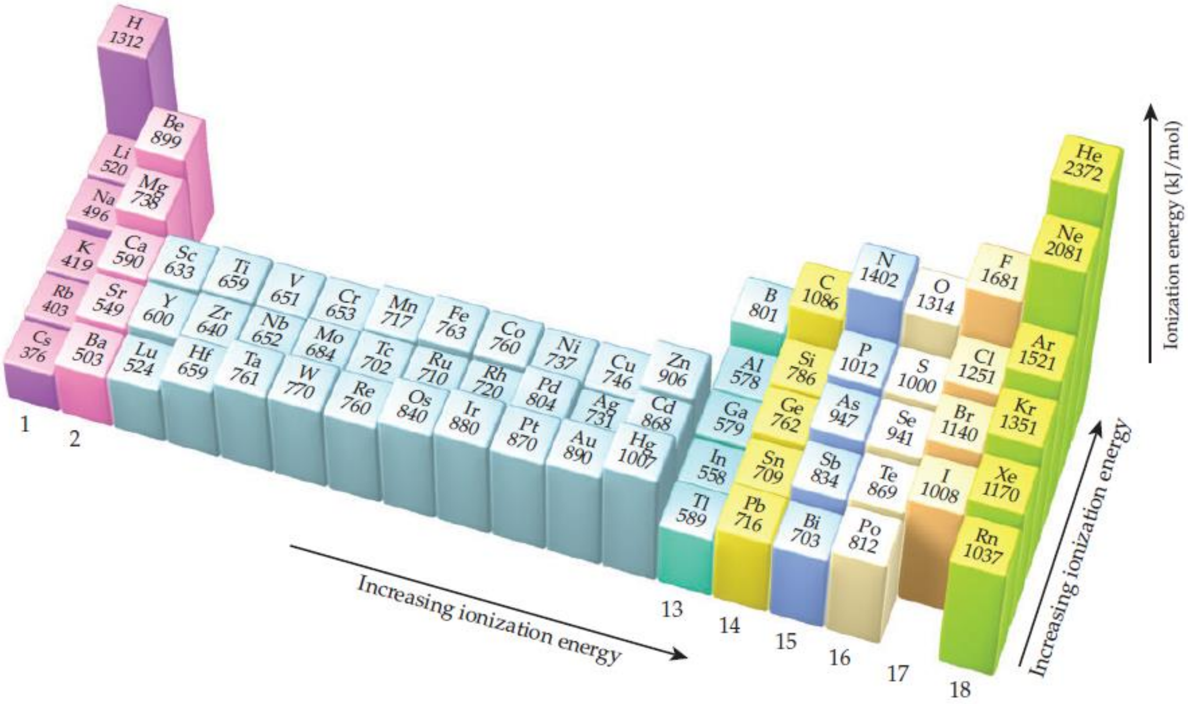


Figure V.3. The first ionization energies of the elements in kJ/mol

3. Electron Affinity: it measures the energy change that occurs when an electron is added to a gaseous atom. The greater the attraction between the atom and the added electron, the more negative its electron affinity.

- Across a period, it generally becomes more negative (atoms more readily gain an electron).
- Down a group, it becomes less negative (atoms gain electrons less easily).
- Example: Cl releases more energy than S when capturing an electron.

Key points to remember:

- Periodicity explains family trends: halogens are highly electronegative, alkali metals are highly electropositive.
- These properties allow prediction of chemical reactivity, bond types, and element behavior.

Chapter VI. Chemical Bonding and Molecular Geometry

This chapter examines, through the lens of Lewis theory, chemical bonding, with a particular focus on three-dimensional arrangement and molecular shape. Geometry of molecules plays a crucial role in the determination of both chemical and physical properties, such as density, melting points, boiling points, and chemical reactivity. It will be demonstrated that molecular shapes can be estimated with reasonable accuracy using a simple method based on Lewis structures. Nevertheless, although this theory is widely used due to its simplicity and practical value, it does not offer a detailed explanation of the fundamental processes underlying bond formation.

The Valence Shell Electron Pair Repulsion (VSEPR) model provides a reliable framework for predicting the geometry of a wide range of molecular systems. Its widespread adoption among chemists can be attributed to its ease of application and the consistency of its predictions. Although the precise role of electron-pair repulsion in determining molecular structure remains a subject of theoretical discussion, this assumption continues to serve as a useful basis for understanding and predicting molecular geometry.

VI.1 Molecular Geometry

Molecular geometry (see Figure VI.1) is the three-dimensional arrangement of atoms in a molecule.

The three-dimensional arrangement of atoms within a molecule has a direct influence on its physical and chemical behavior, including properties such as density, melting and boiling temperatures, and chemical reactivity. Precise values for bond lengths and bond angles are typically obtained through experimental measurements. Nevertheless, when the Lewis structure of a molecule or ion is known, a straightforward method allows for a reasonably accurate prediction of its overall shape. This method is based on counting the regions of electron density—often referred to as electron domains—around a central atom.

An electron domain may correspond either to a lone pair or to a bonding interaction, regardless of whether the bond is single, double, or triple. In this context, the central atom is defined as any atom that is not located at the periphery of a polyatomic species. The underlying principle of this model is that electron pairs in the valence shell tend to repel one another and adopt arrangements that minimize these repulsive interactions.

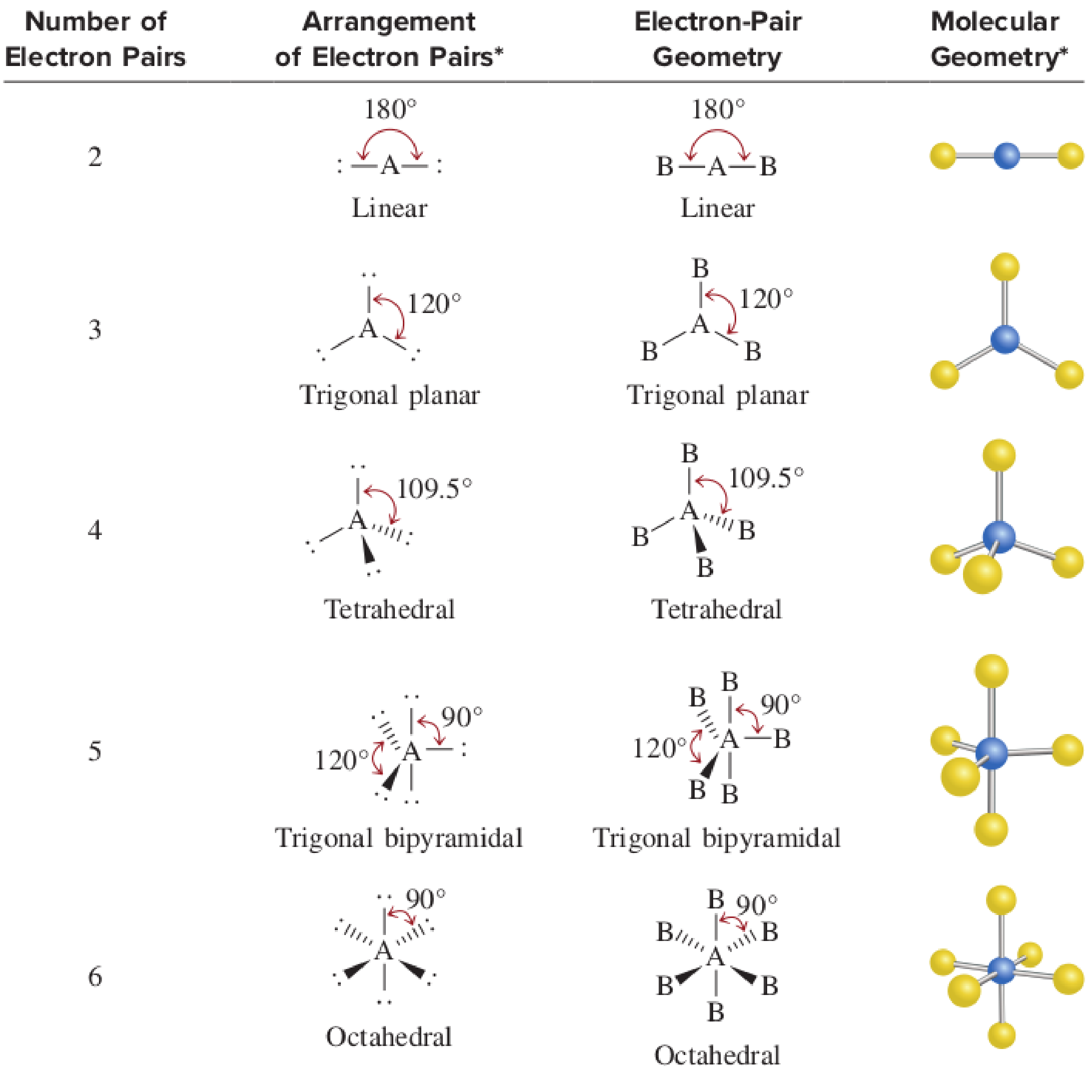
The valence shell corresponds to the outermost energy level of an atom that contains electrons. These electrons are the ones most commonly engaged in the formation of chemical bonds.

In covalent bonding, two atoms are linked through a shared pair of electrons, commonly referred to as a bonding pair. In molecules composed of several atoms, the central atom is typically connected to multiple neighboring atoms through more than one bond. As a result, the electron pairs associated with these bonds repel one another and tend to position themselves at maximum distances from each other. The spatial arrangement adopted by the molecule, defined by the relative positions of its atoms, corresponds to the configuration that reduces these repulsive interactions to a minimum. This method of describing and predicting molecular shape is known as the Valence Shell Electron Pair Repulsion (VSEPR) model, as it explains molecular geometry, by considering the electrostatic repulsion between electron pairs in the valence shell of a central atom (see Figure VI.1).

The application of the VSEPR model is based on two fundamental principles:

1. When considering repulsions between electron pairs, multiple bonds are often approximated as being equivalent to single bonds. This simplification is appropriate for qualitative descriptions of molecular shape. It is important to note, however, that multiple bonds actually occupy a greater volume of space than single bonds. Because two or three shared electron pairs are involved between the same two atoms, the associated electron density is more extensive.

2. When a molecule can be represented by multiple resonance forms, the VSEPR model may be applied using any one of these structures. In such cases, formal charges are generally omitted from the analysis.



*Bonds coming out of the page are represented as solid wedges. Bonds going into the page are represented as dashed wedges. Bonds in the plane of the page are represented as solid lines.

Figure VI.1 Electron-Pair Geometry and Molecular Geometry

1) Molecular Geometry in Species Containing Multiple Central Atoms

Up to this point, our discussion has focused on molecules that contain a single central atom. When a molecule includes more than one nonterminal atom, describing its overall shape becomes more complex. In such situations, it is often more practical to consider the local geometry around each central atom individually.

In methanol, for example, both carbon and oxygen act as central atoms. Around the carbon atom, the three C–H bonds and the C–O bond are arranged in a tetrahedral fashion, leading to bond angles close to 109°. Consequently, both the H–C–H and O–C–H angles are approximately 109°.

The oxygen atom, on the other hand, resembles the oxygen atom in a water molecule, as it is associated with two lone pairs and two bonding pairs. As a result, the H–O–C fragment adopts a bent geometry, with an H–O–C bond angle of about 105° (see Figure VI.2).

2) Guidelines for Using the VSEPR Model

After examining molecular geometries for systems with central atoms that may or may not contain lone pairs, we can outline a general procedure for applying the VSEPR model to a wide variety of molecules.

First, draw the Lewis structure of the species, focusing specifically on the electron pairs associated with the central atom—that is, the atom bonded to two or more other atoms. Next, determine the total number of electron pairs surrounding this atom, including both shared (bonding) pairs and nonbonding (lone) pairs. For the purpose of this model, multiple bonds should be counted as a single electron domain.

The overall spatial arrangement of the electron domains can then be estimated using Figure VI.1. Once this arrangement is known, the molecular shape may be identified with the aid of Figure VI.2.

When estimating bond angles, it is important to recognize that lone pairs exert stronger repulsive forces than bonding pairs, whether they interact with another lone pair or with a bonding pair. Consequently, precise bond-angle predictions become difficult when one or more lone pairs are present on the central atom.

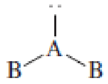
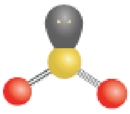
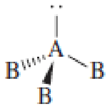
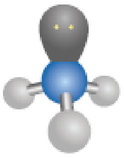
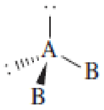
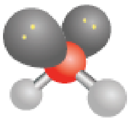
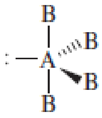
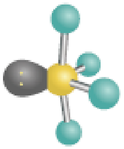
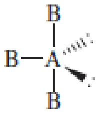
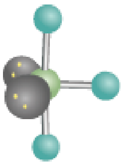
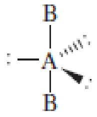
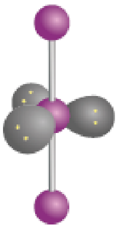
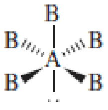
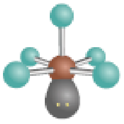
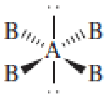
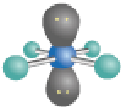
Class of Molecule	Total Number of Electron Pairs	Number of Bonding Pairs	Number of Lone Pairs	Arrangement of Electron Pairs	Geometry of Molecule or Ion	Examples
AB_2E	3	2	1	 Trigonal Planar	Bent (or V-shaped)	 SO_2
AB_3E	4	3	1	 Tetrahedral	Trigonal pyramidal	 NH_3
AB_2E_2	4	2	2	 Tetrahedral	Bent (or V-shaped)	 H_2O
AB_4E	5	4	1	 Trigonal bipyramidal	Seesaw (or distorted tetrahedron)	 SF_4
AB_3E_2	5	3	2	 Trigonal bipyramidal	T-shaped	 ClF_3
AB_2E_3	5	2	3	 Trigonal bipyramidal	Linear	 I_3^-
AB_5E	6	5	1	 Octahedral	Square pyramidal	 BrF_5
AB_4E_2	6	4	2	 Octahedral	Square planar	 XeF_4

Figure VI.2. Molecules and ions: electron – pair geometry and molecular geometry

VI.2 Dipole Moments

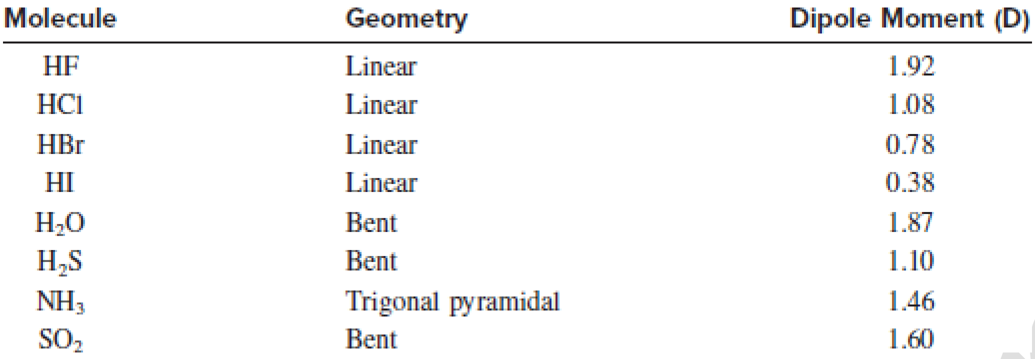
The dipole moment of a molecule quantifies the separation of electrical charge in compounds containing atoms with differing electronegativities (see Figure VI.3). A molecule's overall dipole moment results from the vector sum of the individual bond dipoles. Measuring the dipole moment can therefore provide valuable insight into the molecule's geometry.

For instance, hydrogen fluoride (HF) is a covalent molecule featuring a polar bond. Because fluorine is more electronegative than hydrogen, electron density is drawn toward the fluorine atom. This shift in electron density is often represented in the Lewis structure by a crossed arrow pointing toward the more electronegative atom, indicating the direction of the electron shift.

The resulting separation of charge can be depicted as

δ (delta): a partial charge.

In an electric field, this separation of charges can be confirmed. HF molecules orient their negative ends toward the positive plate, when the field is turned on and their positive ends toward the negative plate. Experimentally, this alignment of molecules can be detected.



Molecule	Geometry	Dipole Moment (D)
HF	Linear	1.92
HCl	Linear	1.08
HBr	Linear	0.78
HI	Linear	0.38
H_2O	Bent	1.87
H_2S	Bent	1.10
NH_3	Trigonal pyramidal	1.46
SO_2	Bent	1.60

Figure VI.3. Dipole moments of some polar molecules

VI.3 Valence Bond Theory

There are two main quantum mechanical approaches for describing how covalent bonds form: valence bond (VB) theory and molecular orbital (MO) theory. According to VB theory, electrons in a molecule occupy atomic orbitals associated with the individual atoms. A stable molecule is achieved when the system reaches a minimum in potential energy (see Figure VI.4).

The VSEPR model, which is largely based on Lewis structures, provides a relatively simple and practical way to predict molecular geometry. However, as mentioned earlier, Lewis theory does not explain why chemical bonds actually exist. While linking covalent bond formation to electron pairing was a step forward, it does not fully account for all observations. For example, Lewis theory represents the single bond in H_2 and the single bond in F_2 in the same way, as simply a pair of electrons. Yet these two molecules differ significantly in bond energy and bond length: H_2 has a bond enthalpy of 436.4 kJ/mol and a bond length of 74 pm, while F_2 has a bond enthalpy of 150.6 kJ/mol and a bond length of 142 pm. Such differences cannot be explained by Lewis theory alone.

To gain a more complete understanding of bond formation and molecular geometry, we turn to quantum mechanics. Currently, VB theory and MO theory are the two primary quantum mechanical frameworks used to describe covalent bonding and the electronic structure of molecules. VB theory preserves the concept of electrons residing in atomic orbitals and emphasizes the role of individual atoms in bond formation. In contrast, MO theory considers that atomic orbitals combine to form molecular orbitals that extend over the entire molecule.

While neither approach fully accounts for every aspect of bonding, both have provided valuable insights into many observed molecular properties.

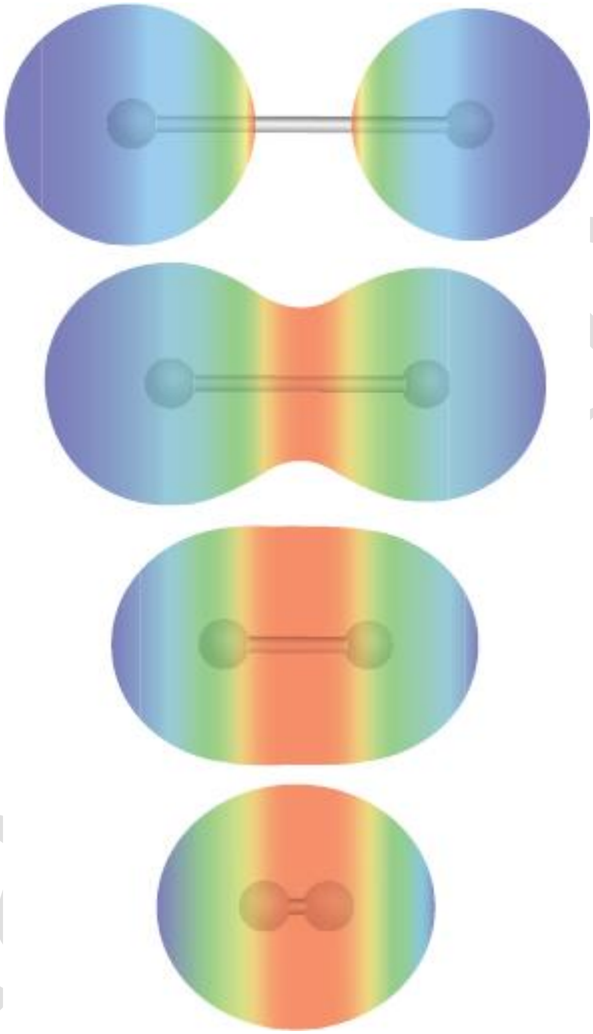


Figure VI.4. From top to bottom: As the two hydrogen atoms come closer, their 1s orbitals start to overlap, and each electron experiences the attraction of both nuclei. This leads to an increased concentration of electron density between the two atoms (indicated in red). Eventually, this interaction results in the formation of a stable H_2 molecule, with an equilibrium bond length of approximately 74 pm.

Valence bond theory provides a more detailed understanding of how chemical bonds form compared with Lewis theory. According to valence bond theory, a stable molecule is produced when the interacting atoms reach a configuration that minimizes the system's potential energy, whereas Lewis theory does not take the energy changes associated with bond formation into account.

VI.4 Hybridization of Atomic Orbitals

In valence bond theory, hybrid orbitals are produced when atomic orbitals from the same atom combine and rearrange. All resulting hybrid orbitals have equal energy and electron density, and the total number of hybrid orbitals equals the number of atomic orbitals involved in the combination. Expansion of the valence shell can be understood by considering the hybridization of s, p, and d orbitals.

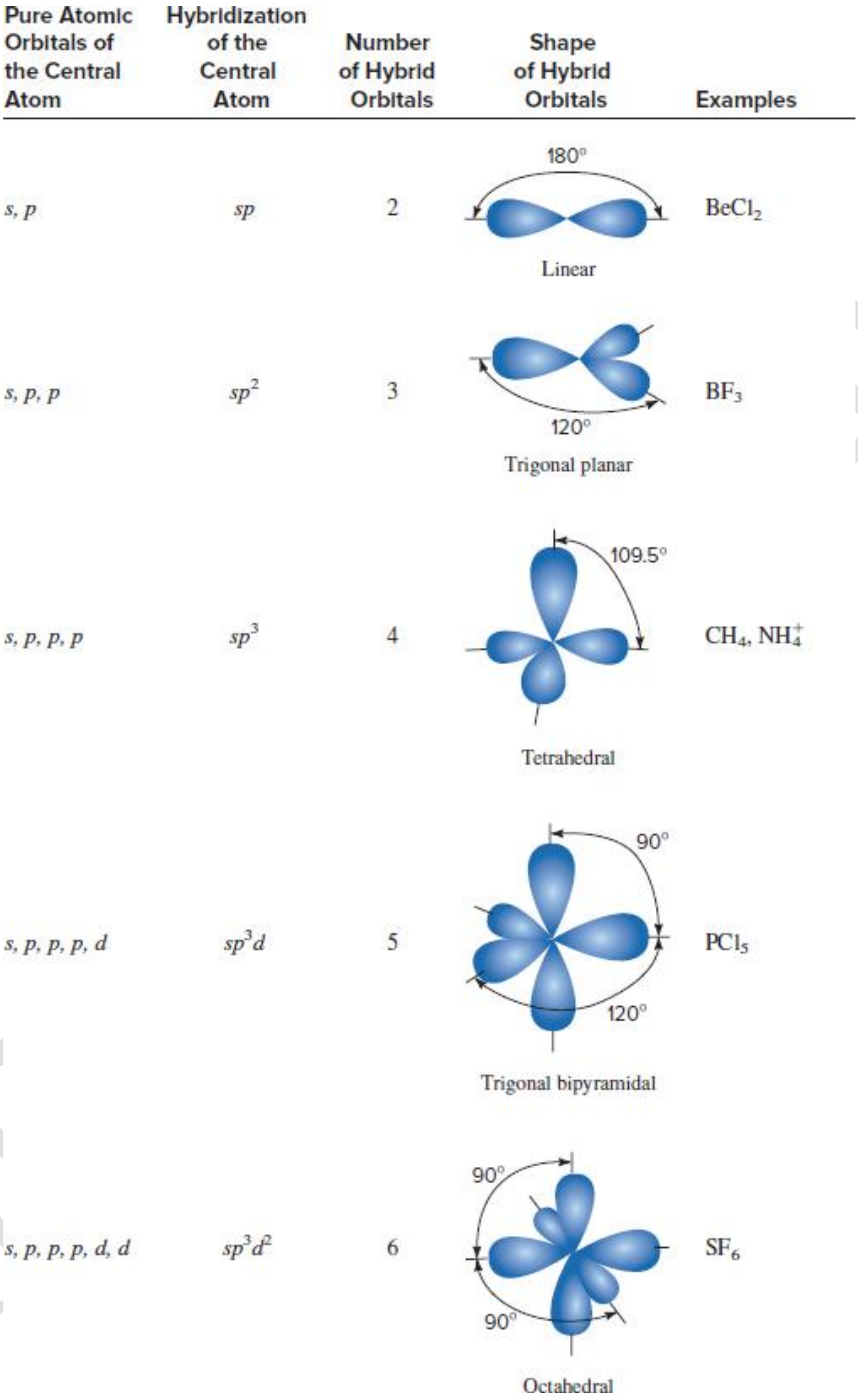
- In **sp hybridization**, two hybrid orbitals are aligned linearly.
- In **sp^2 hybridization**, three hybrid orbitals point toward the corners of an equilateral triangle.
- In **sp^3 hybridization**, four hybrid orbitals are directed toward the corners of a tetrahedron.
- In **sp^3d hybridization**, five hybrid orbitals are arranged in a trigonal bipyramidal geometry.
- In **sp^3d^2 hybridization**, six hybrid orbitals occupy the corners of an octahedron.

Hybridization is a theoretical construct applied to atoms within molecules rather than isolated atoms. It provides a model for understanding covalent bonding :

- Hybridization involves the mixing of at least two atomic orbitals of different types, such as s and p orbitals. As a result, a hybrid orbital is distinct in shape from a pure atomic orbital.
- The number of hybrid orbitals formed always equals the number of atomic orbitals that combine.
- Although energy is required to produce hybrid orbitals, this energy is more than compensated for by the stability gained during bond formation.
- Covalent bonds in molecules and polyatomic ions are formed either by overlap between hybrid orbitals or between hybrid and unhybridized orbitals. This approach

remains within the framework of valence bond theory, as electrons are still considered to occupy atomic-based orbitals, though hybridized.

Figure VI.5 illustrates the geometries of sp, sp^2, sp^3, and other types of hybridization.



Pure Atomic Orbitals of the Central Atom	Hybridization of the Central Atom	Number of Hybrid Orbitals	Shape of Hybrid Orbitals	Examples
s, p	*sp*	2	180° Linear	$BeCl_2$
s, p, p	sp^2	3	120° Trigonal planar	BF_3
s, p, p, p	sp^3	4	109.5° Tetrahedral	CH_4, NH_4^+
s, p, p, p, d	sp^3d	5	90° 120° Trigonal bipyramidal	PCl_5
s, p, p, p, d, d	sp^3d^2	6	90° 90° Octahedral	SF_6

Figure VI.5. Important hybrid orbitals and their shapes

Procedure for hybridizing atomic orbitals:

Hybridization can be thought of as an extension of Lewis theory and the VSEPR model, providing a way to link electron pair arrangements to the shapes of molecules. Assigning the correct hybridization to a central atom requires understanding the geometry of the molecule. The process can be broken down into clear steps:

Step 1: Draw the Lewis structure

Begin by writing the Lewis structure of the molecule. This allows you to identify which atoms are connected, where the bonds are, and how many lone pairs of electrons each atom possesses. The Lewis structure provides the foundation for understanding the distribution of electrons around the central atom.

Step 2: Determine the electron pair arrangement using VSEPR

Next, apply the VSEPR (Valence Shell Electron Pair Repulsion) model to predict the spatial arrangement of all electron pairs—both bonding and nonbonding—around the central atom. VSEPR theory assumes that electron pairs repel each other and will arrange themselves to be as far apart as possible, which determines the general shape of the molecule (see Figure VI.6). By knowing whether the central atom has two, three, four, five, or six regions of electron density, you can predict whether the arrangement is linear, trigonal planar, tetrahedral, trigonal bipyramidal, or octahedral.

Step 3: Assign the hybridization

Once the geometry is known, the hybridization of the central atom can be inferred by comparing the electron pair arrangement with the geometries of hybrid orbitals (see Figures VI.5 and VI.7). For example:

- Two regions of electron density → sp hybridization (linear geometry)
- Three regions → sp^2 hybridization (trigonal planar)
- Four regions → sp^3 hybridization (tetrahedral)
- Five regions → sp^3d hybridization (trigonal bipyramidal)
- Six regions → sp^3d^2 hybridization (octahedral)

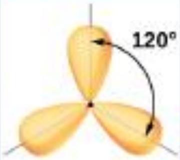
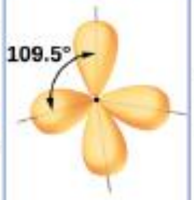
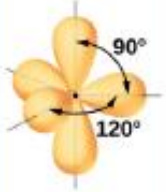
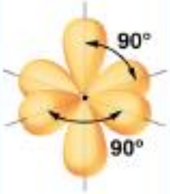
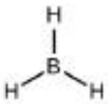
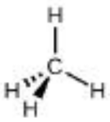
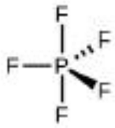
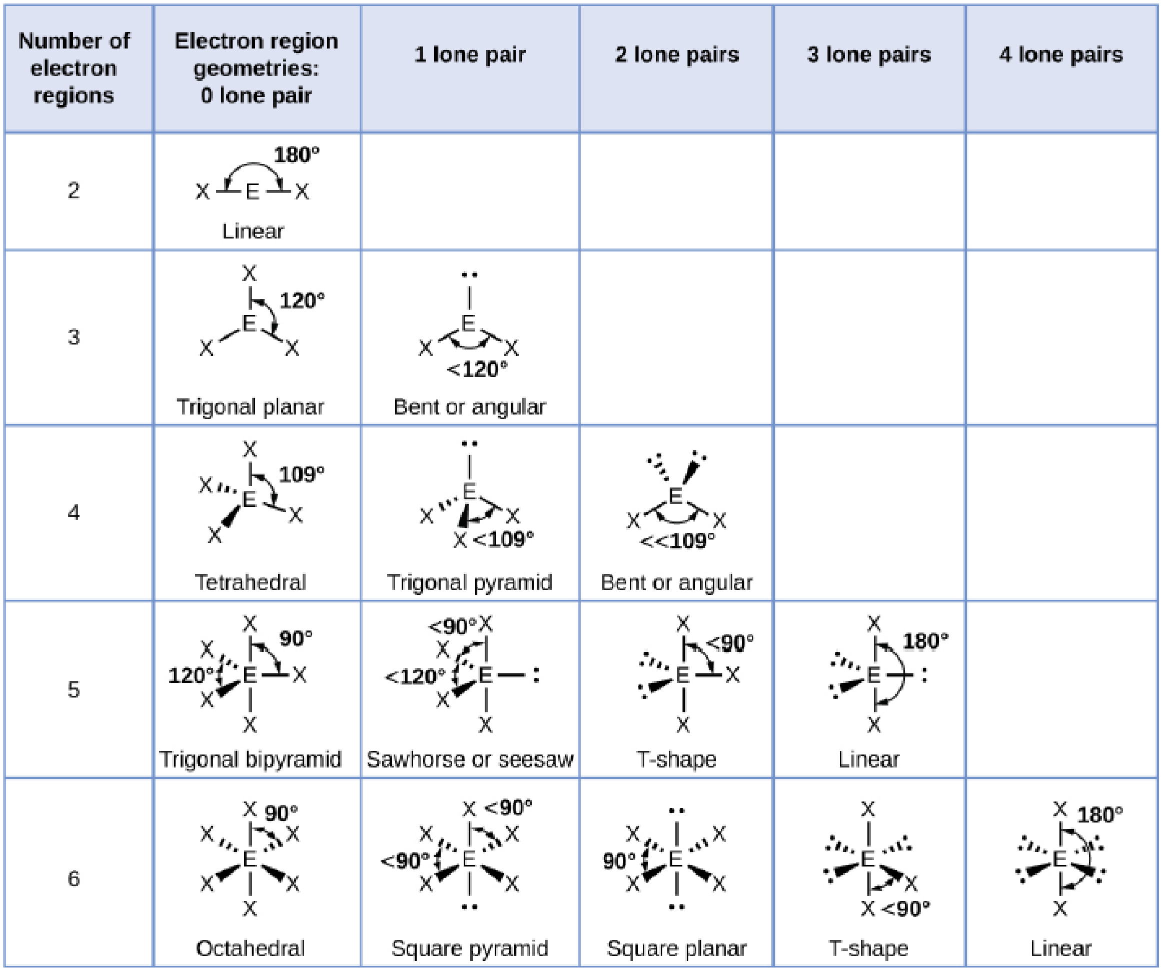
Number of regions	Two regions of high electron density (bonds and/or unshared pairs)	Three regions of high electron density (bonds and/or unshared pairs)	Four regions of high electron density (bonds and/or unshared pairs)	Five regions of high electron density (bonds and/or unshared pairs)	Six regions of high electron density (bonds and/or unshared pairs)
Spatial arrangement					
Line-dash-wedge notation	H—Be—H				
Electron region geometry	Linear; 180° angle	Trigonal planar; all angles 120°	Tetrahedral; all angles 109.5°	Trigonal bipyramidal; angles of 90° or 120° An attached atom may be equatorial (in the plane of the triangle) or axial (above or below the plane of the triangle).	Octahedral; all angles 90° or 180°

Figure VI.6. VSEPR theory



Number of electron regions	Electron region geometries: 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	Linear				
3	Trigonal planar	Bent or angular			
4	Tetrahedral	Trigonal pyramid	Bent or angular		
5	Trigonal bipyramid	Sawhorse or seesaw	T-shape	Linear	
6	Octahedral	Square pyramid	Square planar	T-shape	Linear

Figure VI.7. Electron-pair geometries

Key points to remember:

- MOLECULAR SHAPES: The three-dimensional shapes and sizes of molecules are determined by their bond angles and bond lengths. Molecules with a central atom A surrounded by n atoms B, denoted ABn, adopt a number of different geometric shapes, depending on the value of n and on the particular atoms involved. In the overwhelming majority of cases, these geometries are related to five basic shapes (linear, trigonal pyramidal, tetrahedral, trigonal bipyramidal, and octahedral).

- THE VSEPR MODEL: The valence-shell electron-pair repulsion (VSEPR) model rationalizes molecular geometries based on the repulsions between electron domains, which are regions about a central atom in which electrons are likely to be found. Bonding pairs of electrons, which are those involved in making bonds, and nonbonding pairs of electrons, also called lone pairs, both create electron domains around an atom. According to the VSEPR model, electron domains orient themselves to minimize electrostatic repulsions; that is, they remain as far apart as possible. Electron domains from nonbonding pairs exert slightly greater repulsions than those from bonding pairs, which leads to certain preferred positions for nonbonding pairs and to the departure of bond angles from idealized values. Electron domains from multiple bonds exert slightly greater repulsions than those from single bonds. The arrangement of electron domains around a central atom is called the electron-domain geometry; the arrangement of atoms is called the molecular geometry.

- MOLECULAR POLARITY: The dipole moment of a polyatomic molecule depends on the vector sum of the dipole moments associated with the individual bonds, called the bond dipoles. Certain molecular shapes, such as linear AB_2 and trigonal planar AB_3, lead to cancellation of the bond dipoles, producing a nonpolar molecule, which is one whose overall dipole moment is zero. In other shapes, such as bent AB_2 and trigonal pyramidal AB_3, the bond dipoles do not cancel and the molecule will be polar (that is, it will have a nonzero dipole moment).

- COVALENT BONDING AND ORBITAL OVERLAP: Valence-bond theory is an extension of Lewis's notion of electron-pair bonds. In valence-bond theory, covalent bonds are formed when atomic orbitals on neighboring atoms overlap one another. The overlap region is one of greater stability for the two electrons because of their simultaneous attraction to two nuclei. The greater the overlap between two orbitals, the stronger the bond that is formed.

- HYBRID ORBITALS: To extend the ideas of valence- bond theory to polyatomic molecules, we must envision mixing s and p orbitals to form hybrid orbitals. The process of hybridization leads to hybrid atomic orbitals that have a large lobe directed to overlap with orbitals on another atom to make a bond. Hybrid orbitals can also accommodate nonbonding pairs. A particular mode of hybridization can be associated with each of three common electron-domain geometries 1linear = sp; trigonal planar = sp^2; tetrahedral = sp^3. The bonding in hypervalent molecules—those with more than an octet of electrons—is not as readily discussed in terms of hybrid orbitals.

Lewis dot diagram

- A Lewis dot diagram is a visual tool that shows only an atom's outer-shell (valence) electrons.

- The inner, or core, electrons are not drawn; instead, the element is represented by its chemical symbol, with dots placed around it to indicate valence electrons.

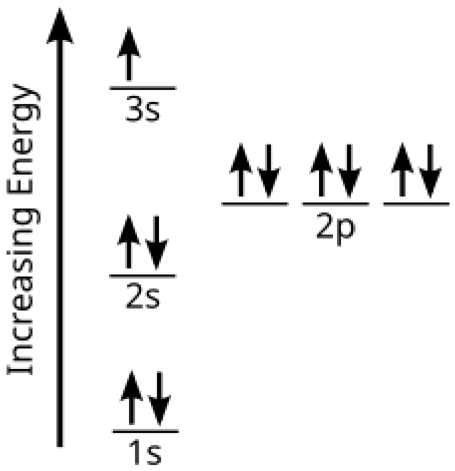
- An atom can display up to eight valence electrons in a Lewis dot diagram, with a maximum of two electrons positioned on each side of the symbol (top, bottom, left, and right).

- Valence electrons are placed according to orbital filling rules: the s orbital is filled first, followed by placing one electron in each p orbital before any p orbital receives a second electron.

- Elements that belong to the same group on the periodic table have identical numbers of valence electrons, so their Lewis dot diagrams contain the same number of dots, differing only in the element symbol.

Valence Electrons (A Simplified Way)

Core electrons include all electrons except those in the outermost shell, known as valence electrons, which drive chemical bonding and reactivity. For sodium, the electron configuration exemplifies this arrangement.
Core electrons occupy inner energy levels and remain stable during reactions, while valence electrons—found in the highest occupied orbital—determine an atom's reactivity. This distinction is fundamental in understanding periodic trends and molecular interactions.



The electron arrangement for sodium is $1s^2 2s^2 2p^6 3s^1$, where the inner electrons ($1s^2 2s^2 2p^6$) form the stable core, and the single outer electron ($3s^1$) acts as the valence electron.

Visualizing Valence Electrons

Lewis introduced a simple graphical method called Lewis dot structures to depict only the reactive valence electrons, ignoring the inert core ones that don't participate in bonding. For sodium, this means showing the Na symbol with one dot nearby, symbolizing its lone valence electron available for reactions.

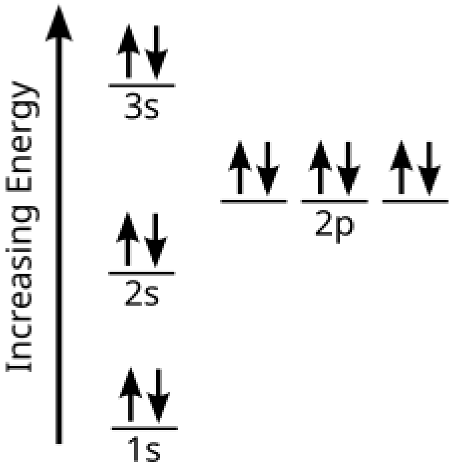
Drawing Process

Imagine the atomic symbol inside a square frame, with each of the four edges standing for the s orbital or one of the three p orbitals in the outer shell. Place the first valence electron on the s side; for sodium, that's it—just one dot for the $3s^1$ electron. Additional electrons (in multi-electron atoms) fill p sides singly before pairing, mirroring orbital filling rules.

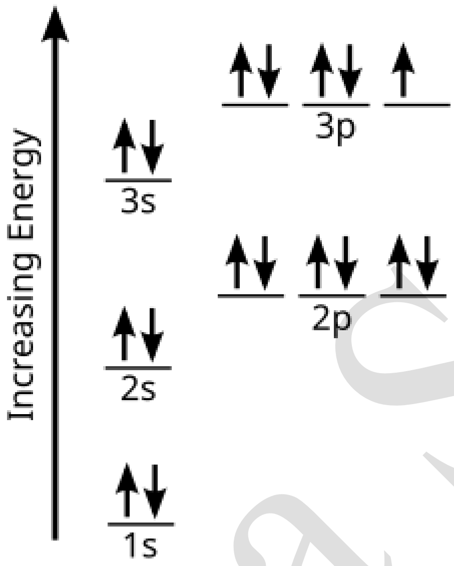
For magnesium: core electrons $1s^2 2s^2 2p^6$

To draw the Lewis electron dot diagram: we place the valence electrons around the sides of the box with each side representing an orbital in the outermost energy level. How many valence electrons does magnesium have? There are 2 valence electrons, $3s^2$.

The Lewis electron dot formula for magnesium is:



For chlorine, the electron configuration is:



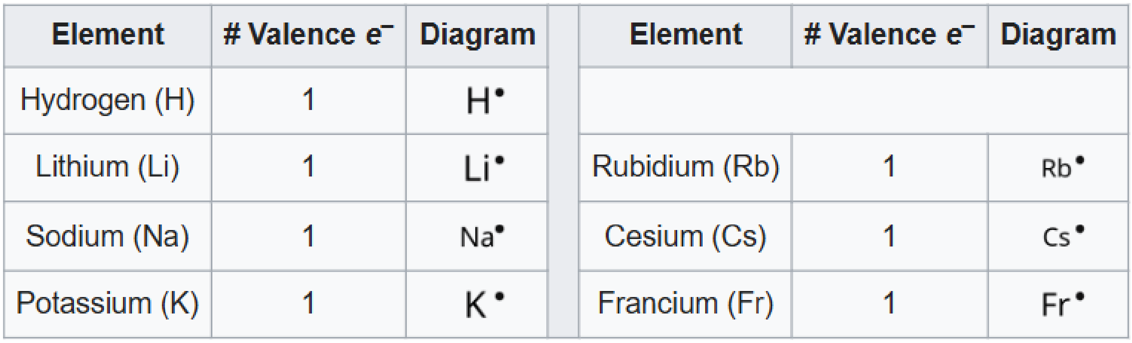
The electron configuration of chlorine: $1s^2 2s^2 2p^6 3s^2 3p^5$. The core electrons: $1s^2 2s^2 2p^6$ while the valence electrons would be in the third shell (or where $n = 3$). Therefore, chlorine has 7 valence electrons.

The electrons are in groups of two. Think of the chlorine in a box and the box has 4 sides. Each side can have 2 electrons on it. Thus, there can be a maximum of 2 × 4 = 8 electrons, normally, on any Lewis electron dot diagram.

Important

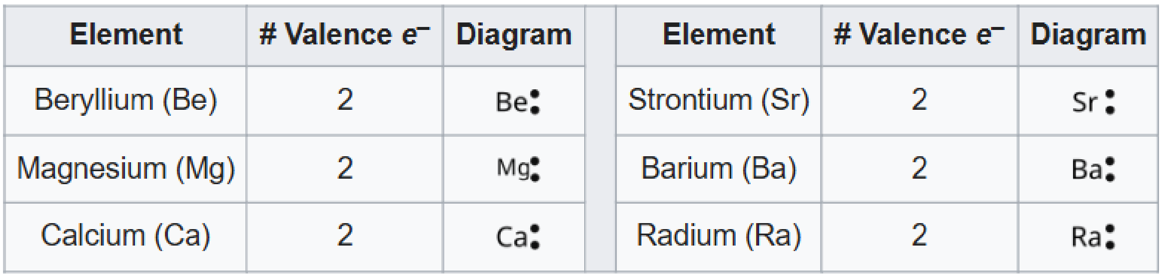
All the elements in a column have the same electron dot diagram (see table 1 and 2).

Table 1. Dot diagram for column 1.



Element	# Valence e⁻	Diagram	Element	# Valence e⁻	Diagram
Hydrogen (H)	1	H•			
Lithium (Li)	1	Li•	Rubidium (Rb)	1	Rb•
Sodium (Na)	1	Na•	Cesium (Cs)	1	Cs•
Potassium (K)	1	K•	Francium (Fr)	1	Fr•

Table 2. Dot diagram for column 2.



Element	# Valence e⁻	Diagram	Element	# Valence e⁻	Diagram
Beryllium (Be)	2	Be:	Strontium (Sr)	2	Sr:
Magnesium (Mg)	2	Mg:	Barium (Ba)	2	Ba:
Calcium (Ca)	2	Ca:	Radium (Ra)	2	Ra:

Inequivalent resonance structures

Resonance principles also extend to structures that are not energetically equivalent. These different forms usually do not contribute equally, since they often have different stability levels. In such situations, the actual molecular structure corresponds to a weighted combination that favors the lower-energy contributors. The cyanate ion (OCN^{-}) provides a clear illustration of this idea. It is possible to draw three distinct octet-based Lewis structures for this ion.

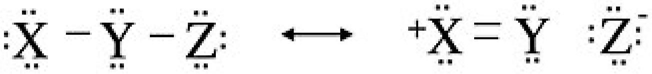
$$^{-}O-C\equiv N \longleftrightarrow O{=}C{=}N^{-} \longleftrightarrow {}^{+}O\equiv C-N^{2-}$$

Among these, the first two can be considered acceptable resonance contributors. However, the structure in which the negative formal charge is located on oxygen is expected to be the most significant, because oxygen is the most electronegative atom involved. The third structure is a poor contributor, as it introduces large formal charges and places a positive charge on oxygen, which is energetically unfavorable. Consequently, the true electronic structure of the cyanate ion is best described as a weighted average dominated by the first two resonance forms.

No-bond resonance

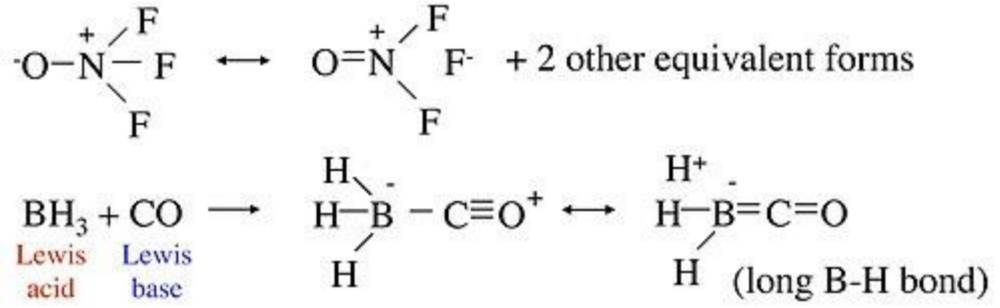
A particularly interesting type of nonequivalent resonance involves structures where the bond order between two atoms is zero. This phenomenon, known as no-bond resonance, plays an important role in explaining bonding behavior in many compounds that contain halogens or hydrogen.

To understand this concept, consider a general molecule X–Y–Z, where Z represents a highly electronegative atom such as fluorine. If the bonding electrons between Y and Z are shifted entirely onto Z, and a lone pair from X is used to form a bond with Y, a new resonance form is produced. In this alternative structure, all atoms still satisfy the octet rule, but there is no direct bond between Y and Z.



It is crucial to note that this no-bond arrangement is only one of several resonance contributors. As a result, the Z atom remains partially bonded to Y in the real molecule. If both resonance forms contribute equally, the bond between X and Y would have a bond order of approximately 1.5, while the Y–Z interaction would have a bond order of about 0.5. This reduced bond order explains why the Y–Z bond would be longer than a typical single bond and more susceptible to breaking.

Several real molecules demonstrate this effect. For example, in ONF_3, the N–F bond is noticeably longer than the corresponding bond in NF_3, where the bond order is one. This difference can be accounted for by including no-bond resonance structures. Likewise, in the Lewis acid–base complex formed between BH_3 and CO, unusually long B–H bonds can be rationalized by resonance forms in which the hydrogen atoms carry partial positive charge.



VI.5 Metallic Bonding

Metallic bonding is a fundamental type of chemical bonding that occurs in metals and alloys. It is responsible for many characteristic physical properties of metals, such as electrical conductivity, thermal conductivity, malleability, ductility, and metallic luster. Unlike ionic or covalent bonding, metallic bonding involves the collective sharing of delocalized valence electrons among a large number of metal atoms.

Understanding metallic bonding is essential for explaining why metals behave differently from nonmetals and why they are widely used in construction, electronics, transportation, and technology.

1. Atomic Structure of Metals

Most metal atoms have:

- Low ionization energies, meaning they lose valence electrons easily
- One to three valence electrons
- Relatively large atomic radii

When metal atoms come together in a solid, their outer orbitals overlap. Instead of forming localized bonds between specific pairs of atoms, the valence electrons become **delocalized**, meaning they are no longer associated with a single atom.

As a result:

- Metal atoms lose control of their valence electrons
- The atoms become **positive metal ions (cations)**
- The electrons move freely throughout the entire structure

2. The Electron Sea Model

The most widely accepted model for metallic bonding is the **electron sea model**.

Key features of the model:

- Metal cations are arranged in a regular, three-dimensional lattice
- Valence electrons form a “sea” of delocalized electrons
- These electrons move freely between the metal ions

- The bonding force is the electrostatic attraction between:
    - Positively charged metal ions
    - Negatively charged delocalized electrons

This attraction holds the metal together and forms the metallic bond.

Importantly, metallic bonding is **non-directional**, meaning the bonding force acts equally in all directions. This is very different from covalent bonding, which is directional.

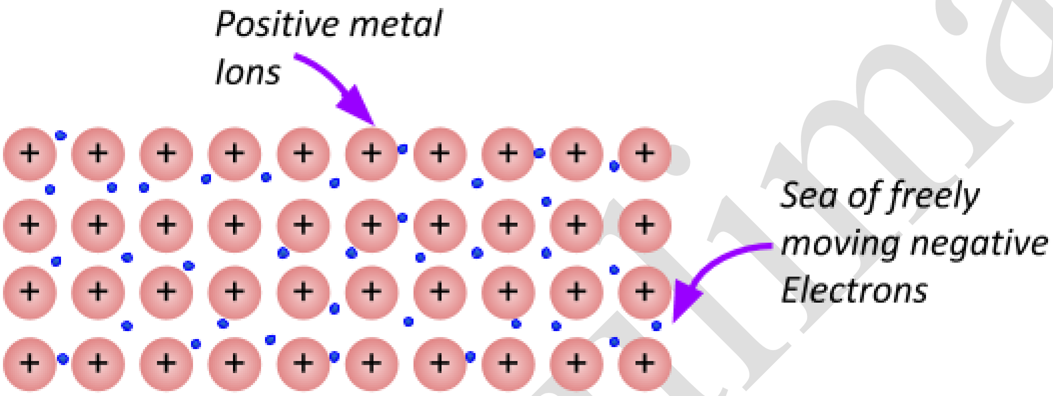


Figure VI.8. Electron Sea Model

3. Strength of Metallic Bonds

The strength of metallic bonding varies between different metals and depends on several factors:

a) Number of delocalized electrons

- More valence electrons → stronger metallic bonding
- Example: aluminum (3 valence electrons) forms stronger metallic bonds than sodium (1 valence electron)

b) Charge of metal ions

- Higher positive charge increases attraction to electrons

c) Size of metal ions

  - Smaller ions allow electrons to stay closer, strengthening attraction

Transition metals typically have strong metallic bonds because they contribute more delocalized electrons and have smaller ionic radii compared to alkali metals.

4. Physical Properties Explained by Metallic Bonding

a) Electrical Conductivity

Because delocalized electrons can move freely, metals conduct electricity efficiently. When an electric potential is applied, electrons flow through the lattice, carrying charge.

This property makes metals essential for:

- Electrical wires
- Circuits
- Electronic components

b) Thermal Conductivity

Metals are good conductors of heat because:

- Delocalized electrons transfer kinetic energy rapidly
- Vibrations move efficiently through the lattice

This explains why metals are used in cookware and heat exchangers.

c) Malleability

Malleability is the ability of a material to be hammered into thin sheets.

In metals:

- Layers of metal ions can slide over each other

- The electron sea adjusts without breaking the bond
- The electrostatic attraction remains intact

This property would not be possible with ionic solids, which fracture when layers shift.

d) Ductility

Ductility is the ability to be drawn into wires.

Because metallic bonds are non-directional:

- Atoms rearrange without bond failure
- The structure remains stable during stretching

This explains why copper and aluminum are ideal for electrical wiring.

Figure VI.9. Properties of metals

5. Metallic Luster

Metallic luster refers to the shiny appearance of metals.

Delocalized electrons:

- Absorb and re-emit light
- Reflect visible wavelengths efficiently

This interaction with light produces the characteristic shine of metal surfaces.

6. Metallic Bonding vs Other Types of Bonding

Metallic bonding is unique because electrons are shared **collectively** rather than between specific atoms.

Table 3. Metallic bond vs ionic and covalent bonds

Bond Type	**Electron Behavior**	**Structure**	**Typical Properties**
Metallic	Delocalized	Lattice of cations + electron sea	Conductive, malleable
Ionic	Transferred	Crystal lattice of ions	Brittle, high melting point
Covalent	Shared (localized)	Molecules or networks	Poor conductivity

7. Alloys and Metallic Bonding

An **alloy** is a mixture of two or more elements, at least one of which is a metal.

Metallic bonding still operates in alloys, but the presence of different-sized atoms can:

- Distort the lattice
- Reduce the ability of layers to slide
- Increase strength and hardness

Examples:

- Steel (iron + carbon)
- Brass (copper + zinc)
- Bronze (copper + tin)

Alloys are often stronger than pure metals due to these lattice disruptions.

Figure VI.10. Applications of common Alloys

Metallic bonding is the force that holds metal atoms together through the attraction between positive metal ions and a sea of delocalized electrons. This bonding model explains the defining properties of metals, including conductivity, malleability, ductility, and luster.

Because of metallic bonding, metals are indispensable materials in modern society, forming the foundation of engineering, electronics, and industrial technology.

VI.6 Intermolecular Attractive Forces

Matter exists in different physical states—solid, liquid, and gaseous—depending on the nature and strength of interactions between its constituent particles. While **chemical bonds** (covalent, ionic, metallic) ensure the internal cohesion of molecules or crystal lattices, a different category of interactions governs the behavior **between** distinct molecules. These interactions are known as **intermolecular attractive forces**.

Although intermolecular forces are significantly weaker than chemical bonds, they play a decisive role in determining the macroscopic properties of substances. Phenomena such as boiling and melting points, solubility, viscosity, surface tension, and even biological structure and function cannot be explained without considering these forces.

1. Nature of Intermolecular Forces

a) Definition

Intermolecular forces are interactions that occur between separate molecules, atoms, or ions without leading to the formation of chemical bonds. These forces arise primarily from **electrostatic interactions** involving permanent or temporary charge distributions.
Unlike intramolecular bonds, intermolecular forces do not involve electron sharing or transfer; instead, they result from subtle attractions between partial charges.

b) Energy Scale

To appreciate their relative weakness, it is useful to compare typical interaction energies:

- Covalent bond energy: 200 – 800 $kJ\cdot mol^{-1}$
- Intermolecular interaction energy: 1 – 50 $kJ\cdot mol^{-1}$

Despite their low individual strength, the collective effect of intermolecular forces can be substantial, especially in condensed phases.

c) Common Features

Intermolecular forces generally exhibit the following characteristics:

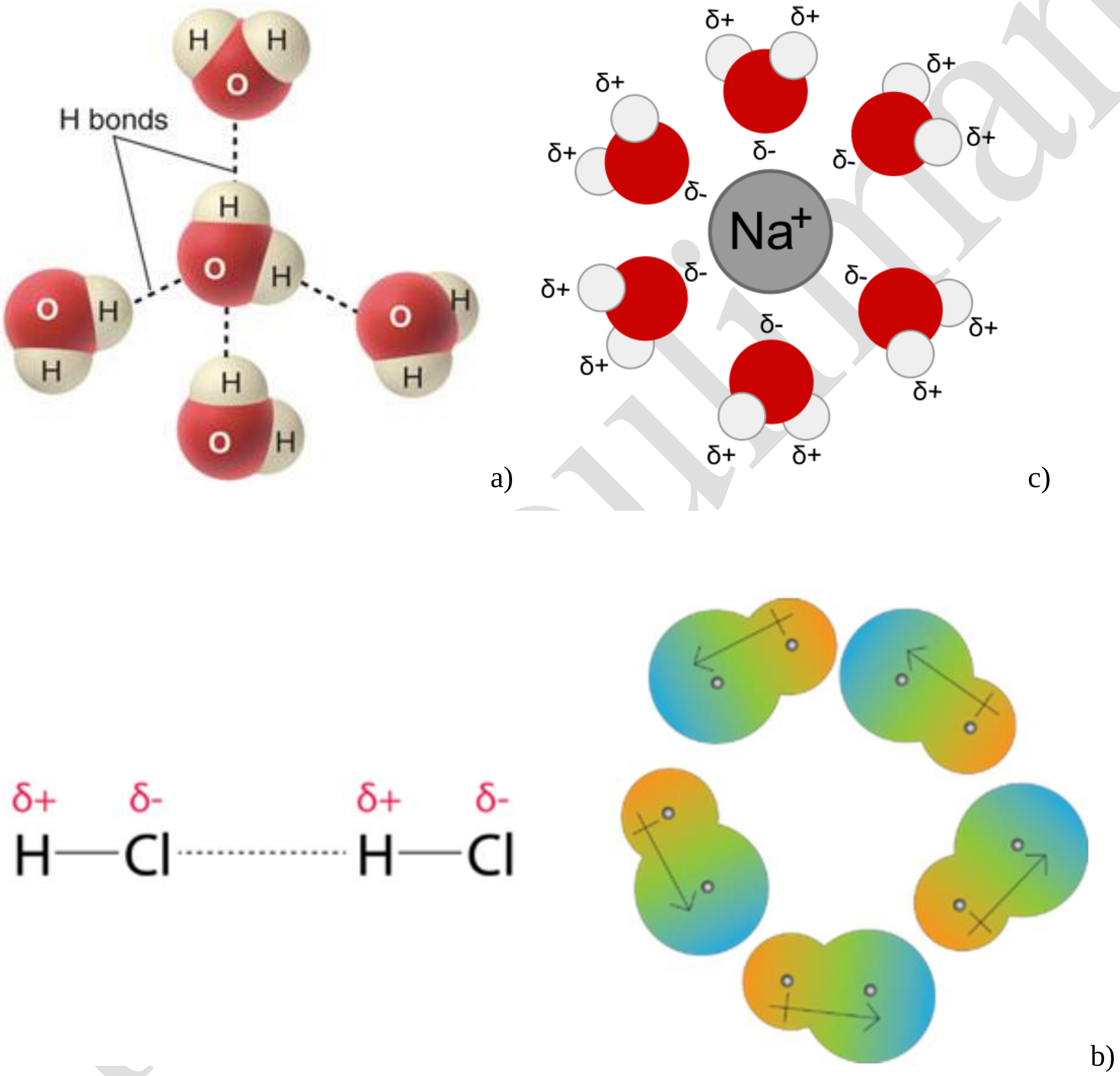
- short-range nature;
- strong dependence on intermolecular distance;
- reversibility;
- sensitivity to temperature and molecular orientation.

2. Classification of Intermolecular Interactions

Intermolecular attractive forces can be grouped into three major categories:

1. Van der Waals interactions;
2. Hydrogen bonding;
3. Ion–dipole interactions.

Each category originates from electrostatic effects but differs in strength, directionality, and physical consequences.



a) c) b)

Figure VI.10. a) Hydrogen bonding b) Van der Waals interactions
c) Ion–dipole interactions

3. Van der Waals Interactions

The term **Van der Waals forces** encompasses a set of weak interactions present between neutral molecules. These forces are universal and contribute to molecular cohesion even in nonpolar substances.

Physical Origin

Van der Waals interactions arise from interactions between molecular dipoles—either permanent or induced—created by uneven electron distribution.

Three main types are distinguished:

- dipole–dipole interactions;
- dipole–induced dipole interactions;
- dispersion forces.

4. Dipole–Dipole Interactions

a) Permanent Molecular Dipoles

A molecule is considered **polar** when it possesses a permanent electric dipole moment. This condition depends on two factors:

- differences in electronegativity between bonded atoms;
- molecular geometry that prevents charge cancellation.

Typical examples include hydrogen chloride (HCl), water (H_2O), and ammonia (NH_3).

b) Interaction Mechanism

Dipole–dipole interactions occur when the positive end of one polar molecule is electrostatically attracted to the negative end of a neighboring molecule. The strength of this interaction depends strongly on molecular orientation.

c) Characteristics

- Moderate interaction strength (approximately 5–25 $kJ \cdot mol^{-1}$);

- Strong angular dependence;
- Rapid decrease in intensity with increasing distance.

5. Dipole–Induced Dipole Interactions

a) Induction Effect

A polar molecule can distort the electron cloud of a nearby nonpolar molecule, inducing a temporary dipole. The resulting attraction is known as a **dipole–induced dipole interaction**.

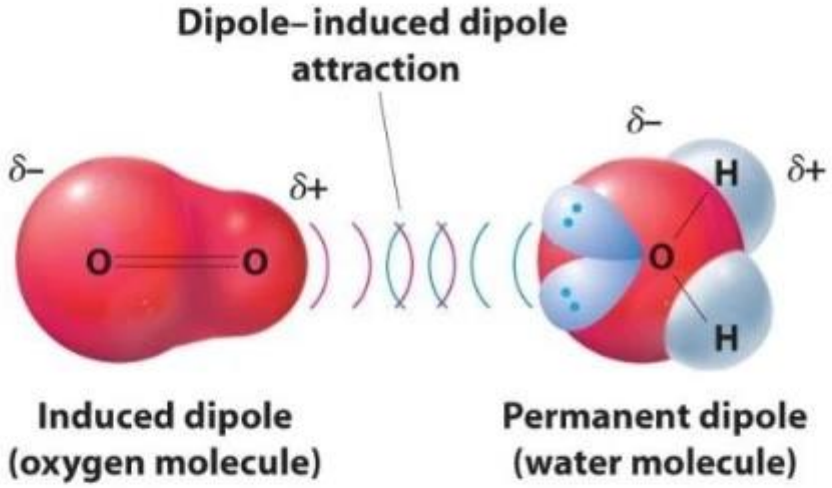


Figure VI.11. Dipole–Induced Dipole Interactions

b) Practical Examples

- Interaction between water molecules and dissolved nonpolar gases;
- Solubility of apolar substances in polar solvents under specific conditions.

c) Relative Strength

These interactions are generally weaker than permanent dipole–dipole interactions but stronger than purely repulsive thermal effects at low temperatures.

6. London Dispersion Forces

a) Quantum Mechanical Origin

Even in nonpolar molecules, electrons are in constant motion. At any instant, fluctuations in electron distribution can generate a transient dipole. This temporary dipole can induce a corresponding dipole in a neighboring molecule, leading to mutual attraction.

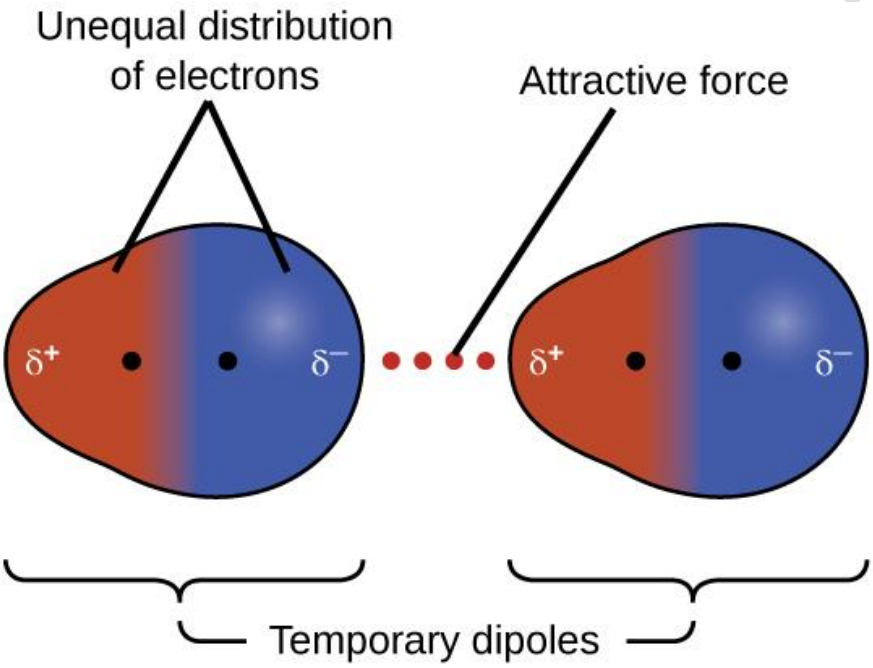


***Figure VI.12. L*ondon Dispersion Forces**

b) Universality of Dispersion Forces

Dispersion forces are present in all atoms and molecules, including noble gases such as helium and argon. They are the only attractive forces operating in purely nonpolar systems.

c) Factors Affecting Strength

The magnitude of dispersion forces increases with:

- molecular size and molar mass;
- electron cloud polarizability;
- molecular surface area and shape.

7. Hydrogen Bonding

a) Definition and Specificity

Hydrogen bonding is a particularly strong and highly directional intermolecular interaction. It occurs when a hydrogen atom covalently bonded to a highly electronegative atom (nitrogen, oxygen, or fluorine) interacts with a lone electron pair on another electronegative atom.

b) Necessary Conditions

For hydrogen bonding to occur, two conditions must be satisfied:

- the presence of a highly polarized X–H bond (X = N, O, or F);
- the availability of a lone electron pair on a neighboring atom.

c) Strength and Directionality

Hydrogen bonds typically exhibit interaction energies ranging from 10 to 40 $kJ \cdot mol^{-1}$, making them stronger than Van der Waals interactions but weaker than covalent bonds.

8. Consequences of Hydrogen Bonding

a) Anomalous Properties of Water

Hydrogen bonding explains many unusual properties of water, including:

- its high boiling point relative to molecular mass;
- its large specific heat capacity;
- the maximum density of liquid water at 4 °C.

b) Biological Significance

Hydrogen bonds play a central role in biological organization:

- stabilization of protein secondary structures (α-helices and β-sheets);
- base pairing in DNA double helices;
- molecular recognition processes in enzymes and receptors.

9. Ion–Dipole Interactions

a) Fundamental Principle

Ion–dipole interactions arise between an ion and a polar molecule. These interactions are particularly important in solutions containing ionic species.

b) Role in Solvation Processes

The dissolution of ionic compounds in polar solvents such as water is driven by strong ion–dipole attractions. For example, sodium and chloride ions are stabilized by surrounding water molecules in aqueous solution.

c) Relative Interaction Strength

Ion–dipole interactions are generally stronger than dipole–dipole interactions but weaker than ionic bonds in crystalline solids.

10. Influence of Intermolecular Forces on Physical Properties

Intermolecular attractive forces govern many observable properties of substances, including:

- melting and boiling temperatures;
- vapor pressure;
- viscosity and fluidity;
- surface tension;
- solubility behavior.

A clear understanding of these forces allows chemists to predict and rationalize material behavior under various conditions.

Intermolecular attractive forces constitute a fundamental aspect of chemical science. While individually weak, their collective influence determines the structure, stability, and properties of matter across all states. Mastery of these interactions is essential for studies in general chemistry, physical chemistry, biochemistry, and materials science.

Key points to remember:

1. **Intermolecular forces (IMFs)** are attractive forces that occur **between molecules**, unlike chemical bonds which occur **within molecules**.
2. Intermolecular forces are generally **much weaker than covalent or ionic bonds**.
3. IMFs strongly influence **physical properties** such as boiling point, melting point, viscosity, and surface tension.
4. Molecules with **stronger intermolecular forces usually have higher boiling and melting points**.
5. Intermolecular forces arise from **electrostatic interactions between partial or temporary charges**.
6. The three main types of intermolecular forces between neutral molecules are:
   - **London dispersion forces**
   - **Dipole–dipole interactions**
   - **Hydrogen bonding**
7. **London dispersion forces (LDF)** occur in **all atoms and molecules**, including nonpolar molecules.
8. Dispersion forces arise from **temporary fluctuations in electron distribution**, creating instantaneous dipoles.
9. The strength of dispersion forces increases with **increasing molecular size and polarizability**.
10. Molecules with **more electrons** generally exhibit **stronger dispersion forces**.
11. **Dipole–dipole interactions** occur between **polar molecules** that possess permanent dipole moments.
12. In dipole–dipole interactions, the **positive end of one molecule attracts the negative end of another molecule**.
13. Dipole–dipole forces are usually **stronger than dispersion forces for molecules of similar size**.
14. **Hydrogen bonding** is a particularly strong type of dipole–dipole interaction.
15. Hydrogen bonding occurs when **hydrogen is covalently bonded to a highly electronegative atom** such as **N, O, or F**.

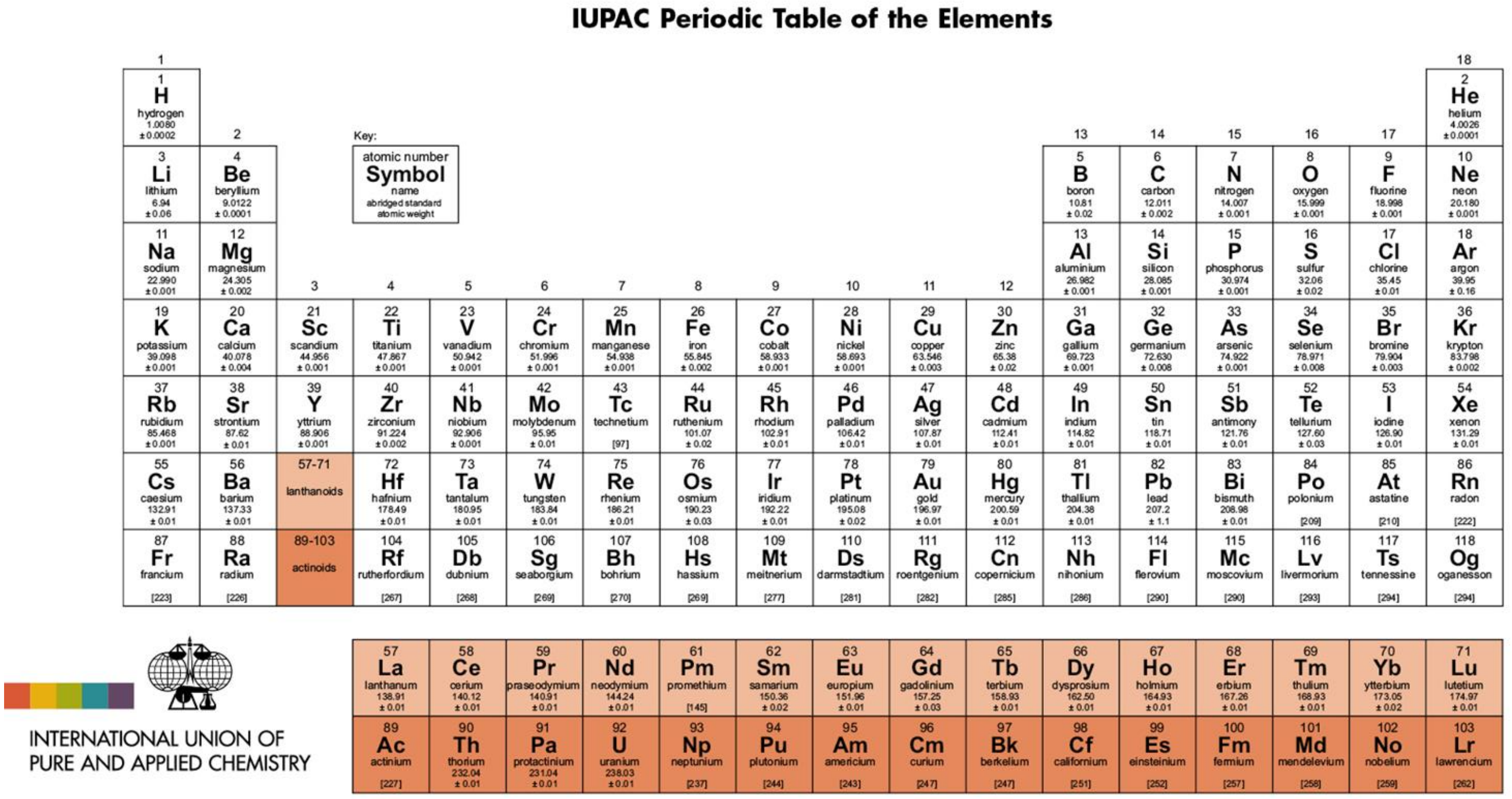
16. The hydrogen atom involved in hydrogen bonding interacts with a **lone pair of electrons on N, O, or F of another molecule**.
17. Hydrogen bonding explains many unusual properties of **water**, including its **high boiling point** and **high surface tension**.
18. Intermolecular forces determine **whether substances exist as gases, liquids, or solids under given conditions**.
19. Strong intermolecular forces generally lead to **higher viscosity and stronger surface tension in liquids**.
20. The relative strength of common intermolecular forces generally follows the order:

Hydrogen bonding > Dipole–dipole interactions > London dispersion forces

(for molecules of comparable size and structure).

Appendices

Appendix A: The periodic table of the elements, basic, in pictures, and color chart for common elements.



IUPAC Periodic Table of the Elements

Key:

atomic number
Symbol
name
abridged standard atomic weight

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** hydrogen 1.0080 ± 0.0002																	2 **He** helium 4.0026 ± 0.0001
3 **Li** lithium 6.94 ± 0.06	4 **Be** beryllium 9.0122 ± 0.0001											5 **B** boron 10.81 ± 0.02	6 **C** carbon 12.011 ± 0.002	7 **N** nitrogen 14.007 ± 0.001	8 **O** oxygen 15.999 ± 0.001	9 **F** fluorine 18.998 ± 0.001	10 **Ne** neon 20.180 ± 0.001
11 **Na** sodium 22.990 ± 0.001	12 **Mg** magnesium 24.305 ± 0.002											13 **Al** aluminium 26.982 ± 0.001	14 **Si** silicon 28.085 ± 0.001	15 **P** phosphorus 30.974 ± 0.001	16 **S** sulfur 32.06 ± 0.02	17 **Cl** chlorine 35.45 ± 0.01	18 **Ar** argon 39.95 ± 0.16
19 **K** potassium 39.098 ± 0.001	20 **Ca** calcium 40.078 ± 0.004	21 **Sc** scandium 44.956 ± 0.001	22 **Ti** titanium 47.867 ± 0.001	23 **V** vanadium 50.942 ± 0.001	24 **Cr** chromium 51.996 ± 0.001	25 **Mn** manganese 54.938 ± 0.001	26 **Fe** iron 55.845 ± 0.002	27 **Co** cobalt 58.933 ± 0.001	28 **Ni** nickel 58.693 ± 0.001	29 **Cu** copper 63.546 ± 0.003	30 **Zn** zinc 65.38 ± 0.02	31 **Ga** gallium 69.723 ± 0.001	32 **Ge** germanium 72.630 ± 0.008	33 **As** arsenic 74.922 ± 0.001	34 **Se** selenium 78.971 ± 0.008	35 **Br** bromine 79.904 ± 0.003	36 **Kr** krypton 83.798 ± 0.002
37 **Rb** rubidium 85.468 ± 0.001	38 **Sr** strontium 87.62 ± 0.01	39 **Y** yttrium 88.906 ± 0.001	40 **Zr** zirconium 91.224 ± 0.002	41 **Nb** niobium 92.906 ± 0.001	42 **Mo** molybdenum 95.95 ± 0.01	43 **Tc** technetium [97]	44 **Ru** ruthenium 101.07 ± 0.02	45 **Rh** rhodium 102.91 ± 0.01	46 **Pd** palladium 106.42 ± 0.01	47 **Ag** silver 107.87 ± 0.01	48 **Cd** cadmium 112.41 ± 0.01	49 **In** indium 114.82 ± 0.01	50 **Sn** tin 118.71 ± 0.01	51 **Sb** antimony 121.76 ± 0.01	52 **Te** tellurium 127.60 ± 0.03	53 **I** iodine 126.90 ± 0.01	54 **Xe** xenon 131.29 ± 0.01
55 **Cs** caesium 132.91 ± 0.01	56 **Ba** barium 137.33 ± 0.01	57-71 lanthanoids	72 **Hf** hafnium 178.49 ± 0.01	73 **Ta** tantalum 180.95 ± 0.01	74 **W** tungsten 183.84 ± 0.01	75 **Re** rhenium 186.21 ± 0.01	76 **Os** osmium 190.23 ± 0.03	77 **Ir** iridium 192.22 ± 0.01	78 **Pt** platinum 195.08 ± 0.02	79 **Au** gold 196.97 ± 0.01	80 **Hg** mercury 200.59 ± 0.01	81 **Tl** thallium 204.38 ± 0.01	82 **Pb** lead 207.2 ± 1.1	83 **Bi** bismuth 208.98 ± 0.01	84 **Po** polonium [209]	85 **At** astatine [210]	86 **Rn** radon [222]
87 **Fr** francium [223]	88 **Ra** radium [226]	89-103 actinoids	104 **Rf** rutherfordium [267]	105 **Db** dubnium [268]	106 **Sg** seaborgium [269]	107 **Bh** bohrium [270]	108 **Hs** hassium [269]	109 **Mt** meitnerium [277]	110 **Ds** darmstadtium [281]	111 **Rg** roentgenium [282]	112 **Cn** copernicium [285]	113 **Nh** nihonium [286]	114 **Fl** flerovium [290]	115 **Mc** moscovium [290]	116 **Lv** livermorium [293]	117 **Ts** tennessine [294]	118 **Og** oganesson [294]

57 **La** lanthanum 138.91 ± 0.01	58 **Ce** cerium 140.12 ± 0.01	59 **Pr** praseodymium 140.91 ± 0.01	60 **Nd** neodymium 144.24 ± 0.01	61 **Pm** promethium [145]	62 **Sm** samarium 150.36 ± 0.02	63 **Eu** europium 151.96 ± 0.01	64 **Gd** gadolinium 157.25 ± 0.03	65 **Tb** terbium 158.93 ± 0.01	66 **Dy** dysprosium 162.50 ± 0.01	67 **Ho** holmium 164.93 ± 0.01	68 **Er** erbium 167.26 ± 0.01	69 **Tm** thulium 168.93 ± 0.01	70 **Yb** ytterbium 173.05 ± 0.02	71 **Lu** lutetium 174.97 ± 0.01
89 **Ac** actinium [227]	90 **Th** thorium 232.04 ± 0.01	91 **Pa** protactinium 231.04 ± 0.01	92 **U** uranium 238.03 ± 0.01	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 [262]

INTERNATIONAL UNION OF
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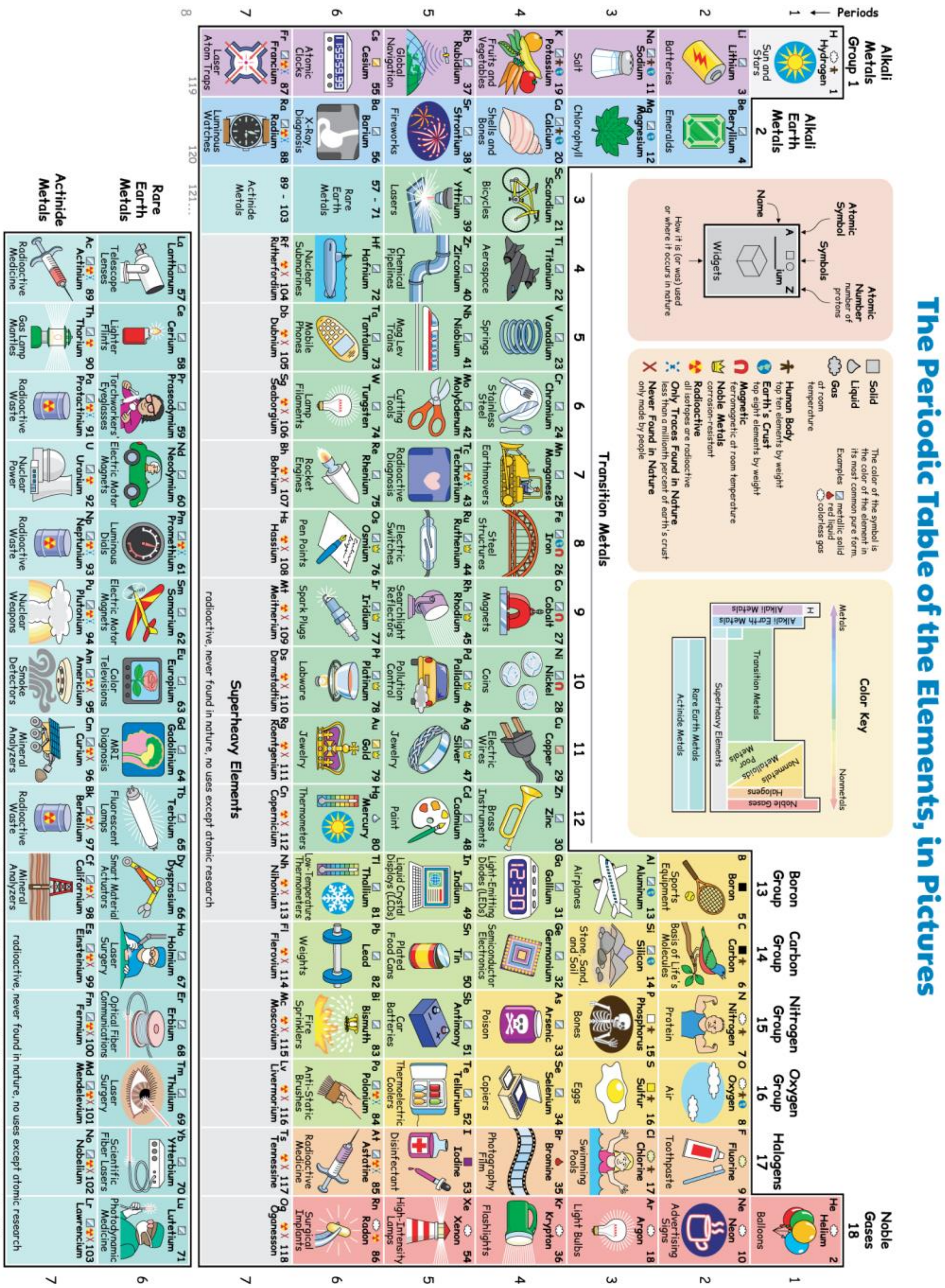
For notes and updates to this table, see www.iupac.org. This version is dated 4 May 2022.



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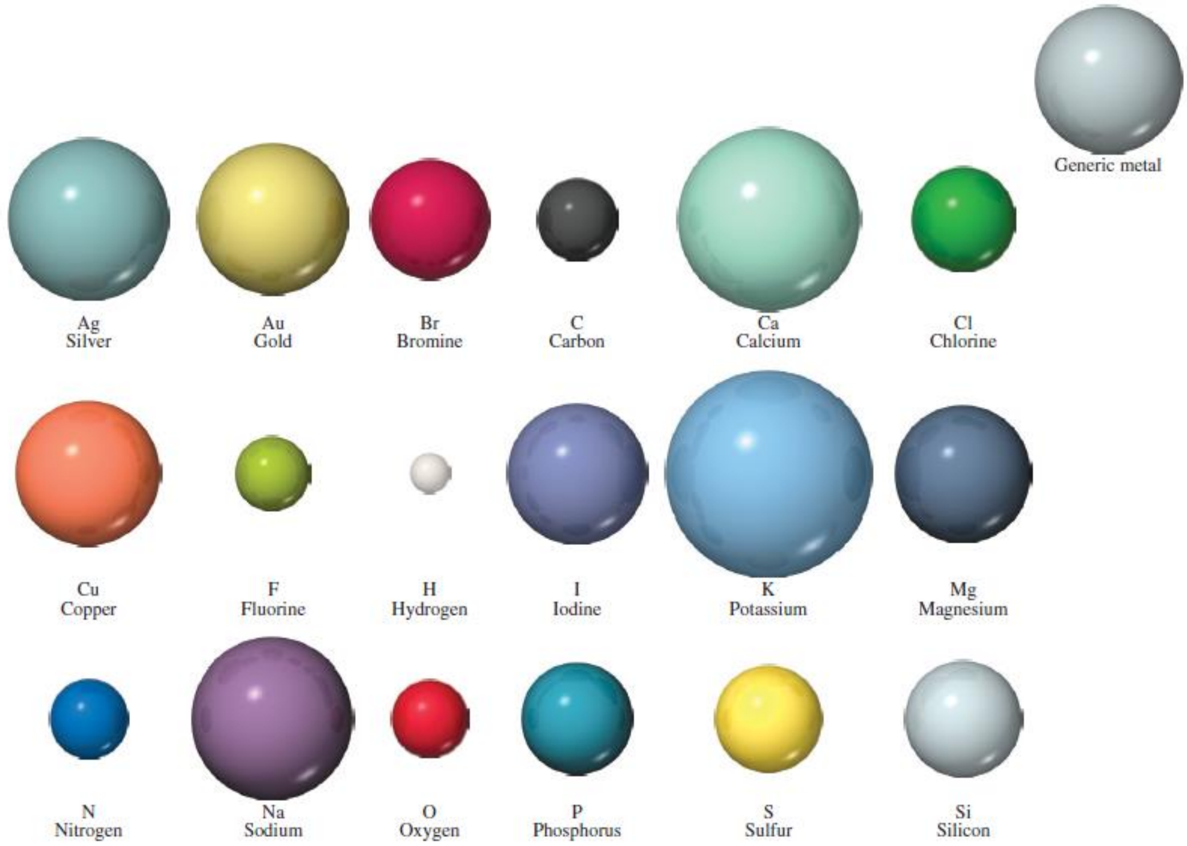
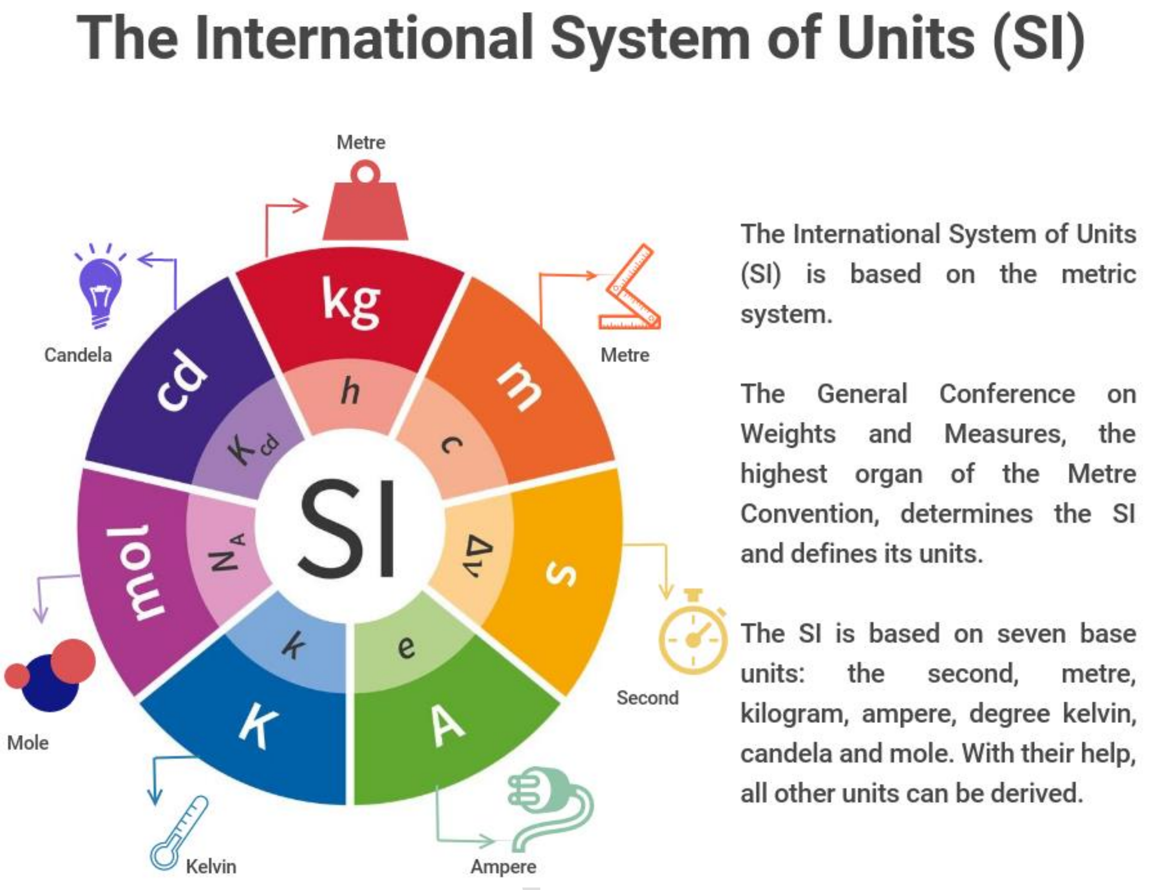
Color Chart for Common Elements



Appendix B: The seven SI base units, SI Prefix

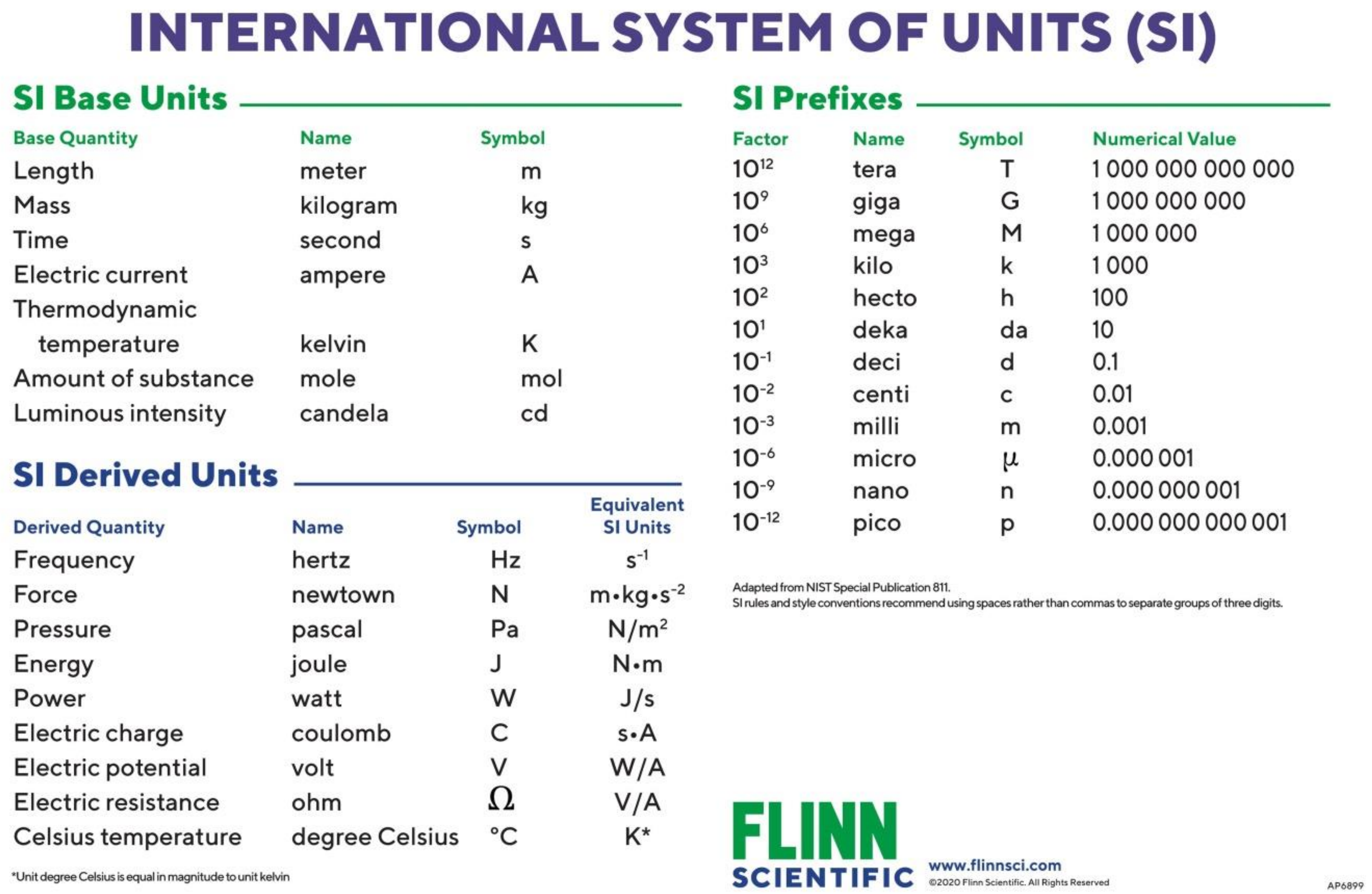

The International System of Units (SI)

The International System of Units (SI) is based on the metric system.

The General Conference on Weights and Measures, the highest organ of the Metre Convention, determines the SI and defines its units.

The SI is based on seven base units: the second, metre, kilogram, ampere, degree kelvin, candela and mole. With their help, all other units can be derived.

Symbol	Name	Base quantity
s	second	time
m	meter	length
kg	kilogram	mass
A	ampere	electric current
K	kelvin	thermodynamic temperature
mol	mole	amount of substance
cd	candela	luminous intensity



INTERNATIONAL SYSTEM OF UNITS (SI)

SI Base Units

Base Quantity	Name	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electric current	ampere	A
Thermodynamic temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

SI Derived Units

Derived Quantity	Name	Symbol	Equivalent SI Units
Frequency	hertz	Hz	s^{-1}
Force	newtown	N	$m \cdot kg \cdot s^{-2}$
Pressure	pascal	Pa	N/m^2
Energy	joule	J	$N \cdot m$
Power	watt	W	J/s
Electric charge	coulomb	C	$s \cdot A$
Electric potential	volt	V	W/A
Electric resistance	ohm	Ω	V/A
Celsius temperature	degree Celsius	°C	K*

*Unit degree Celsius is equal in magnitude to unit kelvin

SI Prefixes

Factor	Name	Symbol	Numerical Value
10^{12}	tera	T	1 000 000 000 000
10^{9}	giga	G	1 000 000 000
10^{6}	mega	M	1 000 000
10^{3}	kilo	k	1 000
10^{2}	hecto	h	100
10^{1}	deka	da	10
10^{-1}	deci	d	0.1
10^{-2}	centi	c	0.01
10^{-3}	milli	m	0.001
10^{-6}	micro	μ	0.000 001
10^{-9}	nano	n	0.000 000 001
10^{-12}	pico	p	0.000 000 000 001

Adapted from NIST Special Publication 811.
SI rules and style conventions recommend using spaces rather than commas to separate groups of three digits.

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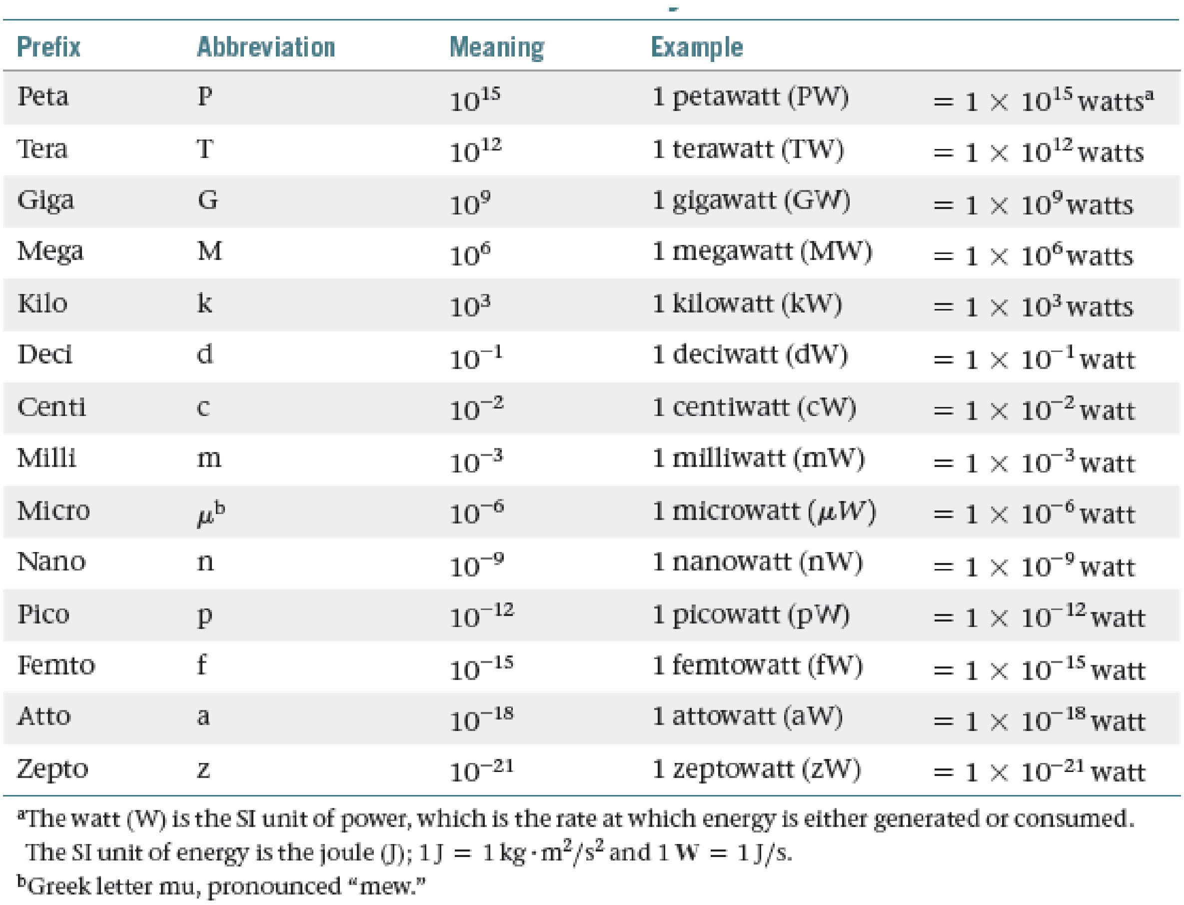
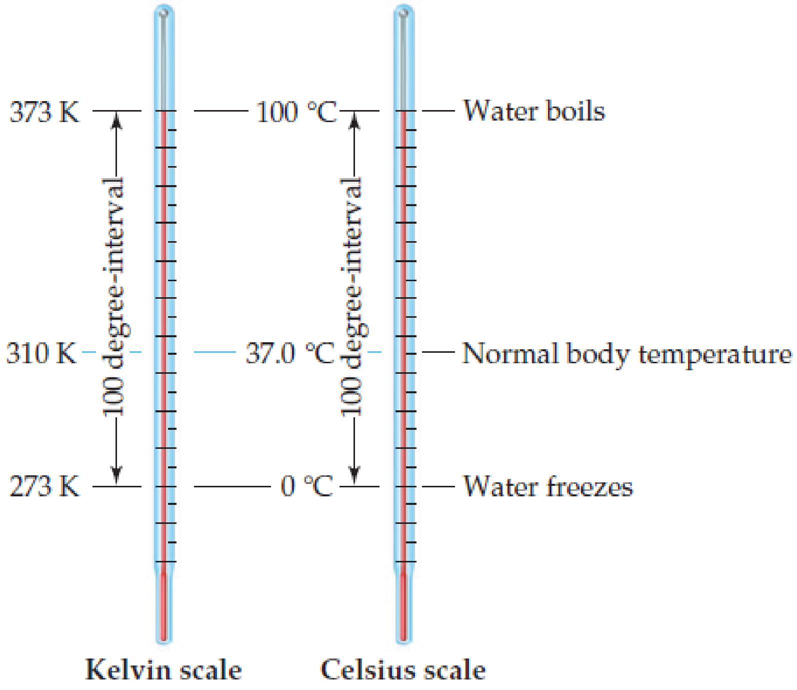


Prefix	Abbreviation	Meaning	Example	
Peta	P	10^{15}	1 petawatt (PW)	$= 1 \times 10^{15}$ watts[a]
Tera	T	10^{12}	1 terawatt (TW)	$= 1 \times 10^{12}$ watts
Giga	G	10^{9}	1 gigawatt (GW)	$= 1 \times 10^{9}$ watts
Mega	M	10^{6}	1 megawatt (MW)	$= 1 \times 10^{6}$ watts
Kilo	k	10^{3}	1 kilowatt (kW)	$= 1 \times 10^{3}$ watts
Deci	d	10^{-1}	1 deciwatt (dW)	$= 1 \times 10^{-1}$ watt
Centi	c	10^{-2}	1 centiwatt (cW)	$= 1 \times 10^{-2}$ watt
Milli	m	10^{-3}	1 milliwatt (mW)	$= 1 \times 10^{-3}$ watt
Micro	μ[b]	10^{-6}	1 microwatt (μW)	$= 1 \times 10^{-6}$ watt
Nano	n	10^{-9}	1 nanowatt (nW)	$= 1 \times 10^{-9}$ watt
Pico	p	10^{-12}	1 picowatt (pW)	$= 1 \times 10^{-12}$ watt
Femto	f	10^{-15}	1 femtowatt (fW)	$= 1 \times 10^{-15}$ watt
Atto	a	10^{-18}	1 attowatt (aW)	$= 1 \times 10^{-18}$ watt
Zepto	z	10^{-21}	1 zeptowatt (zW)	$= 1 \times 10^{-21}$ watt

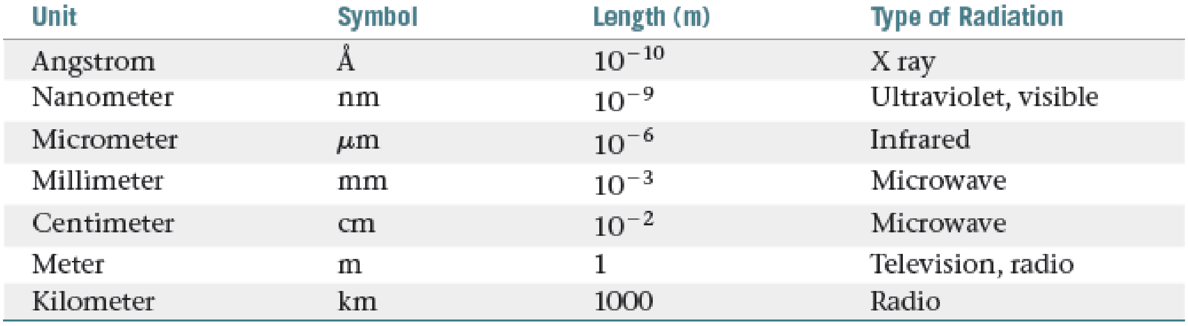
[a] The watt (W) is the SI unit of power, which is the rate at which energy is either generated or consumed. The SI unit of energy is the joule (J); $1\ \text{J} = 1\ \text{kg} \cdot \text{m}^2/\text{s}^2$ and $1\ \text{W} = 1\ \text{J/s}$.

[b] Greek letter mu, pronounced "mew."



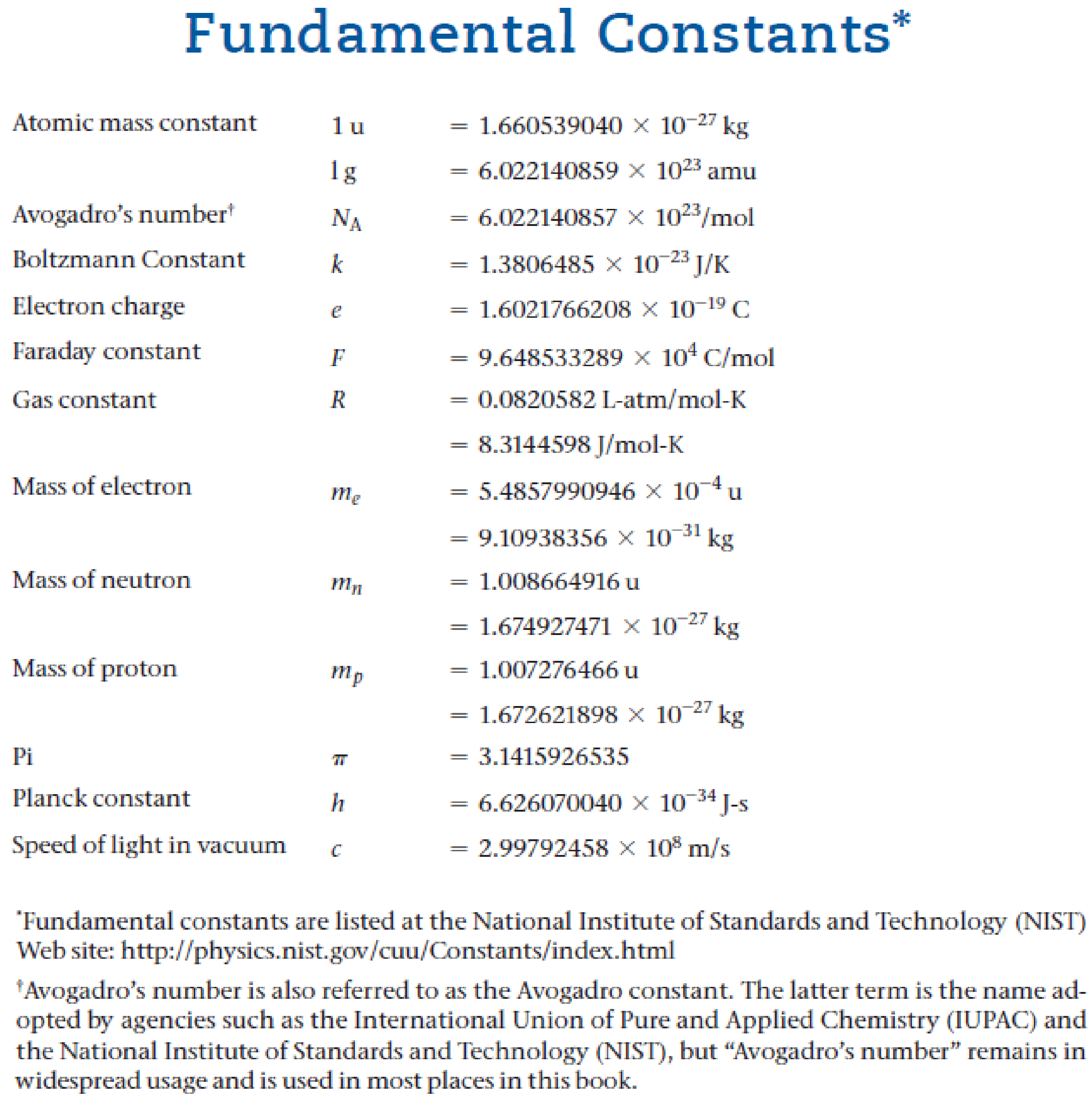
Comparison of the Kelvin and the Celsius temperature scales (K = °C + 273,15).

Appendix C: Common wave length units for electromagnetic radiation



Unit	Symbol	Length (m)	Type of Radiation
Angstrom	Å	10^{-10}	X ray
Nanometer	nm	10^{-9}	Ultraviolet, visible
Micrometer	μm	10^{-6}	Infrared
Millimeter	mm	10^{-3}	Microwave
Centimeter	cm	10^{-2}	Microwave
Meter	m	1	Television, radio
Kilometer	km	1000	Radio

Appendix D: Fundamental Constants, Useful conversion factors and relationships

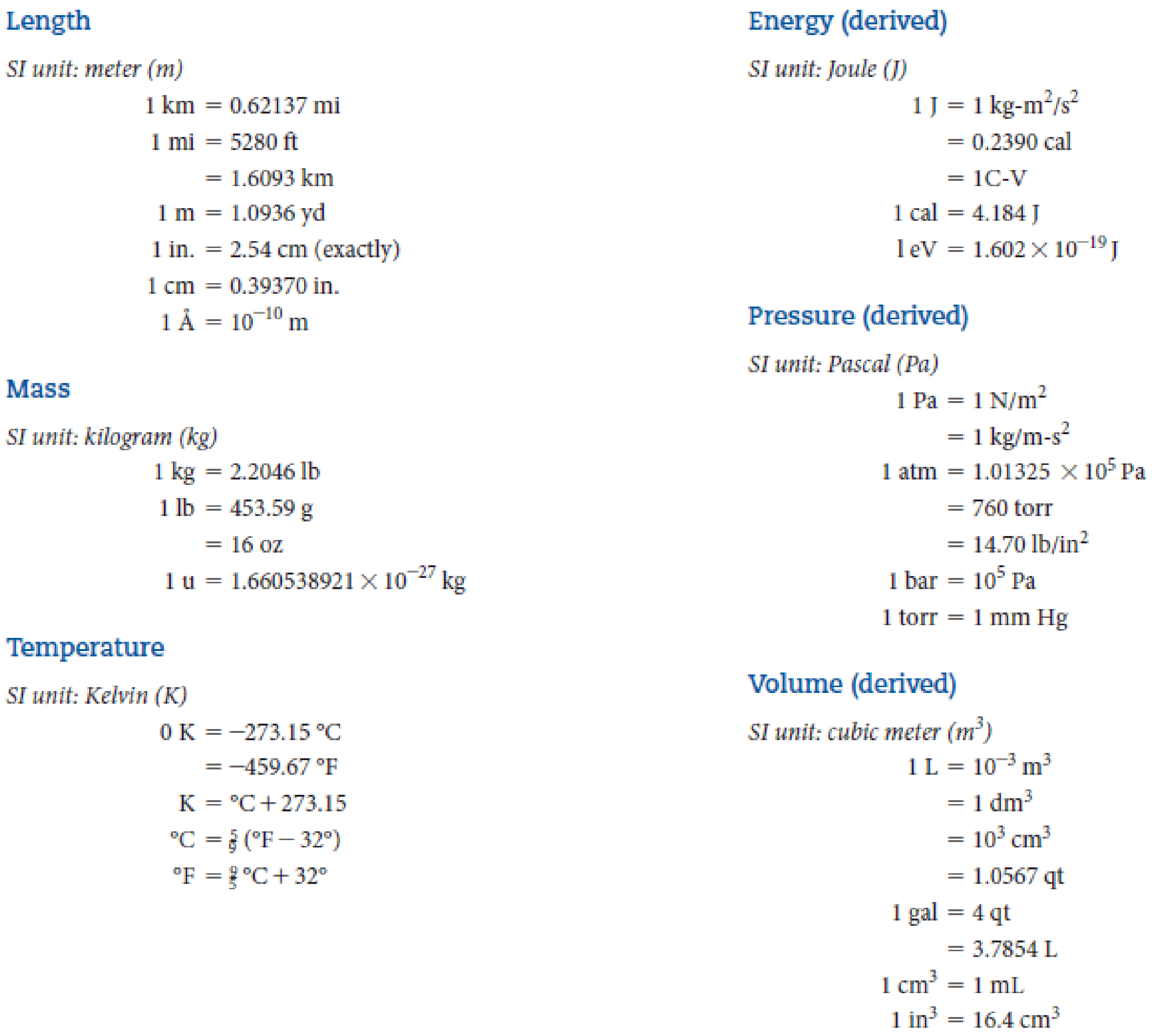


Fundamental Constants[*]

Atomic mass constant	1 u	$= 1.660539040 \times 10^{-27}$ kg
	1 g	$= 6.022140859 \times 10^{23}$ amu
Avogadro's number[†]	N_A	$= 6.022140857 \times 10^{23}$/mol
Boltzmann Constant	k	$= 1.3806485 \times 10^{-23}$ J/K
Electron charge	e	$= 1.6021766208 \times 10^{-19}$ C
Faraday constant	F	$= 9.648533289 \times 10^{4}$ C/mol
Gas constant	R	$= 0.0820582$ L-atm/mol-K
		$= 8.3144598$ J/mol-K
Mass of electron	m_e	$= 5.4857990946 \times 10^{-4}$ u
		$= 9.10938356 \times 10^{-31}$ kg
Mass of neutron	m_n	$= 1.008664916$ u
		$= 1.674927471 \times 10^{-27}$ kg
Mass of proton	m_p	$= 1.007276466$ u
		$= 1.672621898 \times 10^{-27}$ kg
Pi	π	$= 3.1415926535$
Planck constant	h	$= 6.626070040 \times 10^{-34}$ J-s
Speed of light in vacuum	c	$= 2.99792458 \times 10^{8}$ m/s

[*] Fundamental constants are listed at the National Institute of Standards and Technology (NIST) Web site: http://physics.nist.gov/cuu/Constants/index.html

[†] Avogadro's number is also referred to as the Avogadro constant. The latter term is the name adopted by agencies such as the International Union of Pure and Applied Chemistry (IUPAC) and the National Institute of Standards and Technology (NIST), but "Avogadro's number" remains in widespread usage and is used in most places in this book.



Length

SI unit: meter (m)

$1\text{ km} = 0.62137\text{ mi}$

$1\text{ mi} = 5280\text{ ft}$

$= 1.6093\text{ km}$

$1\text{ m} = 1.0936\text{ yd}$

$1\text{ in.} = 2.54\text{ cm (exactly)}$

$1\text{ cm} = 0.39370\text{ in.}$

$1\text{ Å} = 10^{-10}\text{ m}$

Mass

SI unit: kilogram (kg)

$1\text{ kg} = 2.2046\text{ lb}$

$1\text{ lb} = 453.59\text{ g}$

$= 16\text{ oz}$

$1\text{ u} = 1.660538921 \times 10^{-27}\text{ kg}$

Temperature

SI unit: Kelvin (K)

$0\text{ K} = -273.15\ °\text{C}$

$= -459.67\ °\text{F}$

$\text{K} = °\text{C} + 273.15$

$°\text{C} = \frac{5}{9}(°\text{F} - 32°)$

$°\text{F} = \frac{9}{5}°\text{C} + 32°$

Energy (derived)

SI unit: Joule (J)

$1\text{ J} = 1\text{ kg-m}^2/\text{s}^2$

$= 0.2390\text{ cal}$

$= 1\text{C-V}$

$1\text{ cal} = 4.184\text{ J}$

$1\text{ eV} = 1.602 \times 10^{-19}\text{ J}$

Pressure (derived)

SI unit: Pascal (Pa)

$1\text{ Pa} = 1\text{ N/m}^2$

$= 1\text{ kg/m-s}^2$

$1\text{ atm} = 1.01325 \times 10^5\text{ Pa}$

$= 760\text{ torr}$

$= 14.70\text{ lb/in}^2$

$1\text{ bar} = 10^5\text{ Pa}$

$1\text{ torr} = 1\text{ mm Hg}$

Volume (derived)

SI unit: cubic meter (m^3)

$1\text{ L} = 10^{-3}\text{ m}^3$

$= 1\text{ dm}^3$

$= 10^3\text{ cm}^3$

$= 1.0567\text{ qt}$

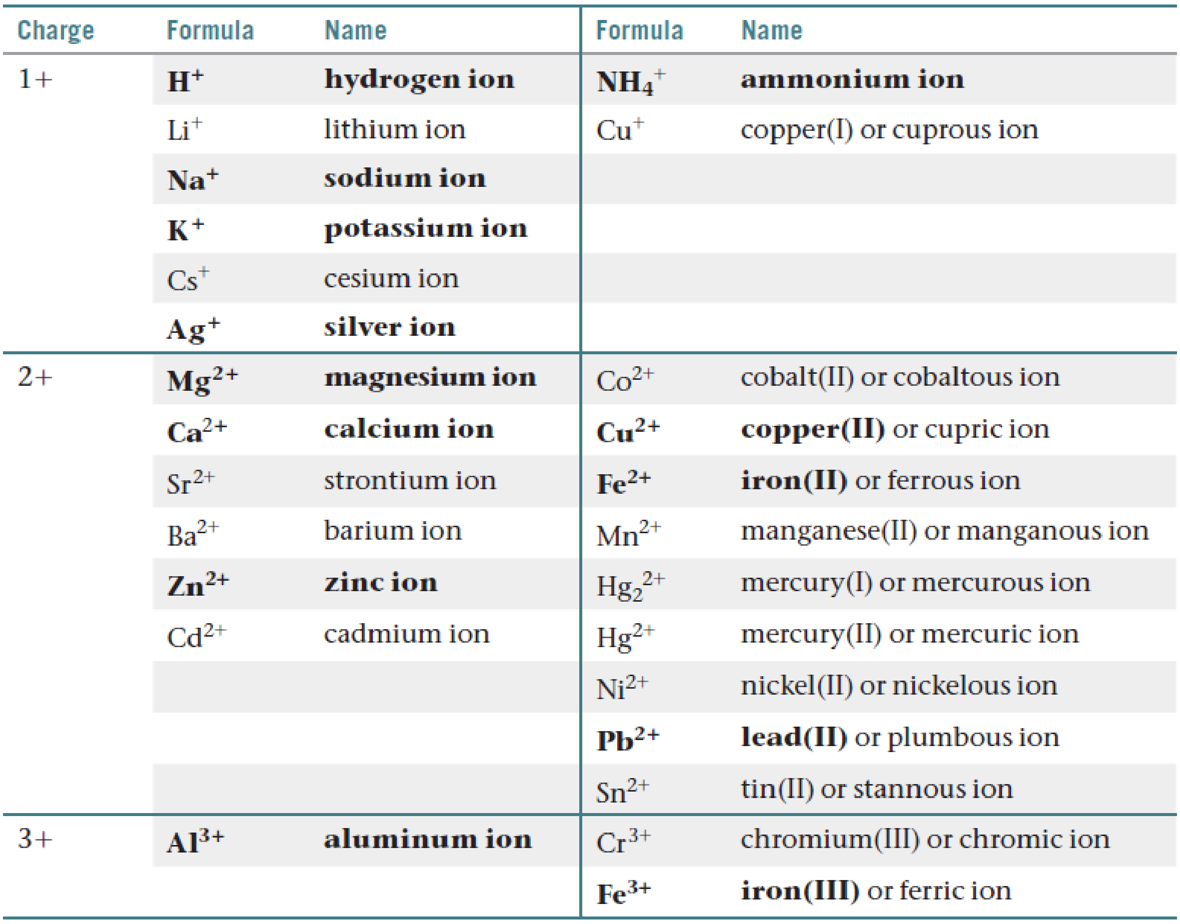
$1\text{ gal} = 4\text{ qt}$

$= 3.7854\text{ L}$

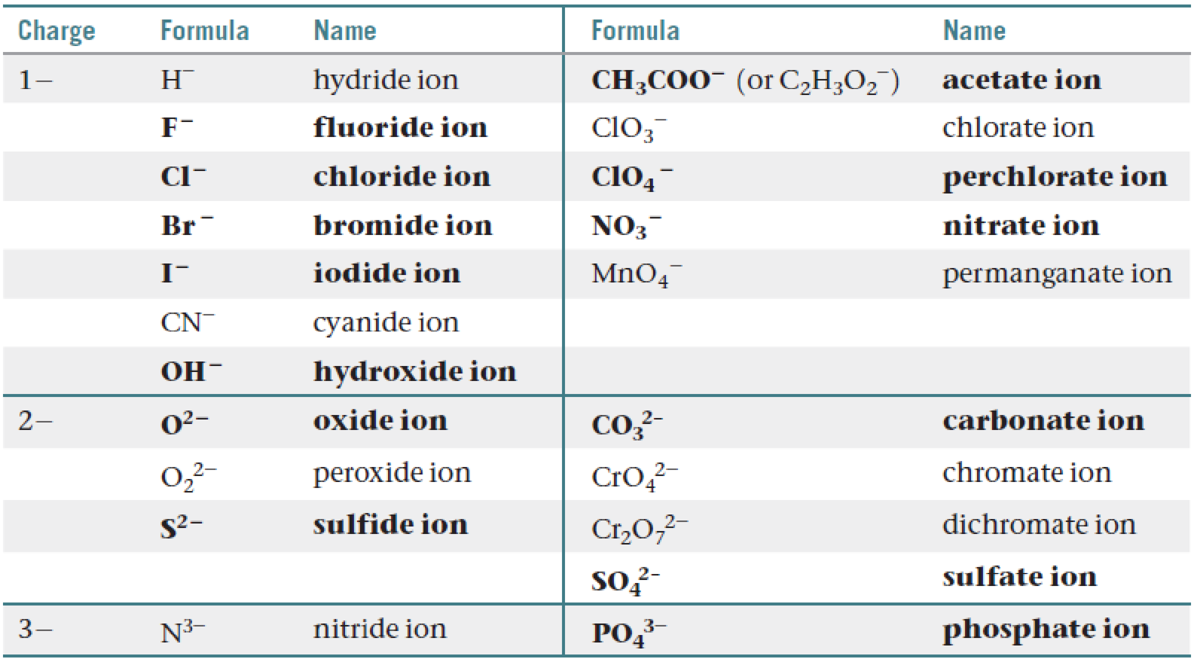
$1\text{ cm}^3 = 1\text{ mL}$

$1\text{ in}^3 = 16.4\text{ cm}^3$

Appendix E: Common cations and common anions

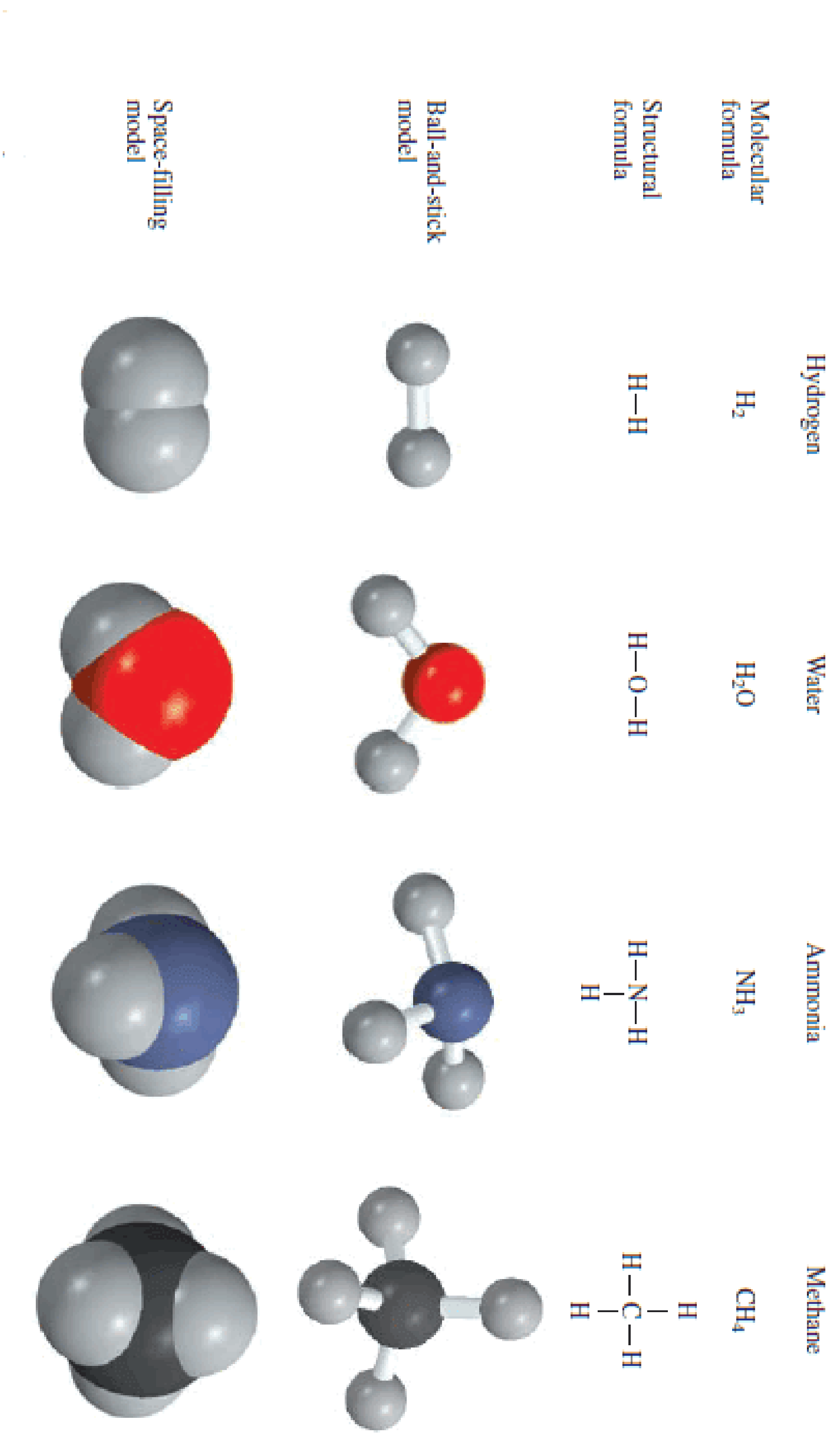


Charge	Formula	Name	Formula	Name
1+	**H^+**	**hydrogen ion**	**NH_4^+**	**ammonium ion**
	Li^+	lithium ion	Cu^+	copper(I) or cuprous ion
	Na^+	**sodium ion**		
	K^+	**potassium ion**		
	Cs^+	cesium ion		
	Ag^+	**silver ion**		
2+	**Mg^{2+}**	**magnesium ion**	Co^{2+}	cobalt(II) or cobaltous ion
	Ca^{2+}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



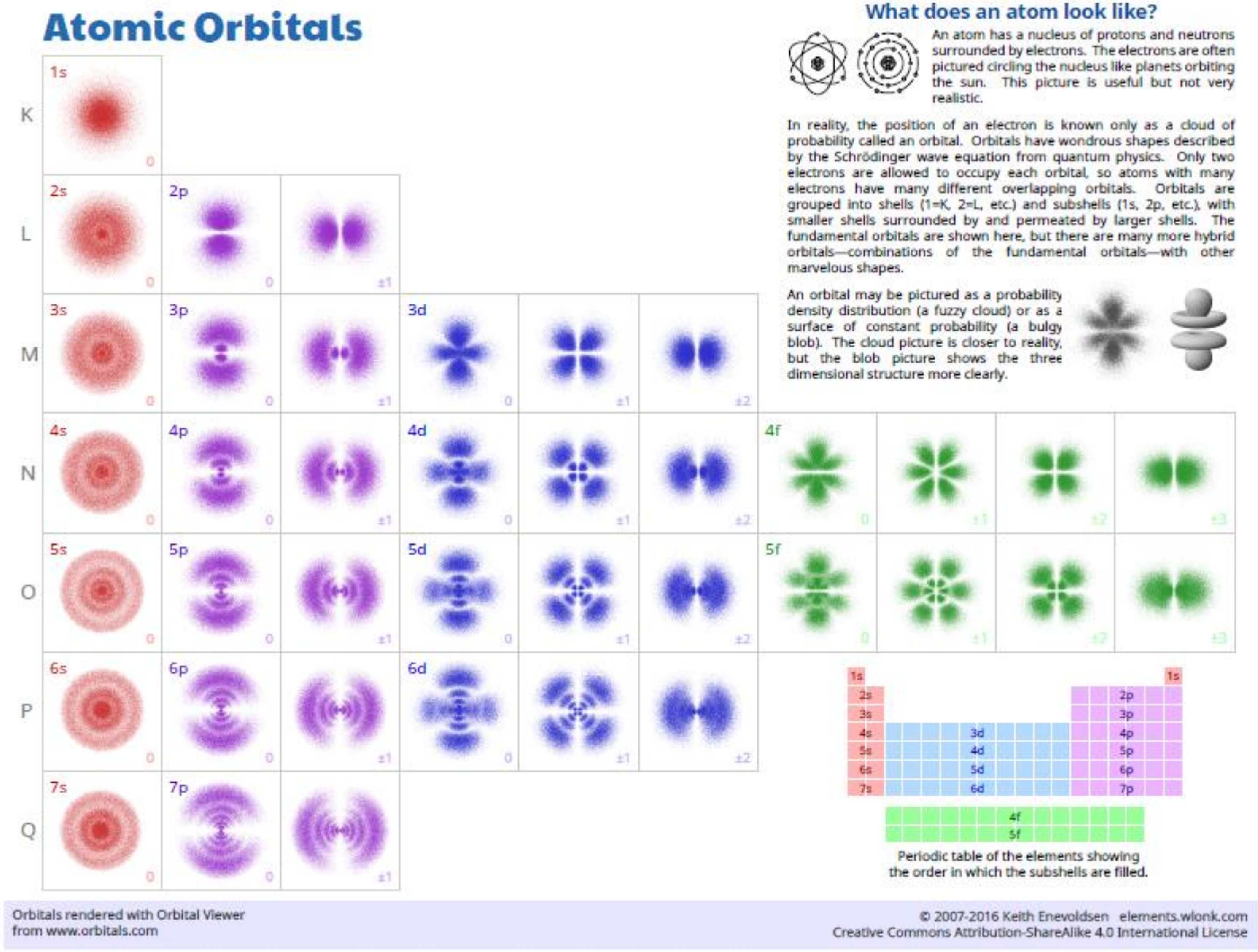
Charge	Formula	Name	Formula	Name
1−	H^-	hydride ion	**CH_3COO^-** (or $C_2H_3O_2^-$)	**acetate ion**
	F^-	**fluoride ion**	ClO_3^-	chlorate ion
	Cl^-	**chloride ion**	**ClO_4^-**	**perchlorate ion**
	Br^-	**bromide ion**	**NO_3^-**	**nitrate ion**
	I^-	**iodide ion**	MnO_4^-	permanganate ion
	CN^-	cyanide ion		
	OH^-	**hydroxide ion**		
2−	**O^{2-}**	**oxide ion**	**CO_3^{2-}**	**carbonate ion**
	O_2^{2-}	peroxide ion	CrO_4^{2-}	chromate ion
	S^{2-}	**sulfide ion**	$Cr_2O_7^{2-}$	dichromate ion
			SO_4^{2-}	**sulfate ion**
3−	N^{3-}	nitride ion	**PO_4^{3-}**	**phosphate ion**

Appendix F: Molecular models



Molecular and structural formulas and molecular models of four common molecules.

Appendix G: Atomic orbitals



Atomic Orbitals

What does an atom look like?

An atom has a nucleus of protons and neutrons surrounded by electrons. The electrons are often pictured circling the nucleus like planets orbiting the sun. This picture is useful but not very realistic.

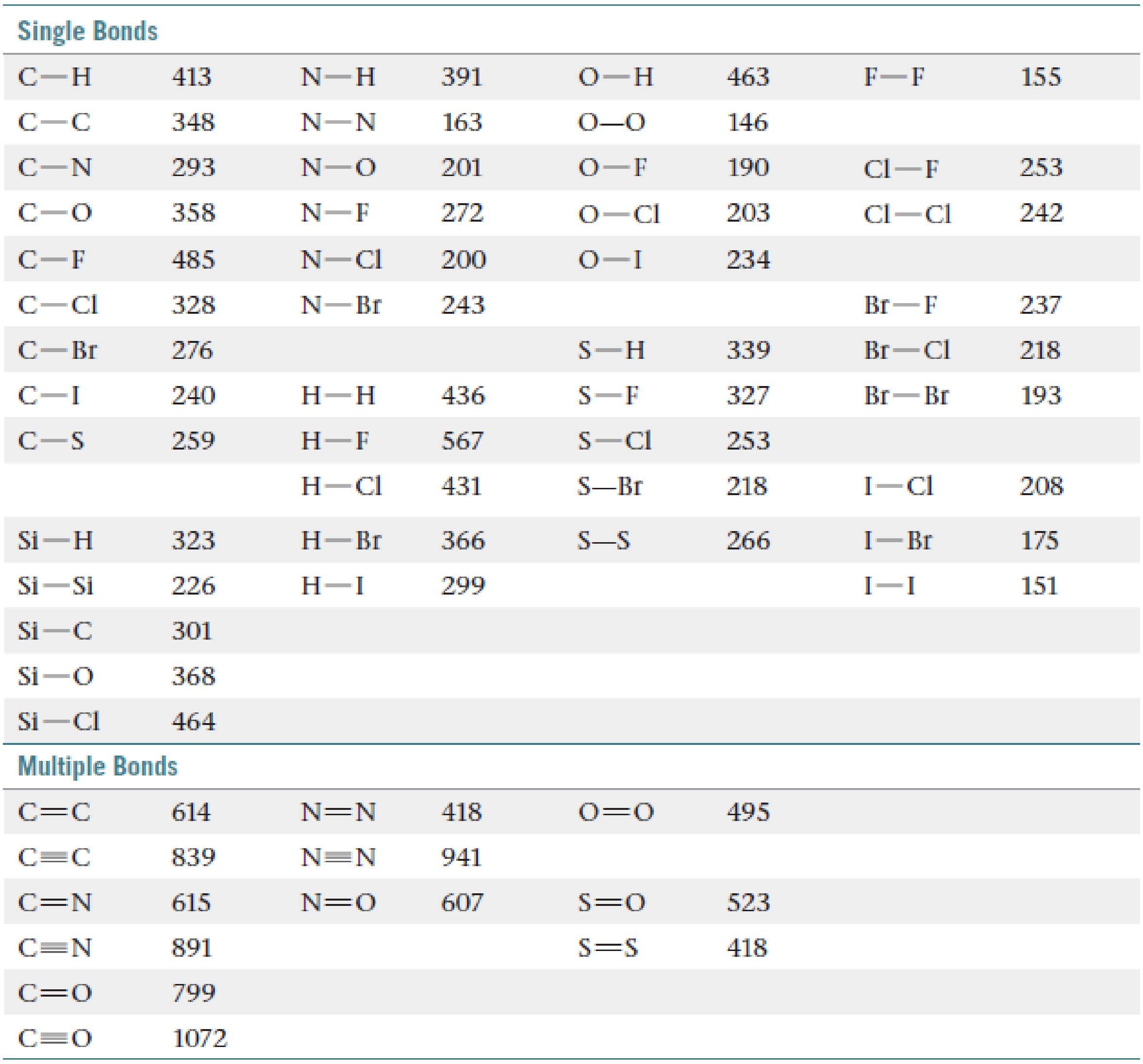
In reality, the position of an electron is known only as a cloud of probability called an orbital. Orbitals have wondrous shapes described by the Schrödinger wave equation from quantum physics. Only two electrons are allowed to occupy each orbital, so atoms with many electrons have many different overlapping orbitals. Orbitals are grouped into shells (1=K, 2=L, etc.) and subshells (1s, 2p, etc.), with smaller shells surrounded by and permeated by larger shells. The fundamental orbitals are shown here, but there are many more hybrid orbitals—combinations of the fundamental orbitals—with other marvelous shapes.

An orbital may be pictured as a probability density distribution (a fuzzy cloud) or as a surface of constant probability (a bulgy blob). The cloud picture is closer to reality, but the blob picture shows the three dimensional structure more clearly.

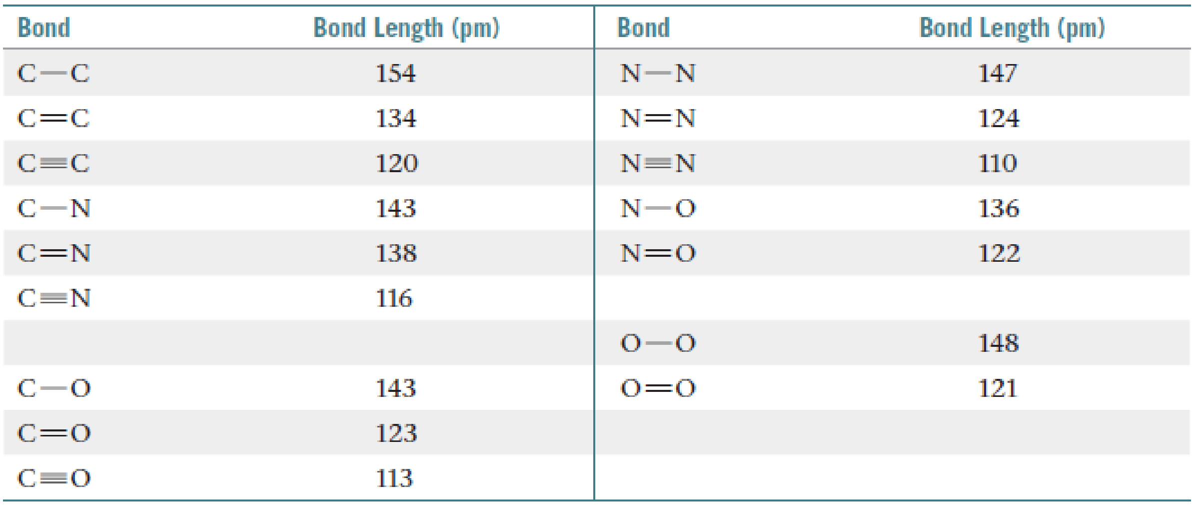
Periodic table of the elements showing the order in which the subshells are filled.

Orbitals rendered with Orbital Viewer
from www.orbitals.com

Appendix H: Average bond enthalpies (kJ/mol) and average bond lengths for some single, double and triple bonds.

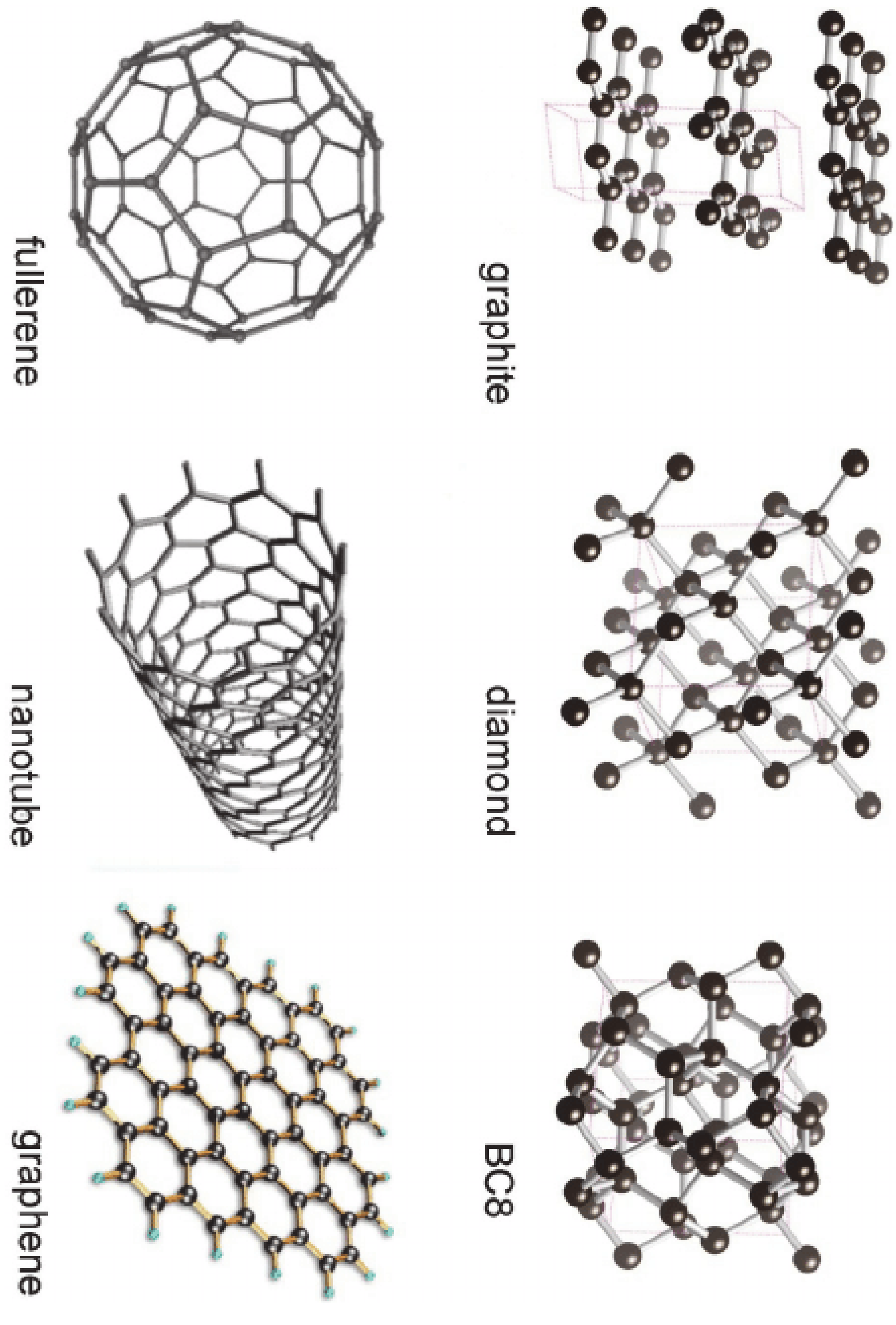


Single Bonds							
C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
Si—H	323	H—Br	366	S—S	266	I—Br	175
Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						
Si—Cl	464						
Multiple Bonds							
C=C	614	N=N	418	O=O	495		
C≡C	839	N≡N	941				
C=N	615	N=O	607	S=O	523		
C≡N	891			S=S	418		
C=O	799						
C≡O	1072						

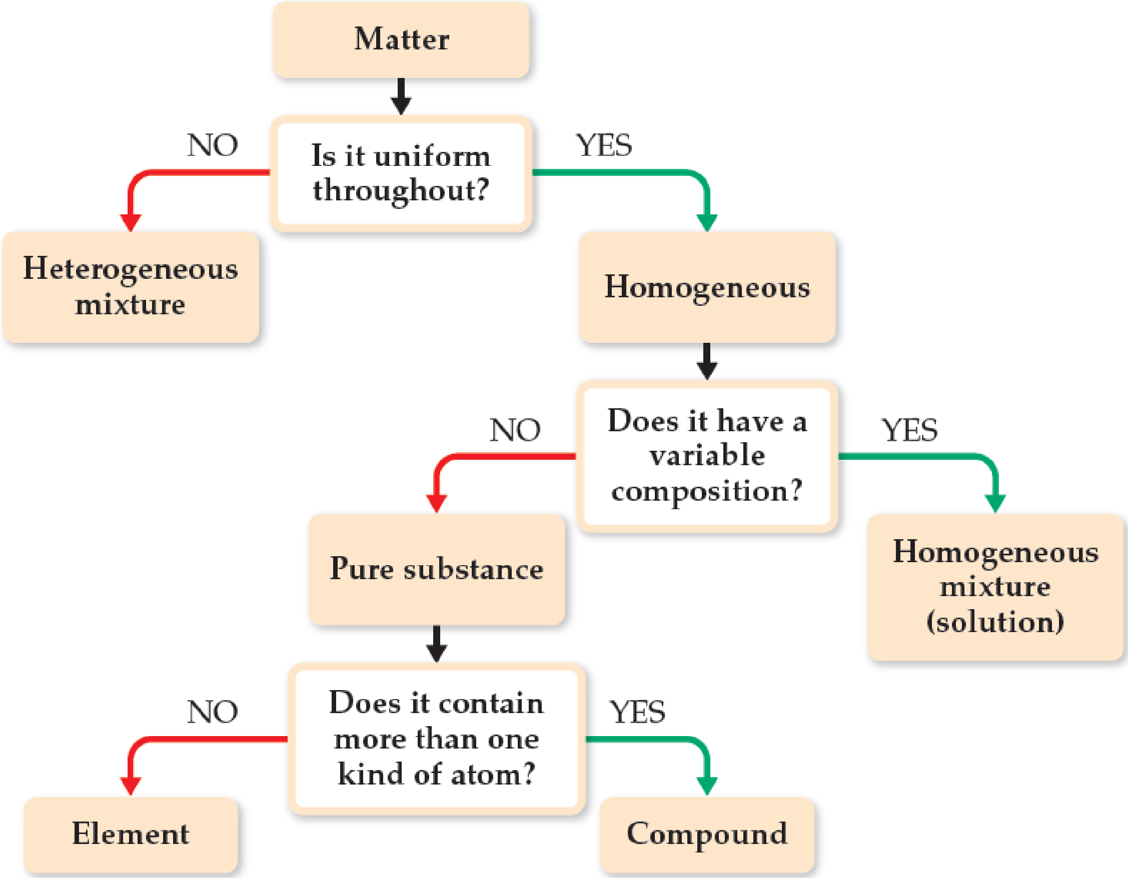


Bond	Bond Length (pm)	Bond	Bond Length (pm)
C—C	154	N—N	147
C=C	134	N=N	124
C≡C	120	N≡N	110
C—N	143	N—O	136
C=N	138	N=O	122
C≡N	116		
		O—O	148
C—O	143	O=O	121
C=O	123		
C≡O	113		

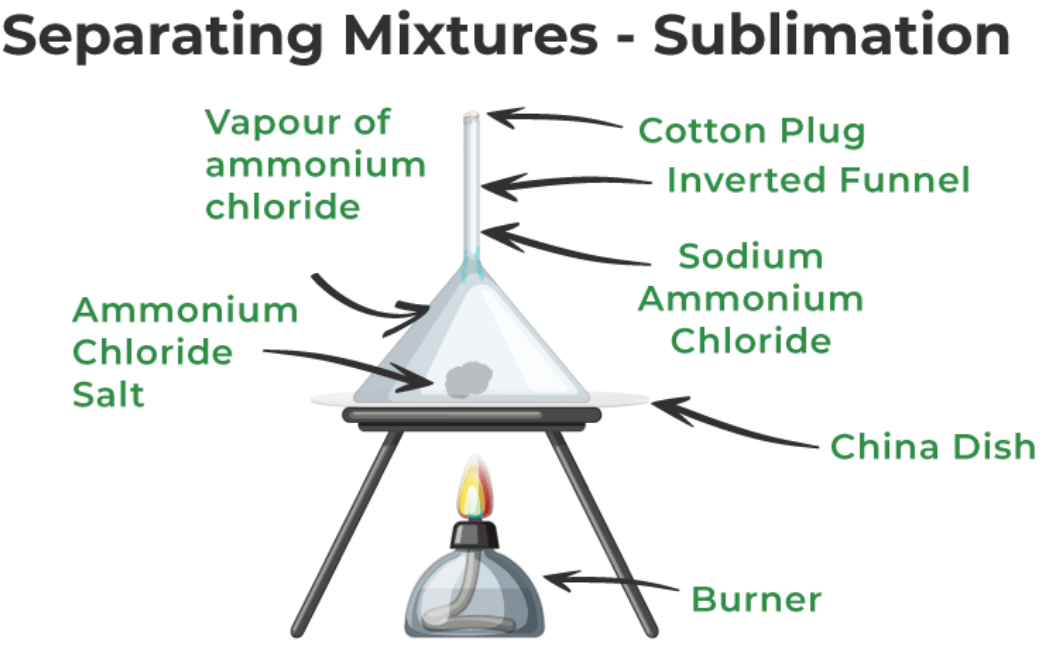
Appendix I: Allotropes of Carbon

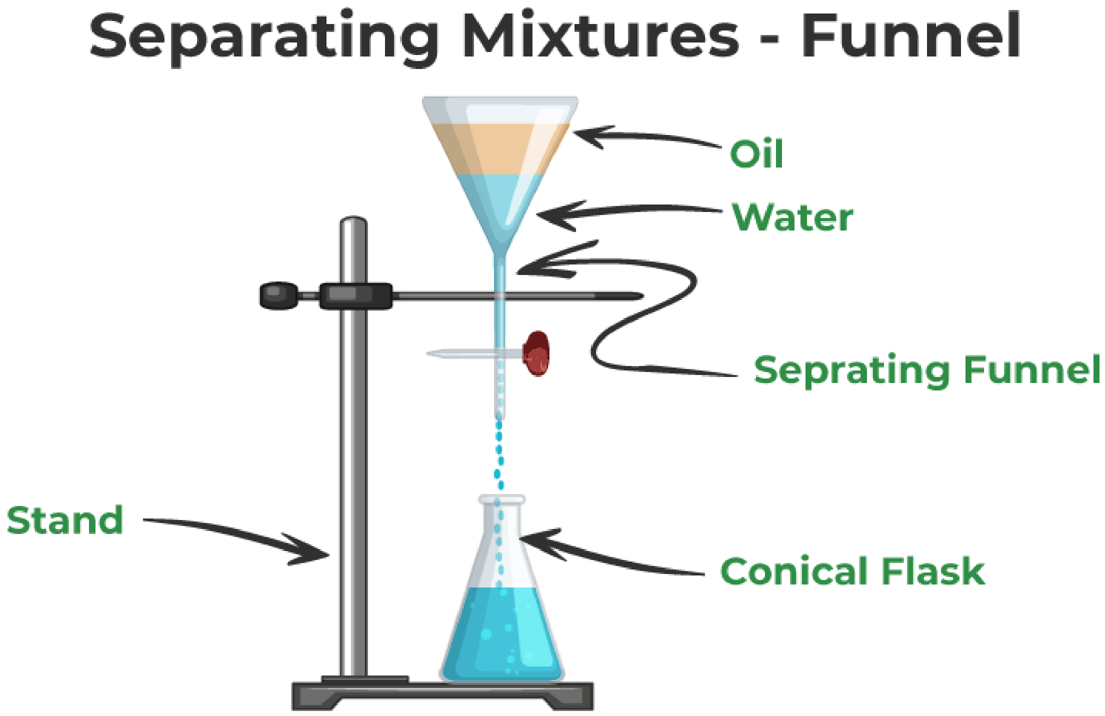


Appendix J: Classification of matter. All pure matter is classified ultimately as either an element or a compound.

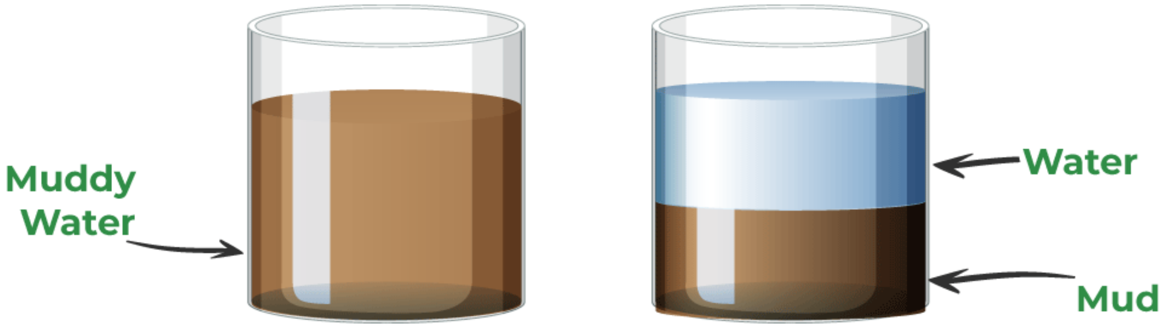


Appendix K: Separation of mixtures

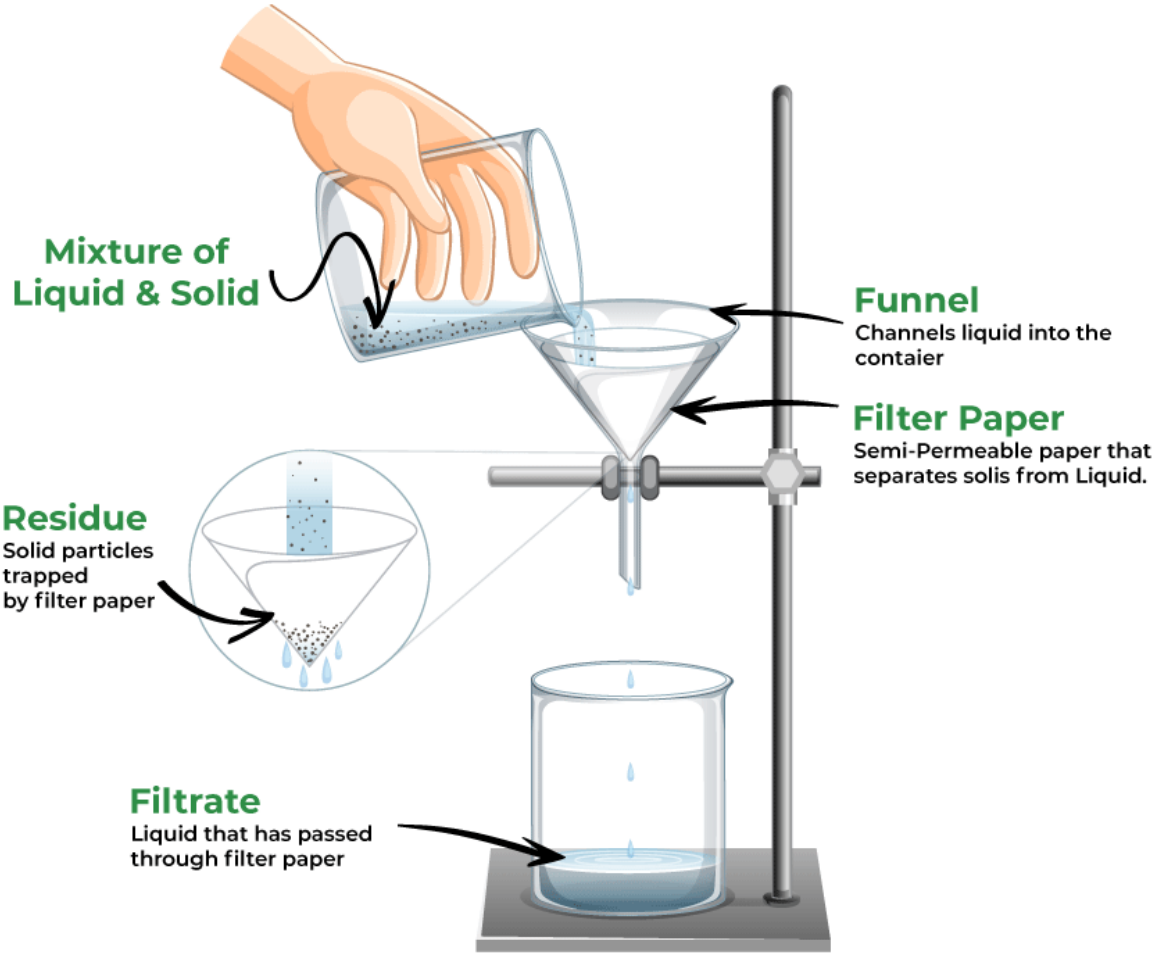




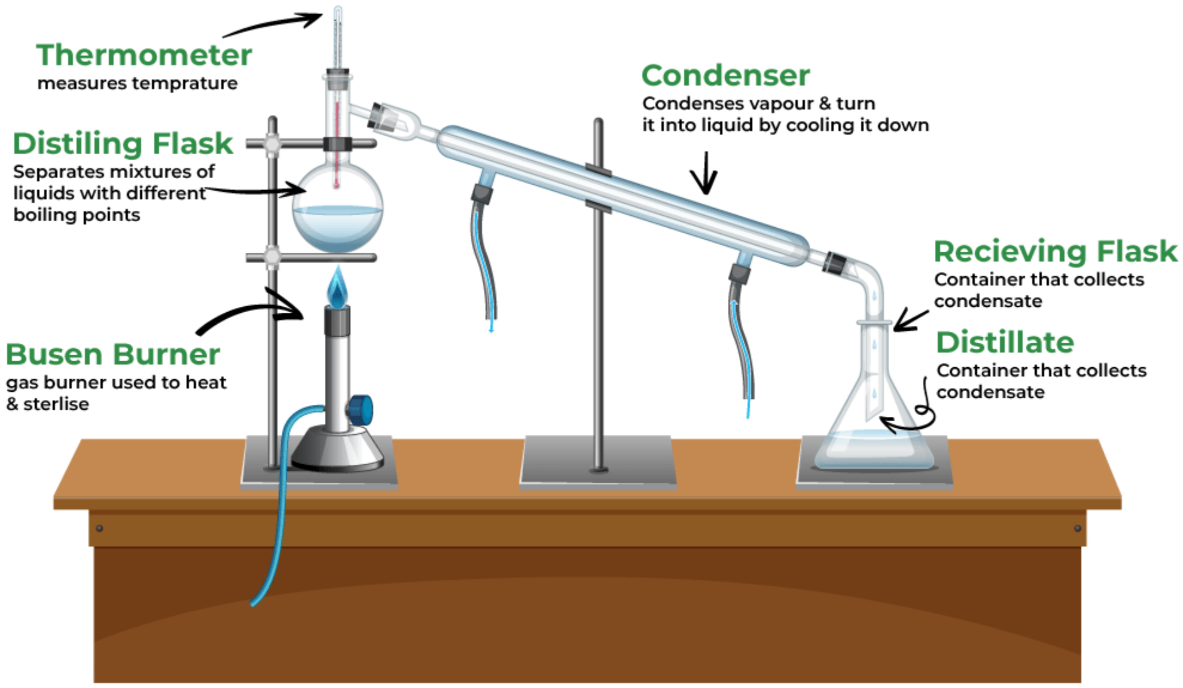
Separating Mixtures - Sedimentation



Separating Mixture - Filtration

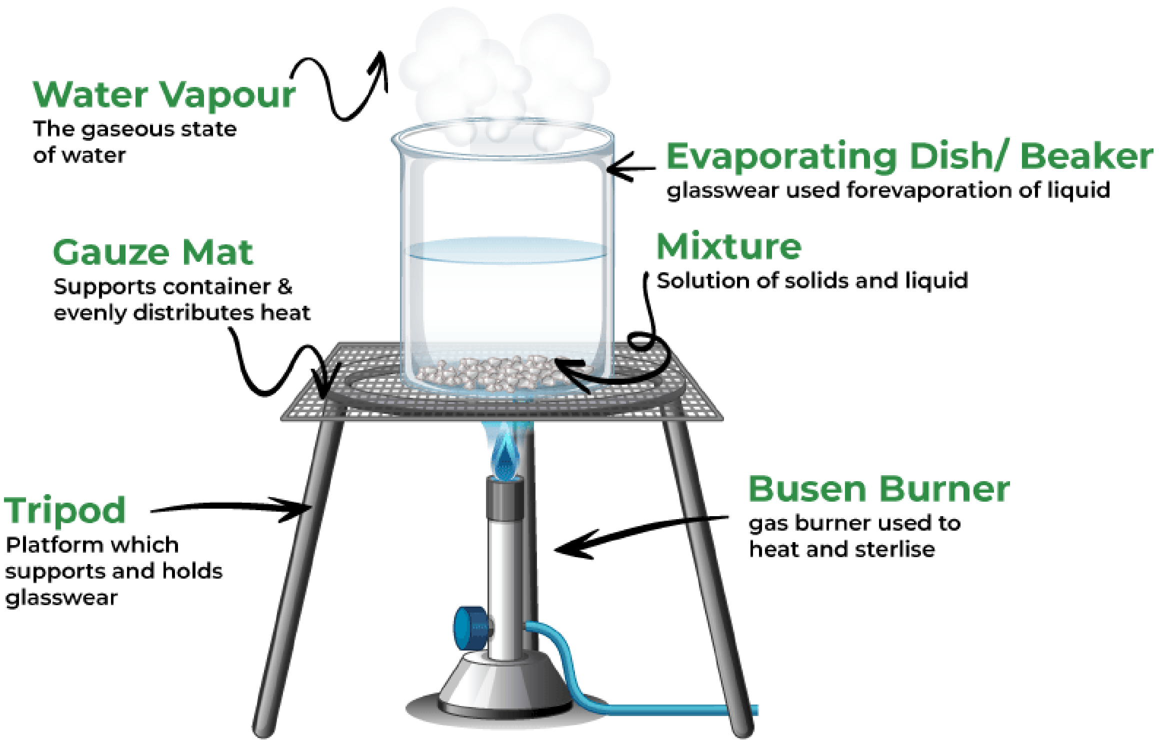


Separating Mixture - Distillation



Seprating Mixtures - Sieving

Separating Mixtures - Evaporation



Separating Mixture - Magnet

Appendix L: Hazard pictograms

GHS Pictogram	Pictogram Name	Hazards Represented
	Exclamation Point	‣ Irritant to eyes, skin and respiratory tract ‣ Skin sensitizer ‣ Acute toxicity (harmful)
	Health Hazard	‣ Carcinogen ‣ Reproductive Toxicity ‣ Target organ toxicity ‣ Mutagenicity ‣ Aspiration toxicity ‣ (Effects usually occur over time)
	Skull and Crossbones	‣ Acute toxicity (severe or fatal) ‣ (Effects can occur shortly after exposure) ‣ (Poisons)
	Corrosive	‣ Corrosive and damaging to skin and eyes ‣ Chemical burns ‣ Corrosive and damaging to metals
	Flame	‣ Flammables ‣ Self-reactives ‣ Pyrophorics ‣ Self-heating ‣ Emits Flammable Gas ‣ Organic Peroxides
	Gas Cylinder	‣ Gases under pressure
	Flame Over Circle	‣ Oxidizing gases, liquids, or solids ‣ (perchlorates) ‣ (Oxidizers are chemicals that facilitate burning or make fires burn hotter and longer.)
	Exploding Bomb	‣ Explosives ‣ Self-reactives ‣ Organic peroxides ‣ (especially reactive or shock sensitive)
	Environment	‣ Aquatic toxicity ‣ (not human health)

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CHEMICAL HAZARD SYMBOLS

Chemical hazard symbols are found on some home products, as well as bottles of chemical reagents in the lab. Here, we take a look at European hazard symbols and the various dangers that they warn of.

ENVIRONMENTAL HAZARD

Indicates substances that are toxic to aquatic organisms, or may cause long lasting environmental effects. They should be disposed of responsibly.

ACUTELY TOXIC

Indicates life-threatening effects, in some cases even after limited exposure. Any form of ingestion and skin contact should be avoided.

GAS UNDER PRESSURE

Container contains pressurised gas. This may be cold when released, and explosive when heated. Containers should not be heated.

CORROSIVE

May cause burns to skin and damage to eyes. May also corrode metals. Avoid skin & eye contact, and do not breathe vapours.

EXPLOSIVE

May explode as a consequence of fire, heat, shock or friction. Chemicals with this label should be kept away from potential ignition sources.

FLAMMABLE

Flammable when exposed to heat, fire or sparks, or give off flammable gases when reacting with water. Ignition sources should be avoided.

MODERATE HAZARD

May irritate the skin, or exhibit minor toxicity. The chemical should be kept away from the skin and the eyes as a precaution.

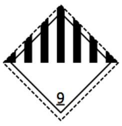
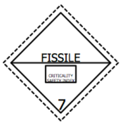
OXIDISING

Burns even in the absence of air, and can intensify fires in combustible materials. Should be kept away from ignition sources.

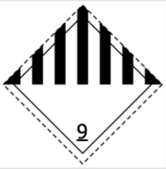
HEALTH HAZARD

Short or long term exposure could cause serious long term health effects. Skin contact and ingestion of this chemical should be avoided.

TRANSPORT PICTOGRAMS

Class 1 Explosives	Class 2 Gases	Classes 3 and 4: Flammable liquids and solids	Other GHS transport classes	Non-GHS transport pictograms	
 Divisions 1.1-1.3	 Flammable gases	 Flammable liquids	 Oxidizing substances	 Infectious substances	 Radioactive material
 Divisions 1.1-1.3	 Non-flammable non-toxic gases	 Flammable solids	 Organic peroxides	 Miscellaneous dangerous substances and articles	 Radioactive material
 Divisions 1.1-1.3	 Toxic gases	 Substances liable to spontaneous combustion	 Toxic substances		 Radioactive material
 Divisions 1.1-1.3		 Substances which in contact with water emit flammable gases	 Corrosive substances		 Radioactive material

ADR pictograms

 1-Explosives	 2.1-Flammable gases	 2.2-Non-toxic and non-flammable gases	 2.3 Poison gases	 3-Flammable liquids
 4.1-Flammable solids	 4.2-Spontaneously combustibles	 4.3-Dangerous when wet	 5.1-Oxidizers	 5.2-Organic peroxides
 6.1-Poison	 6.2-Infectious substances	 7-Radioactive	 8-Corrosive	 9-Miscellaneous dangerous substances

References

1) Brown, T. E., LeMay, H. E., Bursten, B. E., Murphy, C. J., Woodward, P. M., & Stoltzfus, M. E. (2023). *Chemistry: The central science* (15th ed.). Pearson.

2) McQuarrie, D. A., & Simon, J. D. (2023). *Physical chemistry: A molecular approach*. University Science Books.

3) Zumdahl, S. S., Zumdahl, S. A., & DeCoste, D. J. (2023). *Chemistry* (11th ed.). Cengage Learning.

4) Oxtoby, D. W., Gillis, H. P., & Butler, L. J. (2021). *Principles of modern chemistry* (8th ed.). Cengage Learning.

5) Atkins, P. W., de Paula, J., & Keeler, J. (2023). *Atkins' physical chemistry* (12th ed.). Oxford University Press.

6) Atkins, P., Jones, L., Laverman, L., Young, K., & Patterson, J. (2023). *Chemical principles: The quest for insight* (8th ed.). W. H. Freeman.