



Phase 1 Package

Scientific Track – NSMO

Chemistry

Chemistry Scientific Team

Table of Contents

No.	Subject	Contents	Pages
1	Atomic structure	- Composition of the atom - The Law of Conservation of Mass - The Law of Definite Proportions - The Law of Multiple Proportions - Historical Atomic Models - Atomic Structure - Electron Configuration - Quantum Numbers	3 - 27
2	The Periodic Table and Classification of Elements	- Classification of Elements - Modern construction of the periodic table - Division of the elements of the periodic table - Periodic table and electron configuration - Properties of the periodic table - Atomic size - Electronegativity - Ionization energy - Electron affinity	28 - 48
3	Ionic Bonds and Compounds	- Ions - Ionic Bonds - Properties of Ionic Compounds - Steps for Writing the Chemical Formula of Ionic Compounds - Metallic Bond	49 - 61
4	Covalent Bonds and Compounds	- The Covalent Bonding Model - Formation of the Covalent Bond - Properties of the Covalent Bond - Naming Molecules ΔEN (The Electronegativity Difference (ΔEN) and Bond Polarity - The Shapes of Molecules - VSEPR (Valence-Shell Electron-Pair Repulsion (VSEPR) Theory and Molecular Shape	62 - 81

01

Atomic Structure

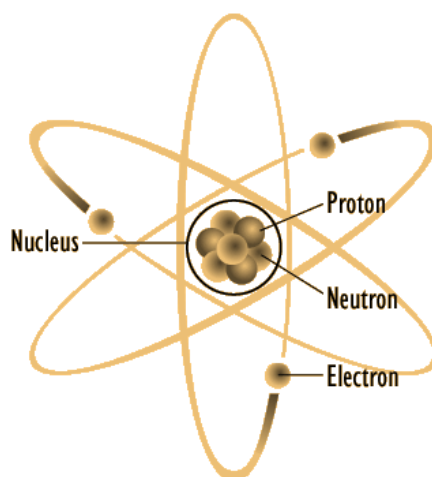
Composition of the atom

If you cut a piece of graphite from the tip of a pencil into smaller and smaller pieces, how far could you go? Could you divide it forever? Would you eventually run into some basic particles that were no longer divisible, not because of their sheer smallness, but because of the nature of matter? This fundamental question about the nature of matter has been asked by thinkers for over two millennia. Their answers have varied over time. On the scale of everyday objects, matter appears continuous, or infinitely divisible. And until about 200 years ago, many scientists thought that matter was indeed continuous—but they were proven wrong. If you were to divide the graphite from your pencil tip into smaller and smaller pieces (far smaller than the eye could see), you would eventually end up with individual carbon atoms. The word atom comes from the Greek atomos, meaning “indivisible”.

You cannot divide a carbon atom into smaller pieces and still have carbon. Atoms compose all ordinary matter. If you want to understand matter, you must begin by understanding atoms.

The Modern Atomic Theory and the Laws it is Based On

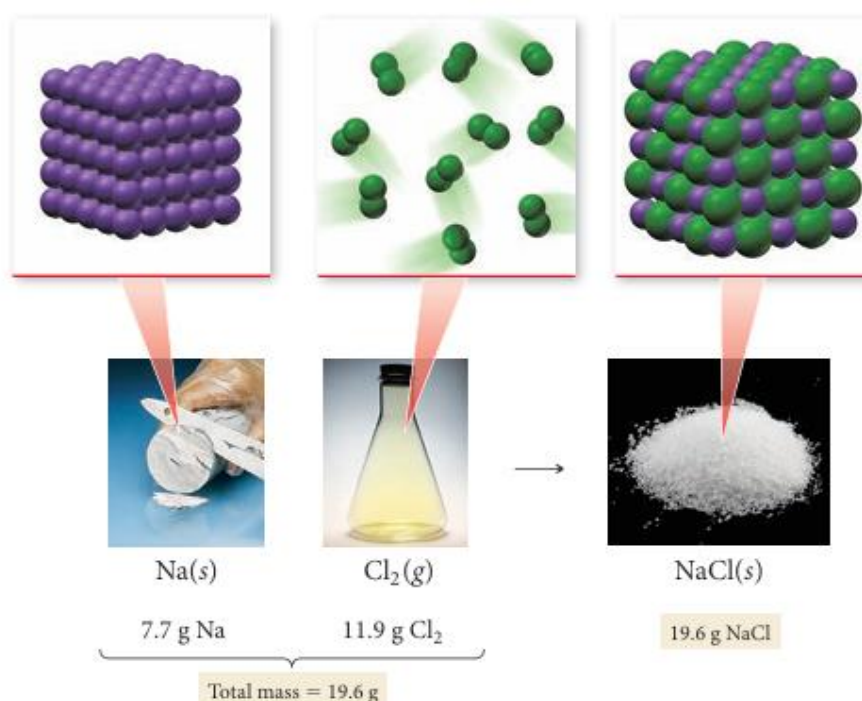
The atomic theory (the idea that all matter is composed of atoms) grew out of observations and laws. The three most important laws that led to the development and acceptance of the atomic theory are the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.



The Law of Conservation of Mass

In 1789, Antoine Lavoisier formulated the law of conservation of mass, which states: In a chemical reaction, matter is neither created nor destroyed.

In other words, when a chemical reaction occurs, the total mass of the substances involved in the reaction does not change. For example, consider the reaction between sodium and chlorine to form sodium chloride. The combined mass of the sodium and chlorine that reacts (the reactants) exactly equals the mass of the sodium chloride that forms (the product) as shown in Figure 1-1. This law is consistent with the idea that matter is composed of small, indestructible particles. The particles rearrange during a chemical reaction, but the amount of matter is conserved because the particles themselves are indestructible (at least by chemical means).



Mass of product = Mass of reactants

Figure 1-1: The reaction of sodium with chlorine

The Law of Definite Proportions

In 1797, the French chemist Joseph Proust made observations on the composition of compounds. He found that the elements composing a given compound always occur in fixed (or definite) proportions in all samples of the compound. In contrast, the components of a mixture can be present in any proportions whatsoever. Proust summarized his observations in the law of definite proportions: samples of a given compound, regardless of their source or how they were prepared, have the same proportions of as their constituent elements.

For example, the decomposition of 18.0 g of water results in 16.0 g of oxygen and 2.0 g of hydrogen, or an oxygen-to-hydrogen mass ratio of:

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

This ratio holds for any sample of pure water, regardless of its origin. The law of definite proportions applies to every compound. Consider ammonia, a compound composed of nitrogen and hydrogen. Ammonia contains 14.0 g of nitrogen for every 3.0 g of hydrogen, resulting in a nitrogen to hydrogen mass ratio of 4.7.

$$\text{Mass ratio} = \frac{14.0 \text{ g N}}{3.0 \text{ g H}} = 4.7 \text{ or } 4.7:1$$

Again, this ratio is the same for every sample of ammonia. The law of definite proportions also hints at the idea that matter is composed of atoms. Compounds have definite proportions of their constituent elements because the atoms that compose them, each with its own specific mass, occur in a definite ratio. Since the ratio of atoms is the same for all samples of a particular compound, the ratio of masses is also the same.

Example 1-1

Two samples of carbon dioxide are decomposed into their constituent elements. One sample produces 25.6 g of oxygen and 9.60 g of carbon, and the other produces 21.6 g of oxygen and 8.10 g of carbon. Show that these results are consistent with the law of definite proportions.

.....

.....

.....

.....

The Law of Multiple Proportions

In 1804, John Dalton published his law of multiple proportions:

When two elements (call them A and B) form different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers.

When two elements A and B combine to form more than one compound, an atom of A combines with either one, two, three, or more atoms of B (AB_1 , AB_2 , AB_3 , etc.). Therefore, the masses of B that react with a fixed mass of A are always related to one another as small whole-number ratios. Let's take the two compounds carbon monoxide and carbon dioxide, which are made of the same two elements: carbon and oxygen. We saw in the previous example that the mass ratio of oxygen to carbon in carbon dioxide is 2.67:1; therefore, 2.67 g of oxygen reacts with 1 g of carbon. In carbon monoxide, however, the mass ratio of oxygen to carbon is 1.33:1, or 1.33 g of oxygen to every 1 g of carbon.

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

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

With the help of the molecular models in the figure below, we can see why the ratio is 2:1.

Carbon dioxide contains two oxygen atoms to every carbon atom, while carbon monoxide contains only one. Of course, neither John Dalton nor Joseph Proust had access to any kind of modern instrumentation that could detect individual atoms. Dalton supported his atomic ideas primarily by using the masses of samples.

Carbon dioxide



Mass oxygen that combines
with 1 g carbon = 2.67 g

Carbon monoxide



Mass oxygen that combines
with 1 g carbon = 1.33 g

Example 1-2

Nitrogen forms several compounds with oxygen, including nitrogen dioxide and dinitrogen monoxide. Nitrogen dioxide contains 2.28 g oxygen to every 1.00 g nitrogen, while dinitrogen monoxide contains 0.570 g oxygen to every 1.00 g nitrogen. Show that these results are consistent with the law of multiple proportions.

.....

.....

.....

.....

Example 1-3

73 g of HCl gas is passed through a solution containing 158 g of sodium thiosulfate, resulting in 117 g of table salt, 64 g of SO₂ gas, 32 g of sulfur, and 18 g of water. Do the results support the law of conservation of mass?

.....

.....

.....

.....

Example 1-4

Two samples of SO₂ gas were analyzed. The first sample contained 3.6 g of oxygen and 1.8 g of sulfur, while the second sample contained 4.8 g of oxygen

and 4.2 g of sulfur. Do these results agree with the law of definite proportions?

.....

.....

.....

.....

.....

Example 1-5

Two unknown compounds were analyzed. The first compound contains 15 g of hydrogen and 120 g of oxygen, and the second compound contains 2 g of hydrogen and 32 g of oxygen. Are the two compounds the same? Explain your answer.

.....

.....

.....

.....

Example 1-6

A 20 g sample of SO_2 was reduced to a mass of 4 g of sulfur remaining. A 28 g sample of SO_3 was reduced to a mass of 4 g of sulfur remaining. Explain that this result is consistent with the law of multiple proportions.

.....

.....

.....

.....

Dalton's Atomic Theory

In 1808, John Dalton explained the laws we just discussed with in his atomic theory:

- Matter is made up of very small particles called atoms.
- Atoms are indivisible and indestructible.
- The atoms that make up an element are similar in size, mass, and chemical properties.
- Atoms of one element differ from atoms of other elements.
- Different atoms combine in a simple numerical ratio to form compounds.
- In chemical reactions: atoms separate, combine, or rearrange.

The Evolution of the Atom

The atom represents the basic building block of matter, and understanding its composition is key to understanding the chemical and physical properties of elements and their interactions. The concept of the atom has evolved from the idea of an "indivisible particle" (Dalton) to the modern model based on quantum mechanics.

Historical Atomic Models

The following diagram in Figure 1-2 shows the historical sequence of the development of our understanding of the structure of the atom:

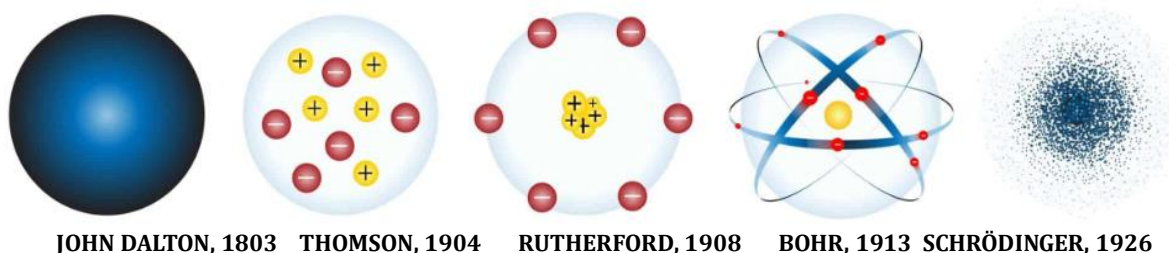


Figure 1-2: The historical sequence of the development of the structure of the atom.

The model	Discovery	Basic Concept
Dalton's model	The first scientific atomic theory	An atom is a solid, indivisible sphere.
Thomson's model	Discovery of the electron	Peach pie: a positive sphere interspersed with negative electrons.
Rutherford's model	Discovery of the proton.	An atom is mostly empty space, and the mass and positive charge are concentrated in the nucleus.
Bohr's model	Atomic emission spectrum of hydrogen	Electrons revolve in fixed orbits with specific quantized energy levels.
Quantum mechanical model	Principles of quantum mechanics	Orbitals are described as probability regions (electron clouds) for the presence of an electron.

Atomic Structure and Subatomic Particles

An atom consists of two main regions: the nucleus and the electron cloud.

The Nucleus

It is the center of the atom. It has a high density and a positive charge, and contains most of the atom's mass.

Subatomic Particles

An atom consists of three fundamental particles: protons, neutrons, and electrons.

Particle	Symbol	Location	Relative Charge	Relative Mass (approximately)
Proton	p^+	Nucleus	+1	1 amu
Neutron	n^0	Nucleus	0	1 amu
Electron	e^-	Electron cloud	-1	$\frac{1}{1840}$ amu

Completing the Atom Model: All atoms are made up of three basic subatomic particles: the electron, the proton, and the neutron. An atom is spherical in shape, containing a small, dense nucleus composed of positive charges surrounded by one or more negatively charged electrons. Most of the atom's volume is empty space containing fast-moving electrons, which move around the nucleus. The electrons are bound to the atom by attraction to the positive charges in the nucleus. The nucleus is composed of neutral neutrons (except for the nucleus of a hydrogen atom, which contains only one proton and no neutrons) and positively charged protons. The nucleus constitutes more than 99.97% of the atom's mass and occupies about 0.0001% of the atom's volume. Because the atom is electrically neutral, the number of protons in the nucleus equals the number of electrons surrounding it. Figure 1-3 shows the components of an atom.

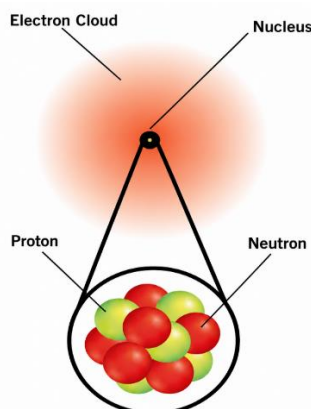
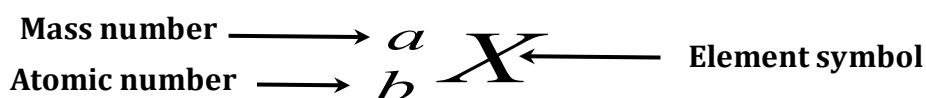


Figure 1-3: An atom consists of a nucleus containing protons and neutrons surrounded by a cloud of electrons.

The components of an atom continue to be of interest to many modern scientists. In fact, scientists have determined that protons and neutrons have their own composition, consisting of particles called quarks. Chemical behavior is explained by the electrons of an atom, as you will study later.

Atomic number and mass

العدد الذري والعدد الكتلي number



Atomic number = number of protons in an atom = number of electrons

Mass number = number of protons + number of neutrons

Number of neutrons = mass number – number of protons

Number of protons = mass number – number of neutrons

Isotopes

All atoms of a given element have the same number of protons; however, they do not necessarily have the same number of neutrons. Since the mass of neutron is approximately equal to the mass of proton (one atomic unit), this means that not all atoms of a given element have the same mass. For example, all hydrogen atoms have one proton, but they may have 0, 1, or 2 neutrons. All three types of hydrogen (protium, deuterium, and tritium) atoms exist, each with a slightly different mass. Atoms with the same number of protons but different numbers of neutrons are known as isotopes.

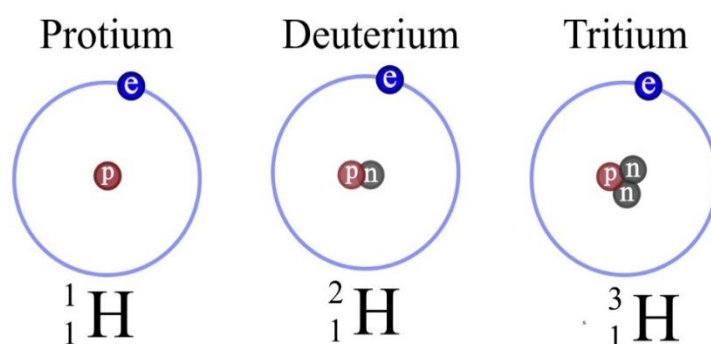


Figure 1-4: Hydrogen isotopes

Some elements, such as beryllium (Be) and aluminum (Al), have only one naturally occurring isotope, while other elements, such as neon (Ne) and chlorine (Cl), have two or more.

The relative amount of each different isotope in a naturally occurring sample of a given element is roughly constant. For example, in any natural sample of neon atoms, 90.48 % of them are the isotope with 10 neutrons, 0.27% are the

isotope with 11 neutrons, and 9.25% are the isotope with 12 neutrons. These percentages are the natural abundance of the isotopes. Each element has its own characteristic natural abundance of isotopes.

Example 1-7

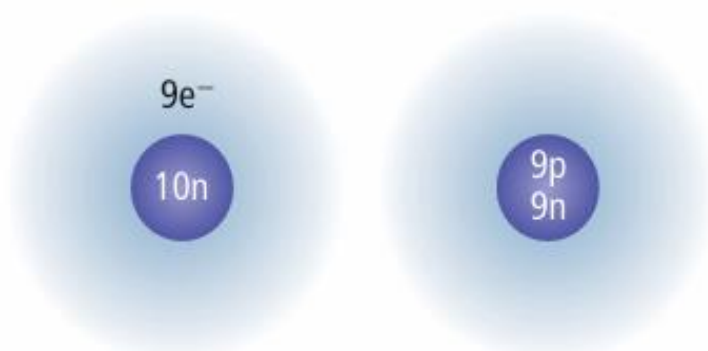
Complete the following two tables: (Use the periodic table)

Element name	Element symbol	Atomic number	Mass number	Number of protons	Number of electrons	Number of neutrons	Isotope name	Isotope symbol
								$^{70}_{30}\text{Zn}$
Calcium	Ca	20	46					
Iron		26	57					
Mercury		80	204					

Isotope	Zn - 64				
Atomic Number				9	11
Mass Number			32		23
Number of Protons			16		
Number of Neutrons		24		10	
Number of Electrons		20			

Example 1-8

Do the atoms shown in the figure below represent the same element? Why?



.....

Mass of Atoms

The mass of a proton and a neutron is approximately 1.67×10^{-27} kg, and the mass of an electron is much smaller, about $\frac{1}{1840}$ the mass of a proton or neutron.

Atomic Mass Unit

Because these masses are so small and difficult to manipulate, scientists developed a new method for measuring the mass of an atom relative to the mass of a standard atom. This standard atom is carbon, which has an atomic mass of 12. Therefore, the atomic mass unit (amu) is $\frac{1}{12}$ the mass of a carbon-12 atom. Therefore, an atomic mass unit is approximately equal to the mass of one proton or one neutron.

Atomic Mass

Because the mass of an atom depends primarily on the number of protons and the number of neutrons in it, and because the mass of both a proton and a neutron is close to 1 amu, you might expect that the atomic mass of an element would always be a whole number! But this is not true; the atomic mass of an element is the average of the masses of its isotopes. Because isotopes have different masses, the average atomic mass is will not be always a whole number. The figure 1-5 shows the calculation of the atomic mass of chlorine.

Chlorine occurs in nature as a mixture of 75.78% chlorine-35 and 24.22% chlorine-37. The atomic mass of chlorine is 35.453 amu.

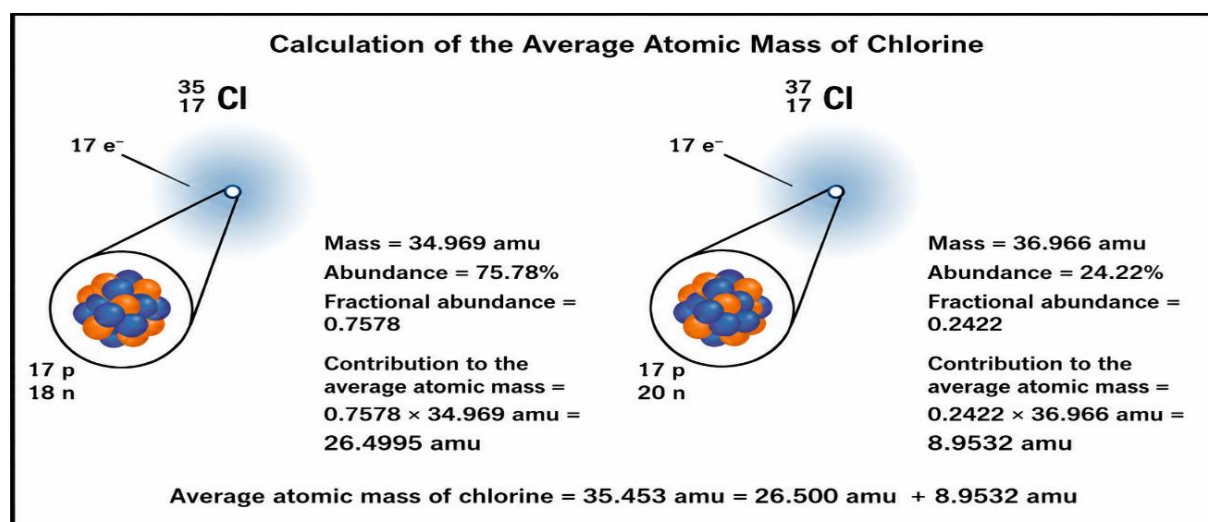


Figure 1-5: Calculating the average atomic mass of the element chlorine

Electron Configuration

Ground -State Electron Configuration

The arrangement of electrons in an atom is called electron configuration. Because lower-energy systems are more stable than higher-energy systems, electrons tend to adopt an arrangement that gives the atom the lowest possible energy. The arrangement of electrons in the lowest, most stable energy position is called the electron configuration in the element's ground state. Principles or rules—including the Aufbau principle, the Pauli principle, and Hund's rule—govern how electrons are arranged in atomic levels.

The Aufbau principle (building upward) states that every electron occupies the lowest energy level. Therefore, determining the ground state electron configuration requires knowing the arrangement of sublevels according to their increasing energy. This sequence is known as the Aufbau diagram, and is illustrated in the Figure 1-6, where each box in the figure represents a sublevel.

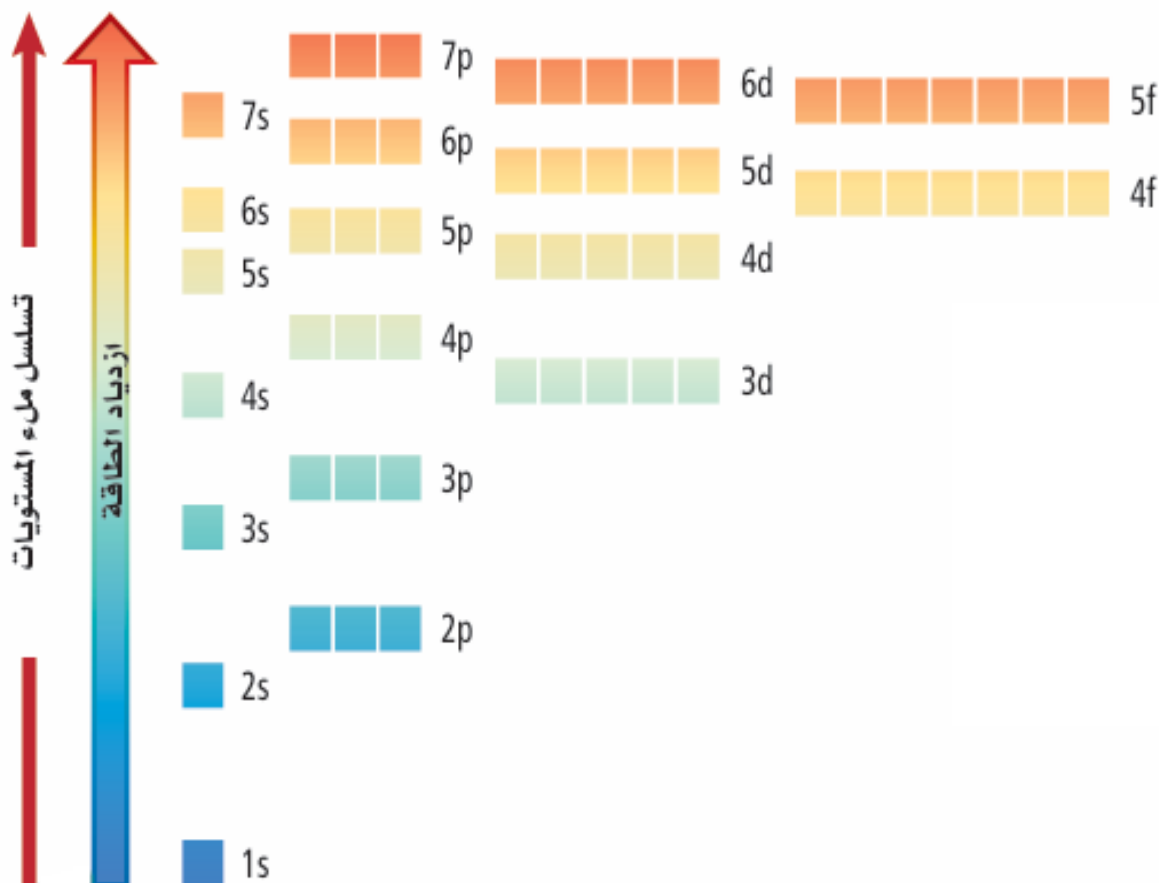


Figure 1-6: Aufbau drawing

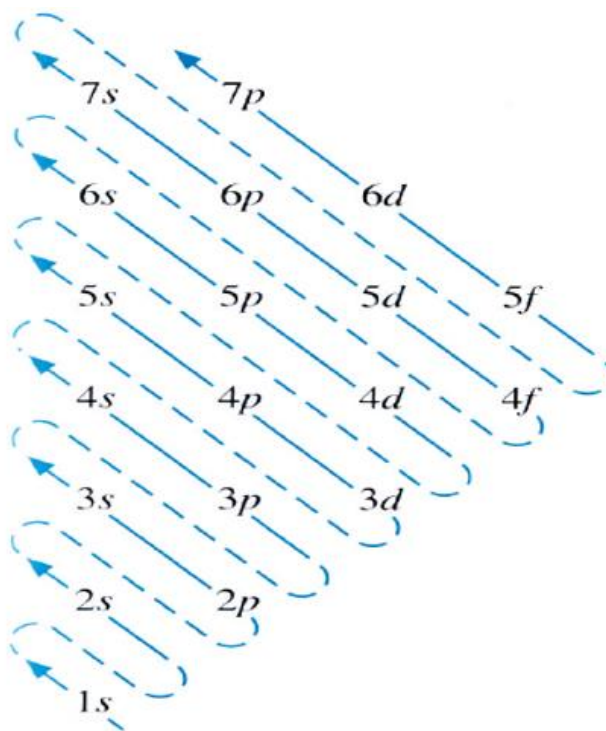


Figure 1-7: Order of filling levels with electrons

The following table shows the properties of the Aufbau diagram:

Example	Feature
All three 2p orbitals are of equal energy.	All orbitals related to an energy sublevel are of equal energy.
The three 2p orbitals are of higher energy than the 2s orbital.	In a multi-electron atom, the energy sublevels within a principal energy level have different energies.
If $n=4$, then the sequence of energy sublevels is 4s, 4p, 4d, and 4f.	In order of increasing energy, the sequence of energy sublevels within a principal energy level is s, p, d, and f.
The orbital related to the atom's 4s sublevel has a lower energy than the five orbitals related to the 3d sublevel.	Orbitals related to energy sublevels within one principal energy level can overlap orbitals related to energy sublevels within another principal level.

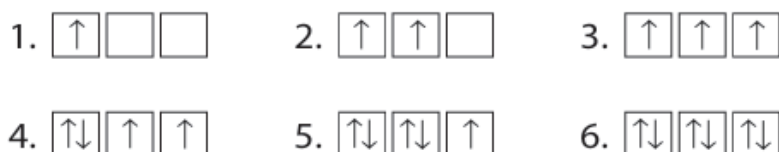
Pauli's Principle: Pauli's principle states that no two electrons in an atom can have the same four quantum numbers .

Sublevels can be represented by squares or circles, and electrons within levels can be represented by arrows within the squares. Each electron has an

associated spin direction, with an upward arrow $\uparrow$ representing the electron's spin in one direction, and a downward arrow $\downarrow$ representing the electron's spin in the opposite direction.

An empty square $\square$ represents a vacant sublevel, a square with one arrow pointing up $\uparrow$ represents a sublevel with one electron, and a square with two arrows, one pointing up and one down $\uparrow\downarrow$, represents a filled sublevel.

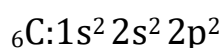
Hund's Rule: The fact that negatively charged electrons repel each other has a significant effect on the configuration of electrons into equal-energy sublevels. Hund's rule states that electrons are distributed in equal-energy sublevels so that they maintain the same spin orientation before additional electrons with the opposite spin orientation occupy the same levels. For example, the three 2p orbitals are filled with single electrons, and then pairing occurs. The following figure illustrates the sequence of six electrons entering the p sublevels.



Methods for representing the electron configuration

You can represent the electron configuration of an atom in one of the following ways: electron notation, Orbital diagram, or noble gas notation.

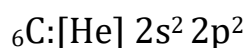
1- Electron Notation:



2- Orbital Diagram:



3- Noble Gas Notation:



Example 1-9

Write the electronic Configuration of magnesium ${}_{12}\text{Mg}$ in three ways?

.....

.....

.....

.....

.....

Electron configuration of positive and negative ions

When an atom loses electrons, it forms a positive ion, and when it gains electrons, it forms a negative ion. The following table shows the electron configuration of some atoms and their ions

H: $1s^1$	H ⁻ : $1s^2$ or [He]
F: $1s^2 2s^2 2p^5$	F ⁻ : $1s^2 2s^2 2p^6$ or [Ne]
O: $1s^2 2s^2 2p^4$	O ²⁻ : $1s^2 2s^2 2p^6$ or [Ne]
N: $1s^2 2s^2 2p^3$	N ³⁻ : $1s^2 2s^2 2p^6$ or [Ne]
Na: $1s^2 2s^2 2p^6 3s^1$	Na ⁺ : $1s^2 2s^2 2p^6$ or [Ne]
Ca: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	Ca ²⁺ : $1s^2 2s^2 2p^6 3s^2 3p^6$ or [Ar]
Al: $1s^2 2s^2 2p^6 3s^2 3p^1$	Al ³⁺ : $1s^2 2s^2 2p^6$ or [Ne]

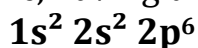
Isoelectronic

When different atoms or ions possess the same total number of electrons and the same electron configuration in the ground state, we say they are isoelectronic.

Examples of electronic isomorphism:

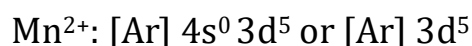
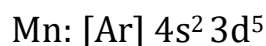
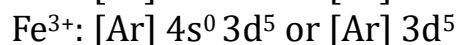
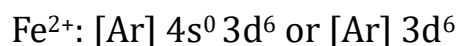
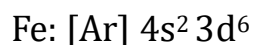
- The oxide ion O²⁻ has 10 electrons.
- The fluoride ion F⁻ has 10 electrons.
- The neon atom Ne has 10 electrons.
- The sodium ion Na⁺ has 10 electrons.

These species are all isoelectronic, having the same electron configuration.



Transition Element Ions

For transition elements, they first lose two electrons from the **s** orbital, then from the **d** orbital.



Exceptions to predicted configuration

You can use the aufbau diagram to write correct ground-state electron configurations for all elements up to and including vanadium, atomic

number 23. However, if you were to proceed in this manner, your configurations for Chrome, $[\text{Ar}] 4s^2 3d^4$, and copper, $[\text{Ar}] 4s^2 3d^9$, would be incorrect. The correct configurations for these two elements are $[\text{Ar}] 4s^1 3d^5$ for Chrome and $[\text{Ar}] 4s^1 3d^{10}$ for copper. The electron configurations for these two elements, as well as those of several other elements, illustrate the increased stability of half-filled and filled sets of s and d orbitals.

Quantum Numbers

When we solve the Schrödinger equation for the hydrogen atom, we find many wave functions (orbitals) that satisfy it. Each of these orbitals is characterized by a series of numbers called quantum numbers, which describe various properties of the orbital:

Principal quantum number (n): Determines the distance of an electron from the nucleus and indicates the energy of the electron and, consequently, the energy of the principal energy level in which it moves. It takes one of the positive integer values: 1, 2, 3, etc. Principal energy levels are sometimes indicated by letters, as follows:

Principal Quantum Number	1	2	3	4	5	6	7
Letter	K	L	M	N	O	P	Q

The maximum number of electrons that can exist in each principal energy level in an atom can be determined from the relationship: $2n^2$.

secondary quantum number (ℓ): has integral values from 0 to n-1 for each value of n. This quantum number is related to the shape of atomic orbitals. The value of ℓ for a particular orbital is commonly assigned a letter: $\ell = 0$ is called s; $\ell = 1$ is called p; $\ell = 2$ is called d; $\ell = 3$ is called f. This system arises from early spectral studies and is summarized in following Table.

Secondary quantum numbers and corresponding letters used to denote atomic orbitals				
Value of ℓ	0	1	2	3
Letter Used	s	p	d	f
Sublevel shape	Spherical	Dumbbell-shaped	Complex	More complex

The magnetic quantum number (m_ℓ): It expresses the direction of the orbitals that make up each sublevel in space and takes integer values from $(-\ell)$ to $(+\ell)$ passing through zero. This means that its values are related to the values of the secondary quantum number. The values of the magnetic quantum number also express the number of fields (electron rooms) present in each sublevel, and it was found that this number is equal to $2\ell + 1$. The

following table shows the relationship between ℓ and m_ℓ and the number of fields.

ℓ	m_ℓ	Number of orbitals
0	0	1
1	-1, 0, +1	3
2	-2, -1, 0, +1, +2	5
3	-3, -2, -1, 0, +1, +2, +3	7

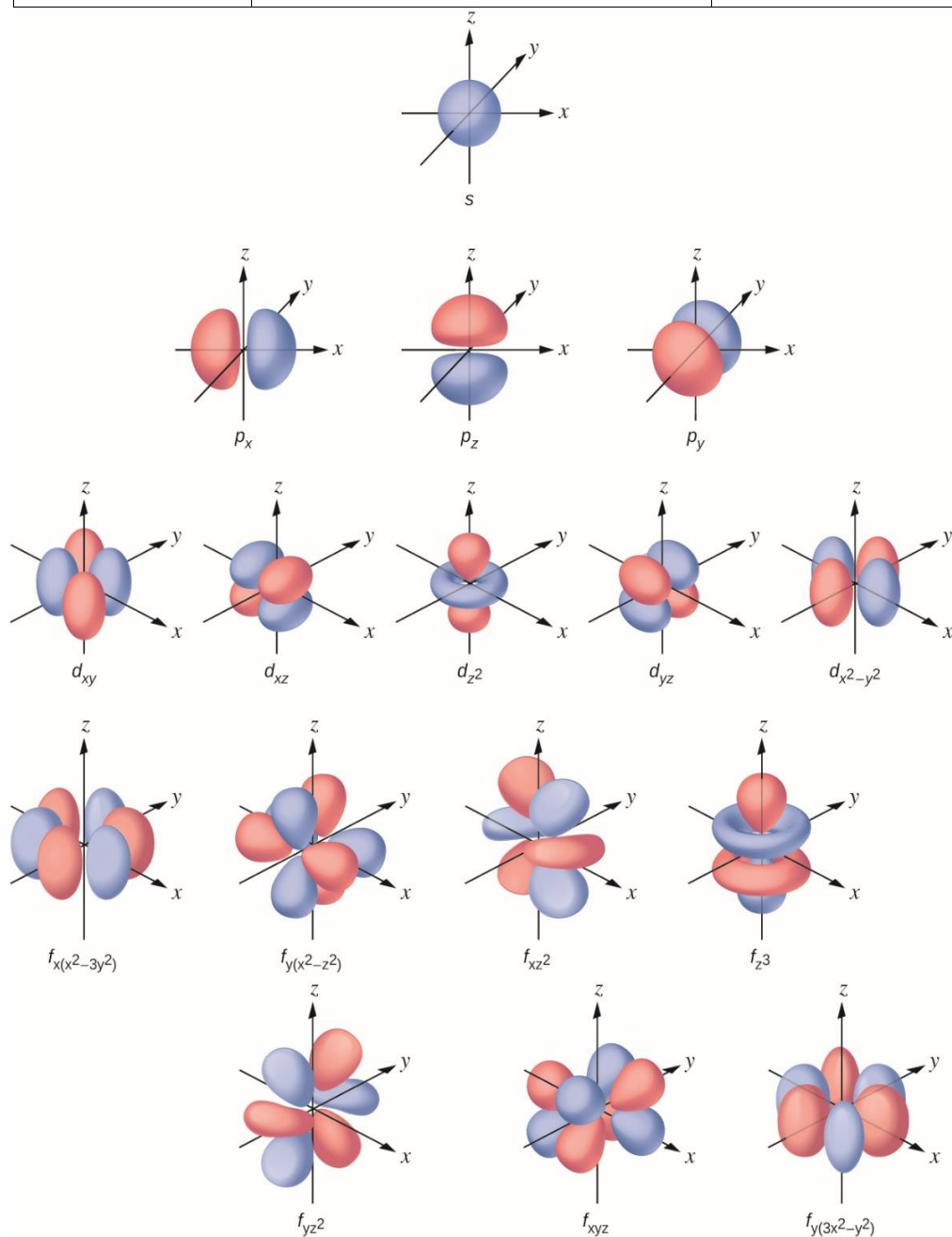


Figure 1-8: Shapes of s, p, d, and f orbitals

Spin quantum number (m_s): It describes how an electron moves around itself (its axis), either clockwise or counterclockwise. There are two values for this number: $+\frac{1}{2}$ and $-\frac{1}{2}$.

It is conventional to represent an electron with ($m_s = +\frac{1}{2}$) by an arrow pointing upward ($\uparrow$), and an electron with ($m_s = -\frac{1}{2}$) by an arrow pointing downward ($\downarrow$), as in the following figure:

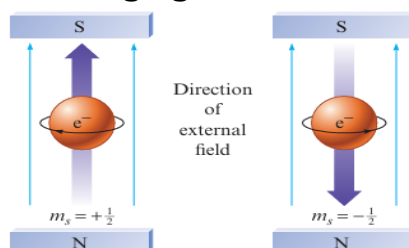


Figure 1-9: Electron Spin and Magnetic Field Direction

Since the spin of an electron is limited to only two directions, when two electrons are in the same field, the spin of one is opposite to the other. This creates two magnetic fields with opposite directions, which attract each other magnetically. This reduces the repulsive force between the negative charges of the electrons in the field, leading to relative stability for the atom.

Since no orbital in an atom can accommodate more than two electrons with opposite spins, this allows any sublevel to be completely filled with a number of electrons equal to twice the number of sublevel orbitals. The maximum number of electrons that sublevels can accommodate can be explained as follows:

Sublevel symbol	Number of orbitals	Maximum number of electrons
s	1	2
p	3	6
d	5	10
f	7	14

By determining the four quantum numbers (n, ℓ, m_ℓ, m_s) for each electron, we arrive at knowing the field in which the electron is located and the direction in which it rotates.

Example 1-10

Complete the following table:

n value	ℓ value	Sublevel symbol	m_ℓ value	m_s value
1				
		2s 2p		
	0 1 2			

Exercises

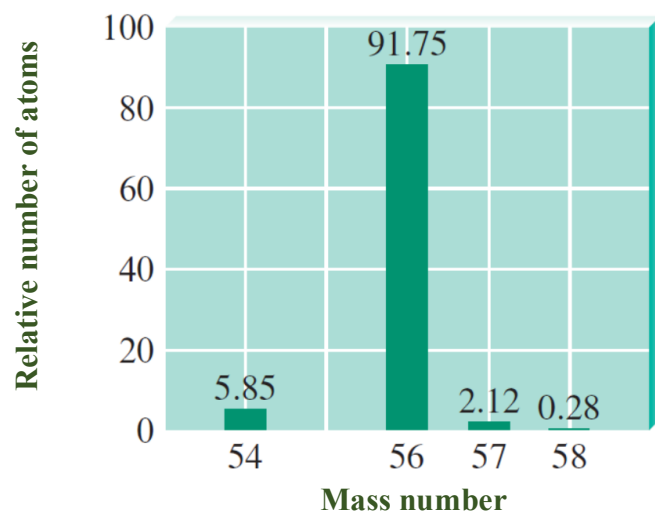
(1-1)

Complete the following table:

Symbol	Number of Protons	Number of Neutrons	Number of Electrons	Net Charge
$^{238}_{92}\text{U}$				
	20	20		+ 2
	23	28	20	
$^{89}_{39}\text{Y}$				
	35	44	36	
	15	16		- 3

(1-2)

The stable isotopes of iron are ^{54}Fe , ^{56}Fe , ^{57}Fe , and ^{58}Fe . The mass spectrum of iron looks like the following:



Use the data on the mass spectrum to estimate the average atomic mass of iron, and compare it to the value given in the Periodic Table.

.....

.....

.....

.....

(1-3)

Nitrogen has three isotopes in nature: ^{14}N , ^{15}N and ^{16}N . To account for nitrogen's atomic mass of 14.00674, what must be the percentage of ^{15}N atoms in natural nitrogen? Ignore the small amount of ^{16}N .

.....

.....

.....

.....

(1-4)

Which of the following fourth-period atoms has the greatest number of Unpaired electrons in its lowest energy state?

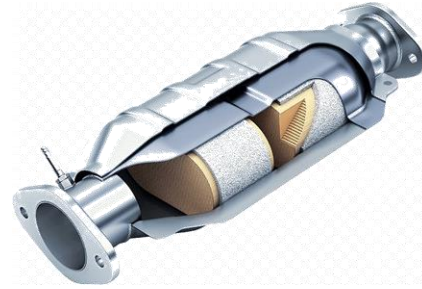
2 2A											13 3A
4 Be 9.012											5 B 10.81
12 Mg 24.31	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 Al 26.98
20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72
			↑	↑	↑			↑			

- (a) Zinc, Zn
- (b) Nickel, Ni
- (c) Manganese, Mn and Vanadium, V
- (d) Chrome, Cr

(1-5)

Palladium Pd (Atomic number 46) is a transition metal used in the catalytic converter at car exhausts to reduce pollution. Which set of quantum numbers could NOT correspond to an electron in a ground-state phase Pd atom?

- (a) $n = 2, \ell = 1, m_\ell = -1, m_s = +\frac{1}{2}$
 (b) $n = 5, \ell = 2, m_\ell = 0, m_s = +\frac{1}{2}$
 (c) $n = 4, \ell = 0, m_\ell = 0, m_s = -\frac{1}{2}$
 (d) $n = 3, \ell = 2, m_\ell = -1, m_s = -\frac{1}{2}$

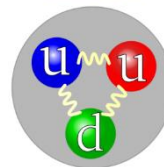


(1-6)

Scientists have identified particles called "quarks" that combine to form protons and neutrons inside the nucleus. If you know that a single proton consists of a union of two "upper quarks" with a charge $(+\frac{2}{3})$ and one of the "lower quarks" with a charge $(-\frac{1}{3})$; What is a neutron composed of?

(Note: the upper quark is denoted by the letter "u" and the lower quark is denoted by the letter "d").

Proton البروتون



- (a) $1u + 1d$ (b) $2u + 1d$
 (c) $2d + 1u$ (d) $4d + 2u$

(1-7)

In the ground state of element 115, Mc:

- 1) how many electrons have $n = 5$ as one of their quantum numbers?
- 2) how many electrons have $\ell = 1$ as one of their quantum numbers?
- 3) how many electrons have $m_\ell = 1$ as one of their quantum numbers?
- 4) how many electrons have $m_s = -1/2$ as one of their quantum numbers?

.....

.....

.....

.....

(1-8)

Determine the maximum number of electrons that can have each of the following designations: $3f$, $2d_{xy}$, $3p$, $5d_{yz}$, and $4p$.

$3f$	$2d_{xy}$	$3p$	$5d_{yz}$	$4p$

(1-9)

Arrange the following electron configurations in order of increasing stability:

$4s^2 3d^{10}$, $4s^2 3d^5$, $4s^2 3d^{10} 4p^6$, $4s^2 3d^6$

.....

.....

.....

.....

Example answers

Example 1-1

For the first sample:

$$\frac{\text{Mass oxygen}}{\text{Mass carbon}} = \frac{25.6}{9.60} = 2.67 \text{ or } 2.67:1$$

For the second sample:

$$\frac{\text{Mass oxygen}}{\text{Mass carbon}} = \frac{21.6}{8.10} = 2.67 \text{ or } 2.67:1$$

The ratios are the same for the two samples, so these results are consistent with the law of definite proportions.

Example 1-2

$$\frac{\text{Mass oxygen to 1 g nitrogen in nitrogen dioxide}}{\text{Mass oxygen to 1 g nitrogen in dinitrogen monoxide}} = \frac{2.28}{0.570} = 4$$

The ratio is a small whole number (4); these results are consistent with the law of multiple proportions.

Example 1-3

Total mass of the reactants = 158 g + 73 g = **231 g**

Total mass of the products = 117 g + 64 g + 32 g + 18 g = **231 g**

This is consistent with the **Law of Conservation of Mass**.

Example 1-4

First sample: Ratio of the mass of oxygen to the mass of sulfur = $\frac{3.6 \text{ g (O)}}{1.8 \text{ g (S)}} = 2$

Second sample: Ratio of the mass of oxygen to the mass of sulfur = $\frac{4.8 \text{ g (O)}}{4.2 \text{ g (S)}} = 1.14$

Since the ratio is different in the two samples, this indicates that the results **do not agree with the Law of Definite Proportions**

Example 1-5

First sample: %H = 11.1% %O = 88.9%

Second sample: %H = 5.9% %O = 94.1%

The samples are **not of the same compound** because their percentage compositions are different.

Example 1-6

$$\frac{\text{Mass ratio for compound 1}}{\text{Mass ratio for compound 2}} = \frac{\frac{16}{4}}{\frac{24}{4}} = \frac{4}{6} = \frac{2}{3} = 3:2$$

Example 1-7

Complete the following two tables: (Use the periodic table)

Element name	Element symbol	Atomic number	Mass number	Number of protons	Number of electrons	Number of neutrons	Isotope name	Isotope symbol
Zinc	Zn	30	70	30	30	40	Zinc-70	$^{70}_{30}\text{Zn}$
Calcium	Ca	20	46	20	20	26	Calcium-46	$^{46}_{20}\text{Ca}$
Iron	Fe	26	57	26	26	31	Iron-57	$^{57}_{26}\text{Fe}$
Mercury	Hg	80	204	80	80	124	Mercury-204	$^{204}_{80}\text{Hg}$

Isotope	Zn - 64	Ca-44	S-32	F-19	Na-23
Atomic Number	30	20	16	9	11
Mass Number	64	44	32	19	23
Number of Protons	30	20	16	9	11
Number of Neutrons	34	24	16	10	12
Number of Electrons	30	20	16	9	11

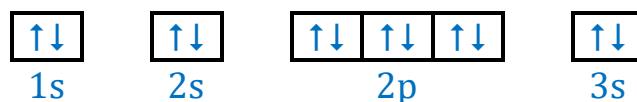
Example 1-8

Yes, because they have the same atomic number.

Example 1-9

1- electron notation : $_{12}\text{Mg}:1s^2 2s^2 2p^6 3s^2$

2- Orbital diagram :



3- gas notation : $_{12}\text{Mg}:[\text{Ne}]3s^2$

Example 1-10

n value	ℓ value	Sublevel symbol	m_ℓ value	m_s value
1	0	1s	0	$\pm\frac{1}{2}$
2	0	2s	0	$\pm\frac{1}{2}$
	1	2p	-1, 0, +1	$\pm\frac{1}{2}$
3	0	3s	0	$\pm\frac{1}{2}$
	1	3p	-1, 0, +1	$\pm\frac{1}{2}$
	2	3d	-2, -1, 0, +1, +2	$\pm\frac{1}{2}$

Exercise answers

إجابات التدريبات

(1-1)

Symbol	Number of Protons	Number of Neutrons	Number of Electrons	Net Charge
${}_{92}^{238}\text{U}$	92	146	92	0
${}_{20}^{40}\text{Ca}^{2+}$	20	20	18	2+
${}_{23}^{51}\text{V}^{3+}$	23	28	20	3+
${}_{39}^{89}\text{Y}$	39	50	39	0
${}_{35}^{79}\text{Br}^{-}$	35	44	36	1-
${}_{15}^{31}\text{P}^{3-}$	15	16	18	3-

(1-2)

The calculated average atomic mass would be:

$$0.0585(54.00) + 0.9175(56.00) + 0.0212(57.00) + 0.0028(58.00) = 55.91 \text{ amu.}$$

The average atomic mass listed in the periodic table is 55.845 amu.

(1-3)

$$14.00674 = (X \times 15) + ((1 - X) \times 14)$$

$$14.00674 = 15X + 14 - 14X$$

$$14.00674 = X + 14$$

$$X = 0.00674$$

$$\% {}^{15}\text{N} = 0.00674 \times 100 = 0.674 \%$$

(1-4)

(d) Chrome, Cr

(1-5)

(b) $n = 5, l = 2, m_l = 0, m_s = +\frac{1}{2}$

(1-6)

(c) $2d + 1u$

(1-7)

Mc: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^{10} 7p^3$

1) $5s^2, 5p^6, 5d^{10}$, and $5f^{14}$; **32** electrons have $n = 5$ as one of their quantum numbers.

2) $\ell = 3$ are (f) orbitals. $4f^{14}$ and $5f^{14}$ are the f orbitals used. They are all filled, so **28** electrons have $\ell = 3$.

3) p, d, and f orbitals all have one of the degenerate orbitals with $m_\ell = 1$. There are 6 orbitals with $m_\ell = 1$ for the various p orbitals used; there are 4 orbitals with $m_\ell = 1$ for the various d orbitals used; and there are 2 orbitals with $m_\ell = 1$ for the various f orbitals used. We have a total of $6 + 4 + 2 = 12$ orbitals with $m_\ell = 1$. Eleven of these orbitals are filled with 2 electrons, and the 7p orbitals are only half-filled. The number of electrons with $m_\ell = 1$ is $11 \times (2 e^-) + 1 \times (1 e^-) = \mathbf{23}$ electrons.

4) The first 112 electrons are all paired; one-half of these electrons ($56 e^-$) will have $m_s = -\frac{1}{2}$. The 3 electrons in the 7p orbitals singly occupy each of the three degenerate 7p orbitals; the three electrons are spin parallel, so the 7p electrons either have $m_s = +\frac{1}{2}$ or $m_s = -\frac{1}{2}$. Therefore, either **56** electrons have $m_s = -\frac{1}{2}$ or **59** electrons have $m_s = -\frac{1}{2}$.

(1-8)

3f	2d _{xy}	3p	5d _{yz}	4p
0	0	6	2	6

(1-9)

$$4s^2 3d^6 < 4s^2 3d^5 < 4s^2 3d^{10} < 4s^2 3d^{10} 4p^6$$

02

The Periodic Table and Classification of Elements

Classification of Elements

Many attempts have been made to classify elements into tables:

- To facilitate the study of elements with similar properties in one group.
- Berzelius' attempt to divide the elements into metals, non-metals, and metalloids.
- Mendeleev arranged the elements in his first periodic table in increasing order of atomic mass, and by doing so he was able to predict the existence of elements that had not yet been discovered. Moseley later demonstrated that the correct arrangement should be based on atomic number, which is the principle used in the modern periodic table.

Modern construction of the periodic table

1) **It consists of periods:** horizontal rows starting from the first period to the seventh.

2) **It consists of groups:** vertical columns numbered from 1 to 18, and each group contains elements that share a common number of valence electrons; thus, their properties are similar.

groups

1 2 13 14 15 16 17 18

3 4 5 6 7 8 9 10 11 12

Division of the elements of the periodic table:

- **Metals:** Most elements are of this type.
- **Non-metals:** are the elements found on the right side of the periodic table.
- **Metalloids:** There are 7 elements found between metals and non-metals in a zigzag row.

The diagram shows a simplified periodic table with the following classification:

- Metals (Blue):** Includes all elements in the first two columns (s-block), the transition metals in the middle, and the lanthanides and actinides at the bottom.
- Non-metals (Green):** Includes elements in the upper right corner, specifically the groups containing Boron, Carbon, Nitrogen, Oxygen, Fluorine, and the noble gases.
- Metalloids (Yellow):** Includes elements along the diagonal line separating metals and non-metals, specifically Boron, Silicon, and Germanium.

Important notes:

- The elements of main groups 1, 2, 13 to 18 are referred to as the representative elements.
- Elements in groups 3 to 12 are referred to as transition elements.
- Group 1 elements (except hydrogen) are called alkali metals.
- Group 2 elements are called alkaline earth metals.
- Transition elements are divided into transition metals and inner transition metals.
- The inner transition metals are known as the lanthanide and actinide series.
- The elements in group 17 are known as halogens.

- The elements in group 18 are known as the noble gases.

Physical states of the elements of the periodic table:

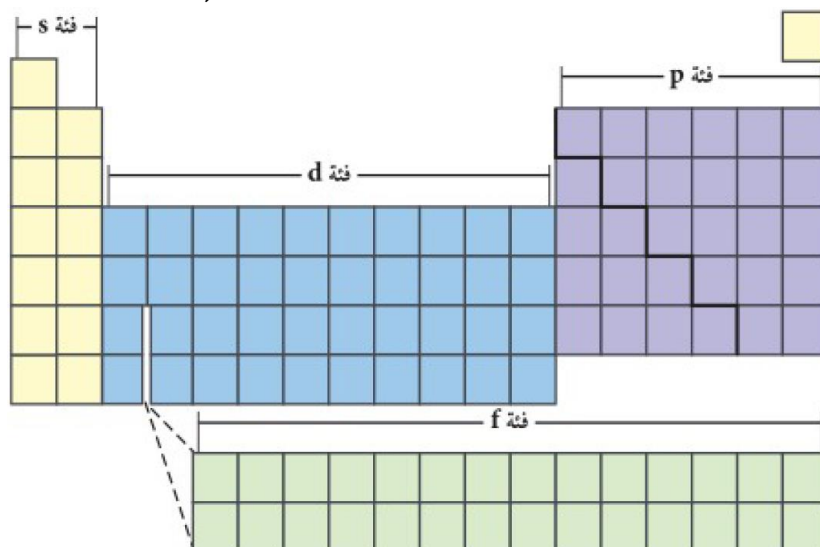
- Gases: F_2 , O_2 , N_2 , H_2 , Cl_2 , which are non-metals.
- Liquid: Bromine Br_2 (non-metal), Mercury Hg (the only metal that is liquid at room temperature).
- Solid elements: the rest of the elements in the periodic table (mostly metals).
- The inert gases are He, Ne, Ar, Kr, Xe, and Rn. They have several uses, especially in the manufacture of lamps. They are called this because they are chemically inert (their final level is completely full).

Physical properties of metals, nonmetals and metalloids:

Metals	Nonmetals	Metalloids
Solid at room temperature except mercury	Including gas, liquid, and solid	Generally, have properties intermediate between those of metals and non-metals.
Their melting points are high.	Its melting and boiling points are low.	
It has a shine and luster.	It has no shine	
conductor of heat and electricity	Non-conductive to heat and electricity	
Malleable and ductile.	The steel is brittle and easy to break.	

Periodic table and electron configuration

The periodic table is divided into categories that represent the secondary energy levels of an atom, which contain the valence electrons.



The above periodic table shows a classification of its elements based on block categories.

s-block

The first column is given a number, which is (1). They are called alkali metals, and they have a general electronic configuration of (ns^1) .

The second column was given a number (2) and was called the alkali earth metals, and it has a general electronic configuration of (ns^2) .

p-block

The first column in the p-block is given a number, which is (13). They are called Boron group and have a general electronic configuration (ns^2np^1) .

The second column in the p-block is given a number, which is (14). It is called the carbon group elements and has a general electronic configuration of (ns^2np^2) .

The third column in the p-block was given a number, which is (15). It was named the nitrogen group elements, and it has a general electronic configuration of (ns^2np^3) .

The fourth column in the p-block is given a number, which is (16). It is called the oxygen group elements, and it has a general electronic configuration of (ns^2np^4) .

The fifth column in the p-block is given a number, which is (17). It is called the halogen group elements, and it has a general electronic configuration of (ns^2np^5) .

The sixth column in the p-block is given a number, which is (18). It is called the noble gas group elements, and it has a general electronic configuration of (ns^2np^6) .

d-block

Column (3) in the d-block has a general electronic configuration of $(ns^2(n-1)d^1)$.

Column (4) in the d-block has a general electronic configuration of $(ns^2(n-1)d^2)$.

Column (5) in the d-block has a general electronic configuration of $(ns^2(n-1)d^3)$.

Column (6) in the d-block has a general electronic configuration of $(ns^1(n-1)d^5)$.

Here an exception occurs and the reason for this is that the half-filled $((n-1)d^5)$ sublevels are slightly more stable and the electrons in the same sublevel (d) have the same energy.

Column (7) in the d-block has a general electronic configuration of $(ns^2(n-1)d^5)$.

Column (8) in the d-block has a general electronic configuration of $(ns^2(n-1)d^6)$.

Column (9) in the d-block has a general electronic configuration of $(ns^2(n-1)d^7)$.

Column (10) in the d-block has a general electronic configuration of $(ns^2(n-1)d^8)$.

Column (11) in the d-block has a general electronic configuration of $(ns^1(n-1)d^{10})$.

An exception occurs here as well, due to the slightly greater stability of the filled $(n-1)d^{10}$ sublevels, and the electrons in the same sublevel (d) have the same energy.

Column (12) in the d-block has a general electronic configuration of $(ns^2(n-1)d^{10})$.

- ❖ To determine the period number and group number of a chemical element, you can rely on the electron configuration of the element.

Specify the period number:

- The period number represents the highest principal energy level (n) in the electron configuration of the element.

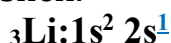


The period: **Second**

Specify group number:

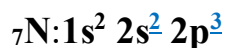
The group number is determined by the last sublevel in the electron configuration of the element:

1. If the configuration ends in s: The group number is the number of electrons in the last s subshell.



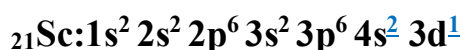
Group: **First**

2. If the configuration ends with p: The group number is the sum of the number of electrons in the last s and p subshells, plus 10.



Group: **Fifteen**

3. If the configuration ends with d: The group number is the sum of the number of electrons in the last s and d subshells.



Group: **Third**

4. For elements whose electron configuration ends at the f subshell (inner transition elements), the group number is determined differently than for other elements.
- **Group number:** All f-block elements are often considered part of Group 3. Individual group numbers are not assigned to these elements because they have very similar chemical properties.
 - **Period number:** To determine the period number, follow the general rule: The period number is the highest number for the principal energy level in the electron configuration.

Properties of the periodic table

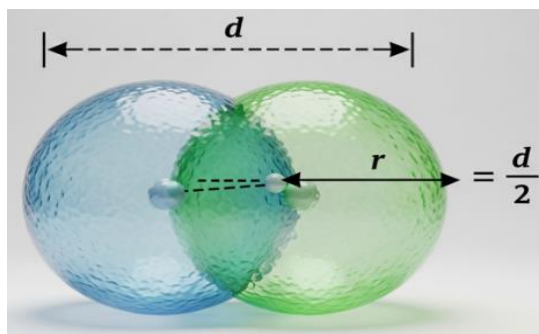
The electron configuration of elements exhibits a periodic change with increasing atomic number, leading to periodic changes in chemical and physical behavior. We will discuss some of these properties.

❖ Atomic size:

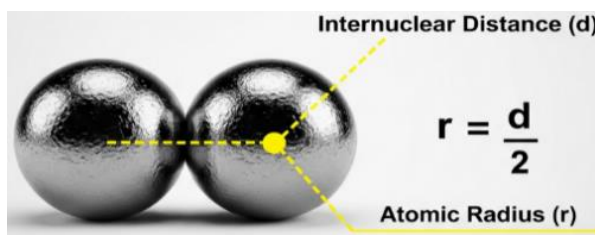
The atomic size is a measure of the size of an atom, and it is commonly defined as the distance between the center of the nucleus and the outermost electron shell of the atom. The atomic size - usually described in terms of the atomic radius - indicates how large the atom is.

Because an atom has no clear outer boundaries, its size is not constant, but varies depending on how it bonds with other atoms. Therefore, atomic size is determined in different ways based on the type of chemical bonds the atom forms.

- Covalent radius: Used for elements that form covalent bonds. It is measured by taking half the distance between the nuclei of two identical atoms bonded by a single covalent bond, as in the case of a hydrogen (H_2) or chlorine (Cl_2) molecule.



- Metallic radius: Used for metallic elements. It is measured by taking half the distance between the nuclei of two adjacent atoms in a solid metallic crystal.

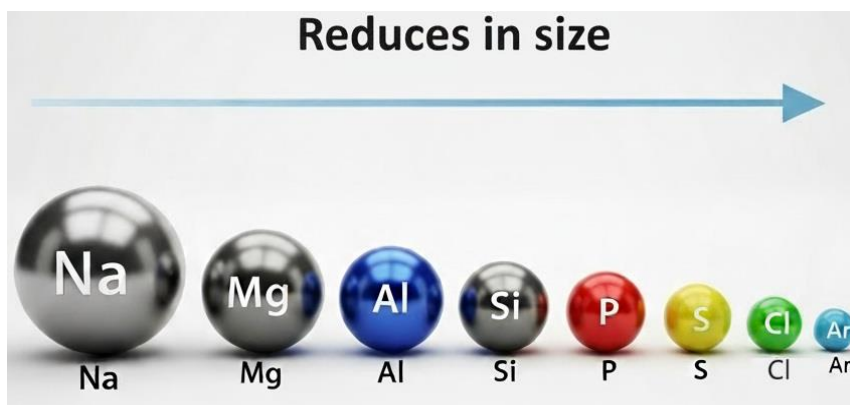


Atomic size gradation in the periodic table

Atomic size follows specific trends in the periodic table, which help us understand the properties of elements:

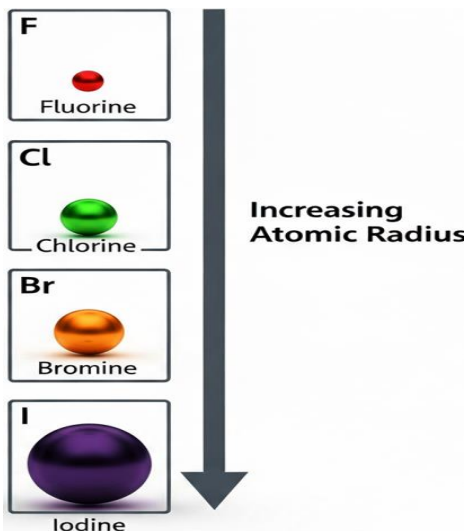
- Across the period (from left to right): Atomic size decreases.

Reason: As the atomic number increases, the number of protons in the nucleus increases. This results in an increase in the effective nuclear charge, which attracts the outer electrons more strongly, pulling them closer to the nucleus and shrinking the size of the atom.



- Across the group (from top to bottom): atomic size increases.

Reason: As you move down a group, new energy levels (orbitals) are added for electrons. These levels move the outer electrons away from the nucleus, increasing the size of the atom.



Ion size

The size of an ion, known as its ionic radius, is determined by the size of the neutral atom from which it originated, as well as the type of charge (positive or negative) and the number of electrons lost or gained. Ionic radius is usually measured in small units such as picometers (pm).

Comparing the size of an ion to a neutral atom:

1. Positive Ion (Cation):

- **How is it formed?** It is formed by a neutral atom losing one or more electrons from its outer energy shell.
- **Size:** It is smaller than the neutral atom from which it was formed.
- **Reason:**
 - **Decrease in energy levels:** Often, the loss of electrons results in the loss of an entire energy level, which significantly reduces the size (as in the case of Na to Na⁺).
 - **Increase in effective nuclear attraction:** The number of protons (the positive charge of the nucleus) becomes greater than the number of remaining electrons, which increases the nucleus's attractive force on the remaining electrons. The electrons are more attracted to the nucleus and the size shrinks.
 - **Effect of charge:** The greater the positive charge (i.e., the more electrons lost, such as the transition from Fe²⁺ to

Fe^{3+}), the greater the size shrinks (the ion becomes smaller).

2. Negative Ions (Anions):

- **How are they formed?** They are formed by a neutral atom gaining one or more electrons in its outer energy shell.
- **Size:** They are larger than the neutral atom from which they were formed.
- **Reason:**
 - **Increased repulsion between electrons:** Adding electrons to the outer shell increases the electrostatic repulsive forces between electrons, causing them to move away from each other and spread out over a larger space.
 - **Decrease in effective nuclear attraction:** The positive charge of the nucleus remains constant, while the number of electrons increases. This reduces the amount of attraction each electron feels individually, allowing the electrons to move away from the nucleus.
 - **Effect of charge:** The greater the negative charge (i.e., the more electrons gained, such as moving from O^{2-} to O^-), the greater the volume expansion (the ion becomes larger).

Factors affecting ion size:

1. **Number of energy levels:** The more principal energy levels occupied by electrons in an ion, the larger its size (for example, positive ions in a group increase in size from top to bottom).
2. **Effective nuclear charge:** The net positive charge felt by the outer electrons after taking into account the shielding effect of the inner electrons.
 - **For cations (positive ions):** An increase in the effective nuclear charge (due to the loss of electrons) reduces the volume.
 - **For anions (negative ions):** A decrease in the effective nuclear charge felt by the outer electrons (due to the gain of electrons) increases the volume.
3. **Repulsion between electrons:** Increasing the number of electrons in the last shell (as in the formation of an anion) increases the repulsion, causing the shell to expand and the volume to increase.

Group 1A	Group 2A	Group 3A	Group 6A	Group 7A
Li^+ Li  0.68 1.34	Be^{2+} Be  0.31 0.90	B^{3+} B  0.23 0.82	O O^{2-}  0.73 1.40	F F^-  0.71 1.33
Na^+ Na  0.97 1.54	Mg^{2+} Mg  0.66 1.30	Al^{3+} Al  0.51 1.18	S S^{2-}  1.02 1.94	Cl Cl^-  0.99 1.81
K^+ K  1.33 1.96	Ca^{2+} Ca  0.90 1.74	Ga^{3+} Ga  0.62 1.26	Se Se^{2-}  1.16 1.98	Br Br^-  1.14 1.96
Rb^+ Rb  1.48 2.11	Sr^{2+} Sr  1.13 1.92	In^{3+} In  0.62 1.44	Te Te^{2-}  1.35 2.94	I I^-  1.35 2.20

❖ Electronegativity:

Electronegativity is a chemical property that describes how strongly an atom attracts shared electrons (bond electrons) to itself in a covalent chemical bond. In other words, it is a measure of an atom's "attract" for electrons when bonded to another atom.

Electronegativity is usually measured on the Pauling scale, with values ranging from about 0.7 (for the least electronegative elements such as cesium and francium) to 4.0 (for the most electronegative element, fluorine).

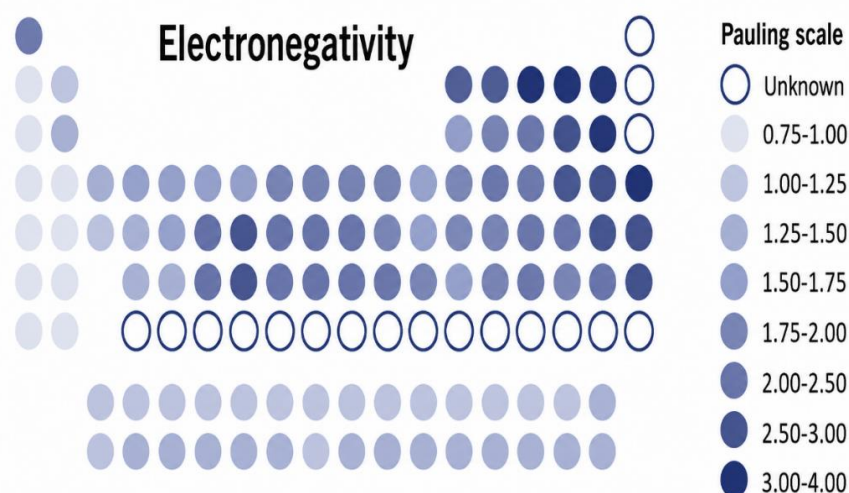
Factors affecting electronegativity:

- 1) **Effective Nuclear Charge:** The greater the positive charge in an atom's nucleus (number of protons), the greater its attraction to electrons, and thus its electronegativity.
- 2) **Atomic Radius:** The closer the valence electrons are to the nucleus (i.e., the smaller the atom's size), the greater the force of attraction generated by the nucleus, and the greater the electronegativity.
- 3) **Shielding Effect:** The inner electrons shield the positive nuclear charge from the valence electrons. The greater the number of inner

electrons, the less influence the nucleus has on the outer electrons, and thus the lower the electronegativity.

Trends in the periodic table:

- **From left to right across a period:** Electronegativity gradually increases. This is because the number of protons (nuclear charge) increases while maintaining the same number of electron shells. This increases the nucleus's attractive force on the electrons in the outer shell, making the atom more able to attract bonding electrons.
- **From bottom to top within a group:** Electronegativity increases. This is because the number of electron shells decreases, which reduces the distance between the nucleus and the valence electrons. This closeness increases the attractive force between them and makes the atom more electronegative.



❖ Ionization energy:

Ionization energy (abbreviated IE) is the amount of energy required to remove one electron from an isolated atom in the gaseous state. In other words, it is the amount of energy required to convert a neutral atom into a positively charged ion.

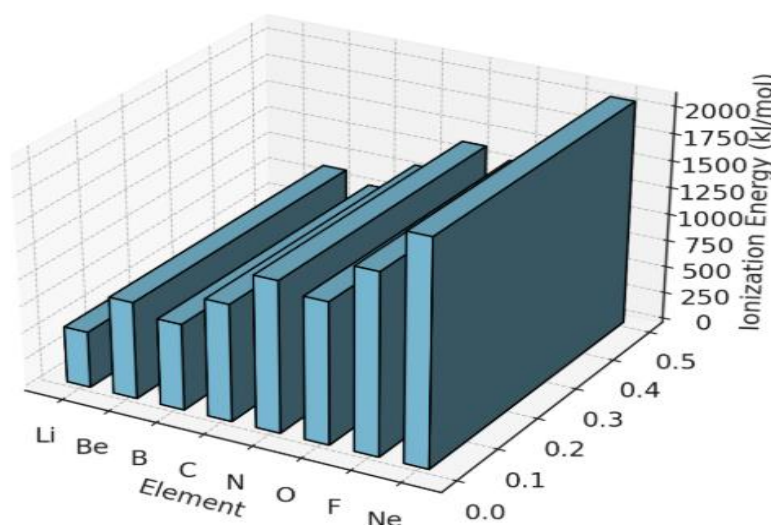
The value of ionization energy is always given as a positive number — the higher it is, the more tightly the electron is bound to the atom.

Ionization energy is affected by several main factors:

- 1) **Nuclear charge (number of protons):** The more protons in the nucleus, the greater the positive charge, and therefore the greater the attractive force on the electrons. This requires more energy to remove an electron, increasing the ionization energy.
- 2) **Electron shielding:** The inner electrons in an atom shield the nuclear charge from the outer electrons. This shielding reduces the effective attractive force felt by the outer electrons, making them easier to remove and lowering the ionization energy.
- 3) **Atomic radius:** The larger the atomic size (radius), the further the outer electrons are from the nucleus. This distance reduces the attractive force, making it easier to extract electrons and lowering the ionization energy.
- 4) **Stability of sub-orbitals:** Atoms with fully or half-filled sub-orbitals (such as s^2 , p^3 , or p^6) are more stable. Breaking this stability requires additional energy, leading to unexpectedly higher ionization energy values.

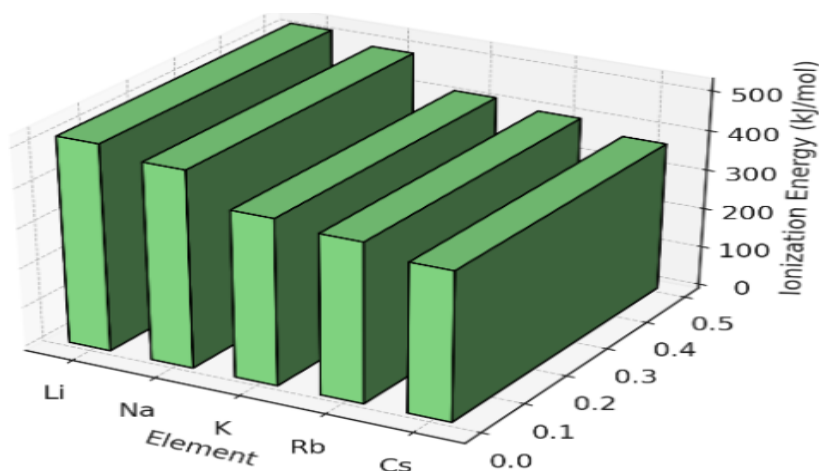
Trends in the Periodic Table:

- **Across a period (left to right):** Ionization energy generally increases.
Reason: The nuclear charge (number of protons) increases while the number of energy levels remains constant. This leads to an increased force of attraction on the outer electrons and a reduced atomic size, making it more difficult to remove an electron.



- **Across the group (from top to bottom):** Ionization energy generally decreases.

Reason: Increasing the number of principal energy levels increases atomic size and electron shielding. This reduces the nucleus's attractive force on the outer electrons, making them easier to separate.



General gradation on the periodic table

	1A	2A		3A	4A	5A	6A	7A	8A
1	H 1311								He 2377
2	Li 520	Be 899		B 800	C 1086	N 1402	O 1314	F 1681	Ne 2088
3	Na 495	Mg 735		Al 580	Si 780	P 1060	S 1005	Cl 1255	Ar 1527
4	K 419	Ca 590		Ga 579	Ge 761	As 947	Se 941	Br 1143	Kr 1356
5	Rb 409	Sr 549		In 558	Sn 708	Sb 834	Te 869	I 1009	Xe 1176
6	Cs 382	Ba 503		Tl 589	Pb 715	Bi 703	Po 813	At (926)	Rn 1042

The Effect of Ionization Energy on Chemical Properties:

Ionization energy determines an element's propensity to form ions, and thus its type and behavior:

- **Metals:** They have low ionization energy (especially Groups 1 and 2), and therefore tend to easily lose electrons and form Cation (Positive ion), making them chemically reactive.
- **Nonmetals:** They have high ionization energy, and therefore it is difficult for them to lose electrons. Instead, they tend to gain or share electrons to form negative ions.

- **Noble gases:** They have the highest ionization energy due to the stability of their complete outer shell, making them nearly inert.

Identifying an Element's Group: By observing large jumps in successive ionization energy values, the number of valence electrons an element and its group can be determined.

The effect of ionization energy on physical properties (indirectly):

While ionization energy is an atomic property that does not directly affect properties such as melting or boiling points, which depend on the bonding forces between atoms or molecules, it does affect them indirectly by determining the type of bond.

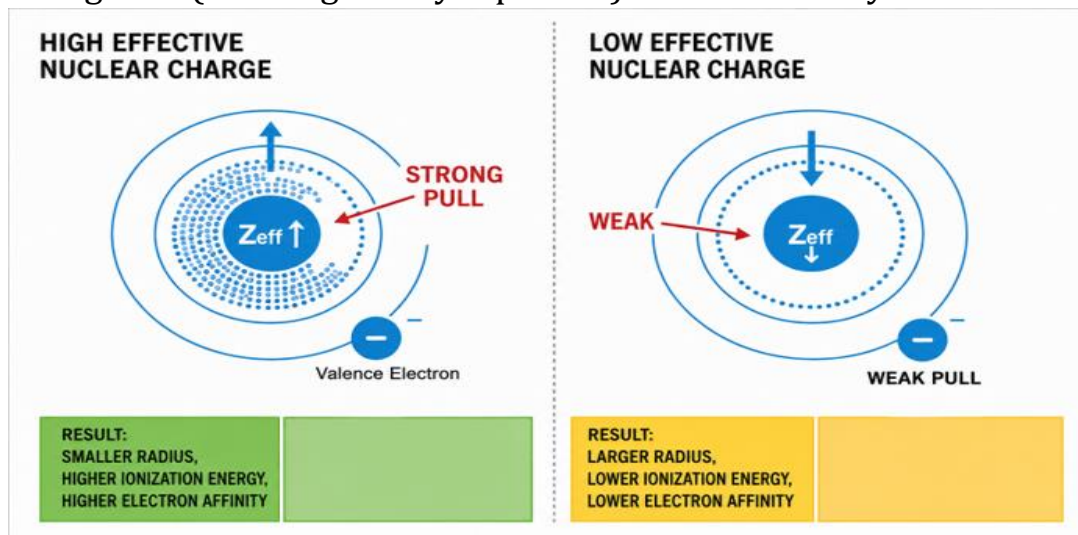
Indirect relationship	The explanation
Electrical and thermal conductivity	Elements with low ionization energy (metals) have weakly bound valence electrons, which become free-floating electrons in the metal lattice. These electrons are responsible for the distinctive physical properties of metals, such as high electrical and thermal conductivity.
Melting and boiling points	The low ionization energy of metals contributes to the formation of strong metallic bonds (attraction between cations and free electrons), resulting in relatively high melting and boiling points for many metals compared to elements that form covalent bonds (nonmetals).
Physical condition	Elements with high ionization energy (such as the nonmetals on the far right) tend to form simple molecules (such as O ₂ , N ₂) with weak covalent bonds, resulting in low melting and boiling points and their existence as gases or liquids at room temperature.

❖ Electron affinity:

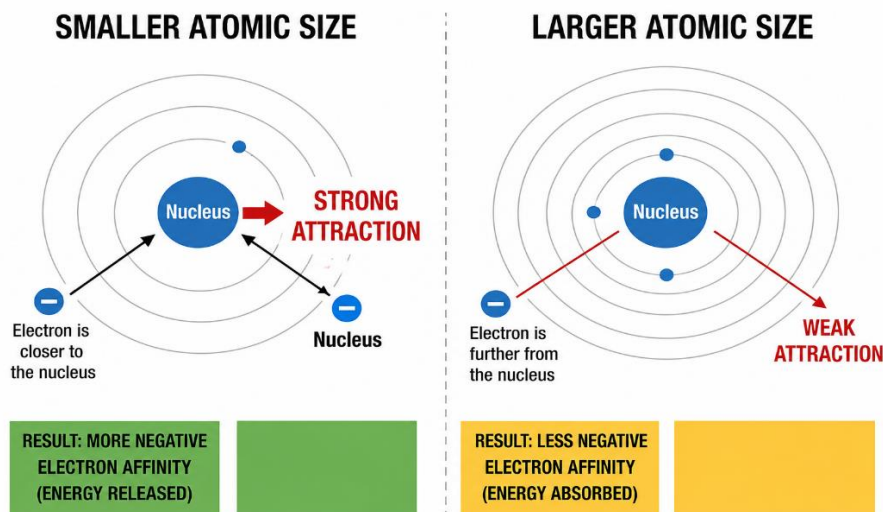
Electron affinity is the amount of energy released or absorbed when an atom in the gaseous state gains an electron. In general, the more negative an element's electron affinity, the greater its willingness to gain an electron and form a negative ion (anion), indicating that the process is exergonic (i.e., releases energy).

Factors affecting electron affinity:

- **Effective nuclear charge:** The higher the effective nuclear charge, the more attractive the nucleus is to the new electron, resulting in a more negative (or energetically repulsive) electron affinity value.



- **Atomic Size:** The smaller the atomic size, the closer the new electron is to the nucleus, which increases the force of attraction and makes the electron affinity more negative.



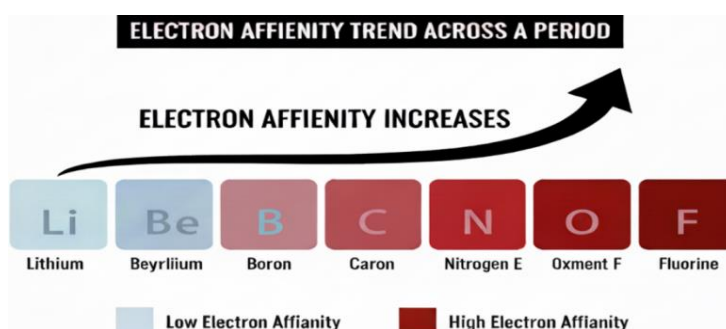
Electron Configuration:

- **Tendency to be full:** Atoms that only need one electron to gain a full outer shell (such as the halogens) have a highly negative electron affinity because they become highly stable after gaining that electron.
- **Tendency to be half-full:** Atoms with a half-full subshell (such as nitrogen) are relatively stable, making adding an electron difficult and requiring energy absorption. Therefore, the electron affinity value is positive or less negative.

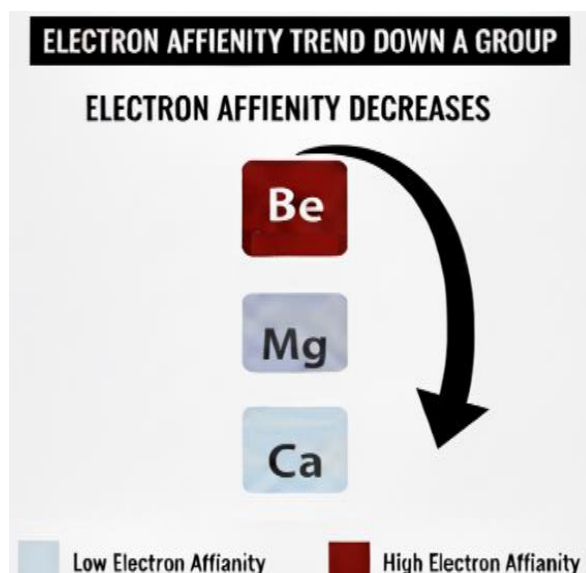
- **Filled shells:** Noble gases have fully filled outer shells, making them very stable. Therefore, their electron affinity value is positive (energy-absorbing), because adding an electron requires a lot of energy.

Periodic trends in electron affinity:

- Across a period (from left to right): Electron affinity tends to become more negative (i.e., more energetically repulsive). This is due to an increase in effective nuclear charge and a decrease in atomic size, which increases the nucleus's attraction to the outer electrons.



- **In a group (from top to bottom):** Electron affinity tends to become less negative (less energy-repellent). The main reason for this is the increase in atomic size and the increase in energy levels, which move the new electron away from the nucleus and reduce the force of attraction.



Effect of electron affinity on physical properties (indirect):

Electron affinity does not directly determine properties such as melting points or density, but it does affect the type of bond, which determines those properties.

Indirect relationship	The explanation
Electrical and thermal conductivity	Nonmetals (high electron affinity) tend to form negative ions or covalent bonds. These materials lack free electrons and are therefore insulators or poor conductors of electricity and heat.
Melting and boiling points	Elements with high electron affinity (such as nonmetals) often form simple molecules (weak covalent bonds), resulting in low melting and boiling points and their presence as gases or liquids under normal conditions (e.g., Cl_2, O_2).
Type of bonding and formation of compounds	The high electronegativity of an atom combined with the low ionization energy of another atom (a metal) predicts the formation of strong ionic bonds. Ionic compounds are characterized by high melting and boiling points and solidity (such as NaCl).

Exercises

(2-1)

Multiple choice questions

- 1) Which of the following elements has the greatest first ionization energy?
a) Na b) K c) Mg d) Ca e) Sr
- 2) Which value is the largest?
a) First ionization energy of potassium.
b) First ionization energy of calcium.
c) First ionization energy of scandium.
d) Second ionization energy of potassium.
e) Second ionization energy of calcium.
- 3) Which of the following chemical particles has the largest radius?
a) Mg^{2+} b) Na^+ c) Ne d) F^- e) O^{2-}
- 4) When considering the following elements: Na, Mg, Ca, F, Cl, K
Which of the following statements is true?
a) The atomic radius of sodium is larger than that of chlorine.
b) The atomic radius of fluorine is larger than that of potassium.
c) Magnesium has the smallest atomic radius and chlorine has the largest.
d) Calcium has the largest atomic radius and fluorine has the smallest.
e) Fluorine has the smallest atomic radius and chlorine has the largest.
- 5) Choose the pair that represents the orbital energies of oxygen (${}_8O$) for the 1s and then the 2s orbitals, from right to left?
a) $-3267 \text{ kJ mol}^{-1}$, $-54264 \text{ kJ mol}^{-1}$
b) $-6335 \text{ kJ mol}^{-1}$, $-3285 \text{ kJ mol}^{-1}$
c) $24098 \text{ kJ mol}^{-1}$, $77568 \text{ kJ mol}^{-1}$
d) 8744 kJ mol^{-1} , 2543 kJ mol^{-1}
e) 9744 kJ mol^{-1} , 3543 kJ mol^{-1}

6) Lithium, sodium, potassium, and rubidium are in Group 1 of the periodic table. The melting and boiling points of each of these elements are given in the following table:

	melting point	boiling point
Li	180°C	1330°C
Na		
K	64°C	774°C
Rb	39°C	688°C

According to values seen in the attached table, which of the following values represent the boiling and melting points of sodium?

- 59° and 910°C
- 113° and 753°C
- 89° and 890°C
- 134° and 1498°C
- 147° and 2498°

(2-2)

Choose the appropriate letter for each of the following:

- Magnesium
- Oxygen
- both of them
- neither of them

- Its atomic radius is larger than that of calcium.
- Its atomic radius is larger than that of beryllium.
- Its atomic radius is larger than its ionic radius.

(2-3)

Fill in the blanks based on the following:

A	B	C	D	E
Metals	Nonmetals	Metalloids	Noble elements	Neither of them

- It always exists in the gaseous state under normal conditions.....
- It exists as a solid or liquid under normal conditions.....
- It may exist in the gaseous state under normal conditions.....
- It has a shine and lustre.....
- Their melting points are low.....
- It has the properties of metals and non-metals.....

- 7) Including oxygen and iodine.....
- 8) Including iron and gold.....
- 9) Transition elements include.....
- 10) Does not react under normal conditions.....
- 11) It does not contain electrons.....
- 12) Sometimes it does not contain neutrons.....

(2-4)

Based on the following figure of the periodic table containing the following hypothetical elements, answer the following questions?

[illegible]

- 1) Identify an element that is in the third period and group 13?
- 2) What is the transition element found in the fourth period?
- 3) How many protons does element Q have?
- 4) Which of the elements in the previous table is :an alkali metal.....
An alkaline earth metal.....A halogen..... An inert
gas.....
- 5) Write the electronic configuration of element B
- 6) The elements that may exist in the gaseous state at room temperature
are

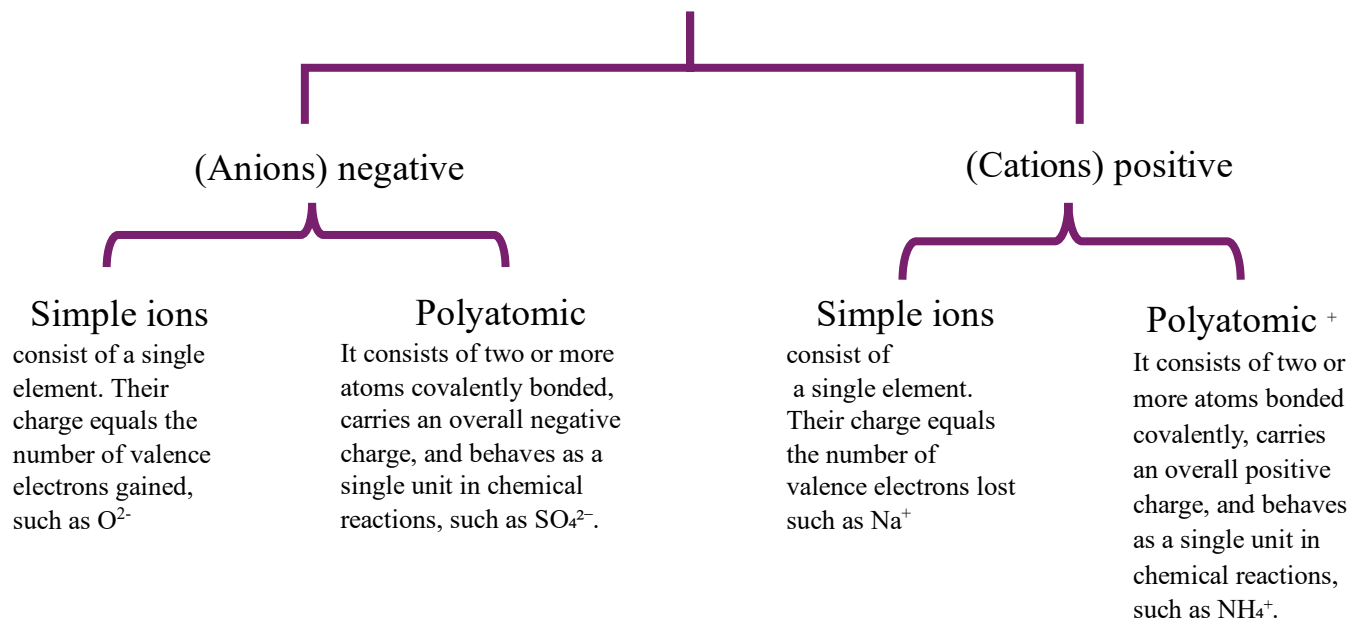
Exercise answers	
(2-1)	
c	(1
d	(2
e	(3
a	(4
a	(5
c	(6
(2-2)	
d	(1
a	(2
b	(3
(2-3)	
D	(1
A	(2
B, D	(3
A	(4
B, D	(5
C	(6
B	(7
A	(8
A	(9
D	(10
E	(11
B	(12
(2-4)	
D	(1
C	(2
17	(3
A, B, Q, Z	(4
$_{18}[\text{Ar}] 4s^2$	(5
G, Q, Z	(6

03

Ionic Bonds and Compounds

Ions

Ions are atoms of elements but that carrying charges.



1 1A											13 3A	14 4A	15 5A	16 6A	17 7A	18 8A	
2 2A																	
Li ⁺													C ⁴⁻	N ³⁻	O ²⁻	F ⁻	
Na ⁺	Mg ²⁺	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	Al ³⁺		P ³⁻	S ²⁻	Cl ⁻	
K ⁺	Ca ²⁺				Cr ²⁺ Cr ³⁺	Mn ²⁺ Mn ³⁺	Fe ²⁺ Fe ³⁺	Co ²⁺ Co ³⁺	Ni ²⁺ Ni ³⁺	Cu ⁺ Cu ²⁺	Zn ²⁺				Se ²⁻	Br ⁻	
Rb ⁺	Sr ²⁺									Ag ⁺	Cd ²⁺				Te ²⁻	I ⁻	
Cs ⁺	Ba ²⁺									Au ⁺ Au ³⁺	Hg ₂ ²⁺ Hg ²⁺				Pb ²⁺ Pb ⁴⁺		

*Remember



CATS ARE GOOD
(POSITIVE)



ONIONS
ARE
BAD
(NEGATIVE)

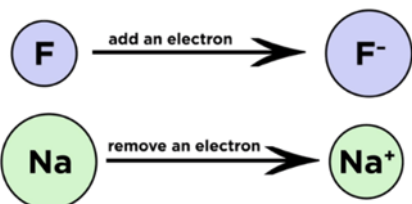
Anions are negatively charged ions

Cations are positively charged ions.

Major Cations and Anions

	Strontium	Barium	Calcium	Magnesium	Silver	Potassium	Sodium	Lithium	Hydrogen	Zinc	Ion
Symbol	Sr^{2+}	Ba^{2+}	Ca^{2+}	Mg^{2+}	Ag^+	K^+	Na^+	Li^+	H^+	Zn^{2+}	Positive Ions (Cations)
	Ammonium	Aluminum	Chromium III	Chromium II	Iron III	Iron II	Copper II	Copper I	Zinc		
Symbol	NH_4^+	Al^{3+}	Cr^{3+}	Cr^{2+}	Fe^{3+}	Fe^{2+}	Cu^{2+}	Cu^+	Zn^{2+}		Negative Ions (Anions)
	Cyanide	Hydroxide	Nitrid	Sulfide	Oxide	Iodide	Bromide	Chloride	Fluoride	Ion	
Symbol	CN^-	OH^-	N^{3-}	S^{2-}	O^{2-}	I^-	Br^-	Cl^-	F^-	Symbol	Negative Ions (Anions)
	Silicate	Perchlorate	Chlorate	Chlorite	Hypochlorite	Permanganate	Bicarbonate	Nitrate	Nitrite	Ion	
Symbol	SiO_3^{2-}	ClO_4^-	ClO_3^-	ClO_2^-	ClO^-	MnO_4^-	HCO_3^-	NO_3^-	NO_2^-	Symbol	

Remember: Size of Ions



Note: The size of ions varies. An anion is larger than its neutral atom, while a cation is smaller. This change is due to electron loss or gain.

Exercises

(3-1)

Write the names of the following compounds, indicating the number of oxygen atoms in each compound.

$\text{Al}(\text{NO}_3)_3$	$\text{Ca}(\text{OH})_2$	Na_2CO_3	Compound
			Name
			Number of oxygens

(3-2)

A hypothetical element with the symbol (X) is in the fourth period and group 16, and its electron configuration ends in the $4P^4$ sublevel.

a) Write its electron configuration.

.....

b) How many protons does it have?

.....

c) When this element reaches stability, how many electrons surround its nucleus? Write its ion symbol and its electron configuration.

.....

d) Which of the noble gases does its electron configuration resemble?

.....

(3-3)

A transition element with the hypothetical symbol (Y) is in the fourth period and group 8.

a) Write its electron configuration and identify how many protons does it has.

.....

b) What are the possible positive ions this element can form to reach stability? Write their electron configuration.

.....

(3-4)

A hypothetical representative element with an ion symbol of ${}_{13}X^{3+}$

a) Write the electron configuration of this ion.

.....

b) How many electrons does it have?

.....

(3-5)

Fill in the blanks in the following table:

Element symbol	Electron configuration	period	Group	Ion symbol	Element type	Electrical conductivity
${}^3\text{Li}$						
${}^7\text{N}$						
${}^{19}\text{K}$						
${}^{17}\text{Cl}$						
${}^{20}\text{Ca}$						
${}^{16}\text{S}$						

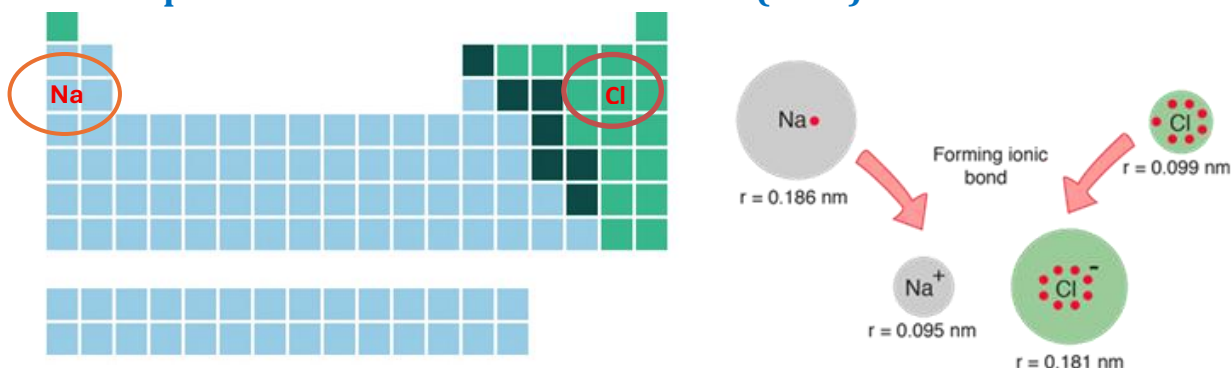
Chemical bond

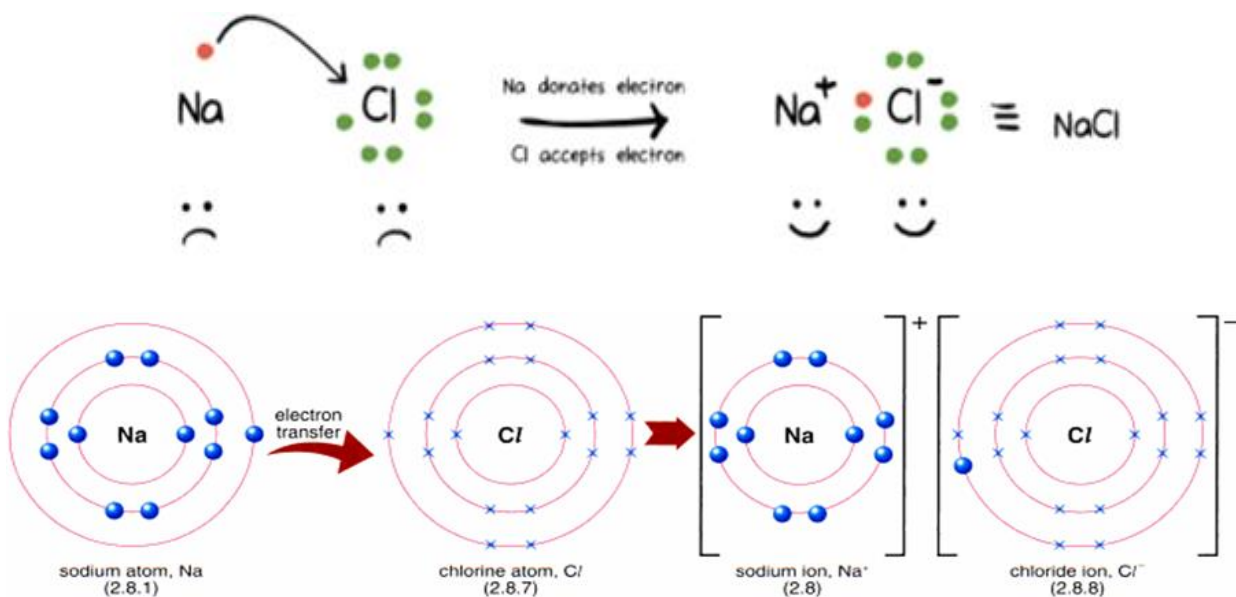
A chemical bond is the attractive force that arises between two or more atoms through electron loss, gain, or sharing.

Ionic Bonds

Ionic bonds are formed between a metal and a non-metal due to the electrostatic attraction of oppositely charged ions, resulting in ionic compounds.

Example: how is sodium chloride formed (NaCl)





A metal atom loses its valence electrons until it stabilizes and forms a positive ion.

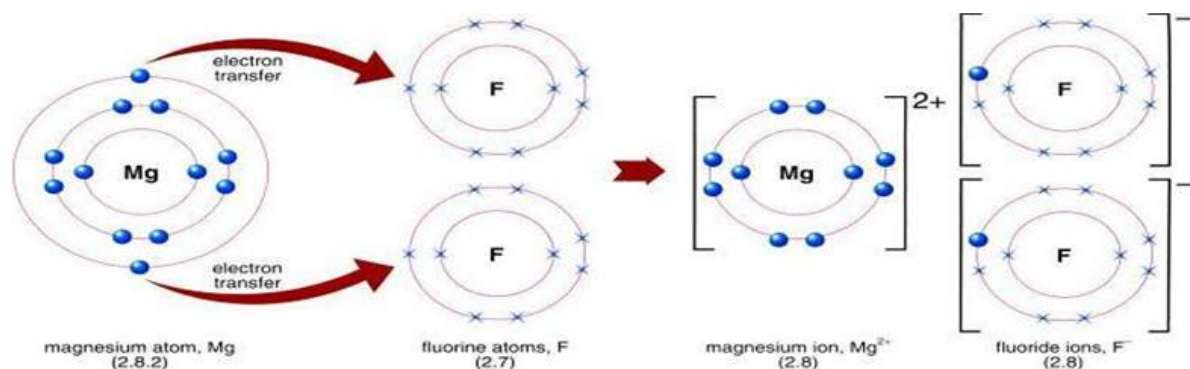
A nonmetal atom gains Electrons until it stabilizes and forms a negative ion

Negative and positive ions attract each other, The number of atoms is such that the overall charge of the compound is zero

The following table summarizes several methods for forming ionic compounds, including sodium chloride.

Chemical Equation
$\text{Na} + \text{Cl} \rightarrow \text{Na}^+ + \text{Cl}^- + \text{energy}$
Electron Configuration
Electron transfer $[\text{Ne}] 3s^1 + [\text{Ne}] 3s^2 3p^5 \rightarrow [\text{Ne}] + [\text{Ar}] + \text{energy}$ Na Cl Na⁺ Cl⁻
Electron Configuration by Drawing Orbital Boxes
Electron transfer Na Cl Na⁺ Cl⁻ + energy
Electron Dot Representation (Lewis Representation)
Electron transfer $\text{Na} \cdot + \cdot \ddot{\text{Cl}} : \rightarrow [\text{Na}]^+ + [: \ddot{\text{Cl}} :]^- + \text{energy}$
Atomic Models
Sodium atom Chlorine atom Sodium ion Chloride ion

How Magnesium Fluoride is Formed (MgF₂)



Properties of Ionic Compounds

1. Solid at room temperature.
2. High melting and boiling points.
3. Mostly soluble in water and dissociate into ions.
4. They conduct electricity in the molten state or in aqueous solution, but not in the solid state.
5. Ionic crystals are hard, brittle, and strong.

Lattice Energy of Ionic Compounds

Definition: The energy required to separate one mole of an ionic solid into gaseous ions. It reflects the strength of electrostatic forces between cations and anions.

Factors:

1. Ion charge: Higher charges → stronger attraction → higher lattice energy.
Example: MgO > NaCl because Mg²⁺ and O²⁻ carry higher charges.
2. Ion size: Smaller ionic radius → stronger attraction.
Example: NaF > NaI because fluoride is smaller than iodide.

Importance:

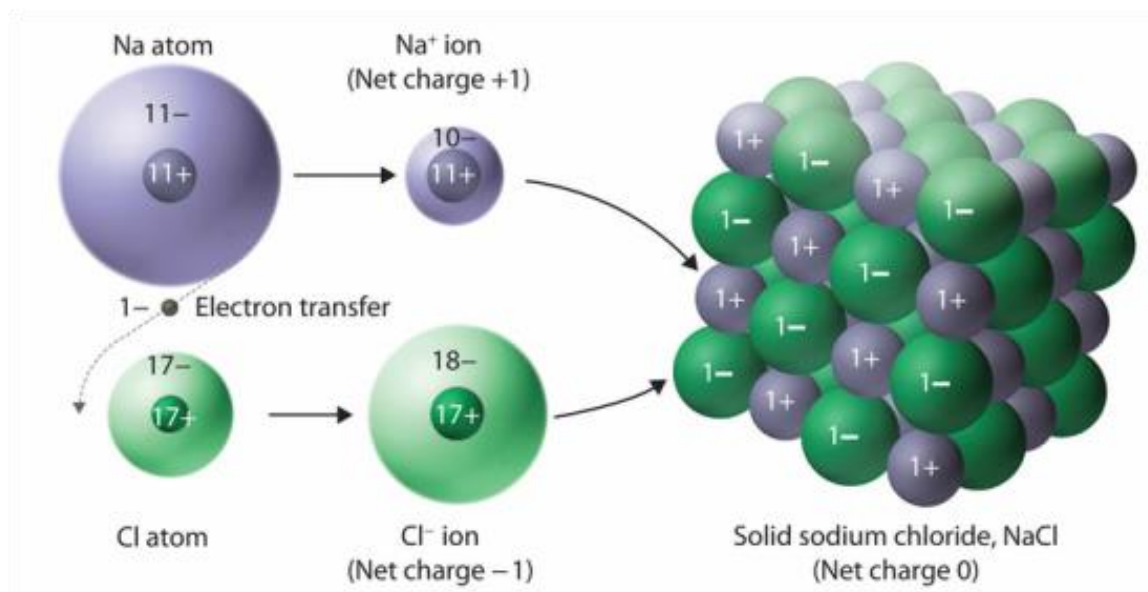
- Explains hardness and high melting points.
- Explains solubility: if lattice energy is too high, solubility decreases.
- Directly related to density and thermal resistance.

Illustrative Examples

The lattice energy of MgO is higher than that of NaF, making MgO more solid and having a higher melting point.

Compounds such as SrCl_2 or AgCl vary in energy depending on the size and charge of their constituent ions.

Summary: The crystal lattice energy is a measure of the strength of the bond between ions. It increases with increasing charge and smaller ionic radius. It is key to understanding the stability and properties of ionic compounds.



Steps for Writing the Chemical Formula of Ionic Compounds

Example

Write the chemical formula of the ionic compound formed from chloride ion and magnesium ion.

Solution

Must follow these steps:

First: Write the symbol of the positive ion on the left and the negative ion on the right.



Second: Write the charge of each ion below it (without the positive or negative sign)



Third: We exchange between charge numbers



Charges are in their simplest ratio, so abbreviations may be necessary.
For example: The formula for silicon oxide is SiO_2 , not Si_2O_4 .

If the number of charges is equal, it is not written.
For example: The formula for calcium oxide is CaO , not Ca_2O_2 .

If the number of charges is +1 or -1, it is not written.
For example: The formula for magnesium chloride is MgCl_2 , not Mg_1Cl_2 .

Polyatomic ions are enclosed in parentheses.
For example: The formula for barium hydroxide is $\text{Ba}(\text{OH})_2$, not BaOH_2 .

Exercises

(3-6) Write the chemical formula for the following compounds.

IronII sulfide	Sodium silicate	Sodium sulfate	IronIII sulfate	Calcium hydroxide
Copper II fluoride	Aluminum carbonate	Potassium bromide	Magnesium bicarbonate	Zinc hydroxide

(3-7) Based on the ions in the following table, write the chemical formulas of the compounds, then write the name of each compound below Formula.

Ions	OH ⁻	O ²⁻	PO ₄ ³⁻	NO ₃ ⁻	SO ₄ ²⁻	SO ₃ ²⁻
Ca ²⁺						
K ⁺						
Al ³⁺						
Cu ²⁺						
Fe ²⁺						
Fe ³⁺						
NH ₄ ⁺						

(3-8) Based on the following elements, fill in the blanks in the table and then answer the subsequent questions.

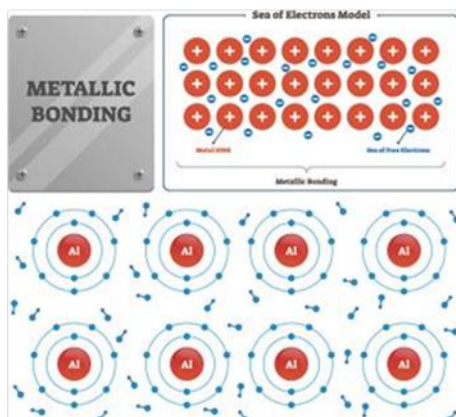
19K	17Cl	8O	7N	12Mg	Element
					Symbol of Ion
					Number of Bonds

- The type of bond between (O, N) is
- The type of bond between (Mg, O) is
- The type of bond between (Mg, Cl) is.
- The formula of the compound formed by the bonding of (Mg, N) is.....
- The formula of an oxygen molecule is.
- Which has the higher boiling point, KCl or Cl₂O

Metallic Bond

A force of attraction between positive metal ions and the surrounding sea of electrons.

The sea of electrons is responsible for many of the distinctive properties of metals.



Factors Affecting the Strength of a Metallic Bond:

- Number of free electrons.
- Amount of positive charge in the nucleus.
- Size of the metal atom.

The Metals

"They are chemical elements that lose electrons to form positive ions. They are mainly located in the left and center of the periodic table, including Groups 1 and 2, most transition metals, and some Group 13 elements."

Properties of Metals:

1. They are shiny and have a luster.
 2. They are ductile and malleable.
 3. They have a high melting point.
 4. They are solid and have a high density.
 5. They are good conductors of electricity and heat.
 6. They are chemically active, reacting with oxygen to form metal oxides.
- Their chemical activity also varies.

Exercise answers

(3-1)

$\text{Al}(\text{NO}_3)_3$	$\text{Ca}(\text{OH})_2$	Na_2CO_3	Compound
Aluminum nitrate	Calcium hydroxide	Sodium carbonate	Name
9	2	3	Number of oxygen atoms

(3-2)

(a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$

(b) 34 protons

(c) electrons 36

${}_{36}\text{X}^{2+}: 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

(d) Kr

(3-3)

(a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ 26 protons

(b) ${}_{26}\text{Fe}^{3+}: 1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$

${}_{26}\text{Fe}^{2+}: 1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$

(3-4)

(a) $1s^2 2s^2 2p^6$, (b) 10

(3-5)

Element symbol	Electron configuration	period	Group	Ion symbol	Element type	Electrical conductivity
${}_3\text{Li}$	$1s^2 2s^1$	2	1	Li^+	metal	yes
${}_7\text{N}$	$1s^2 2s^2 2p^3$	2	5	N^{3-}	nonmetal	no
${}_{19}\text{K}$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	4	1	K^+	metal	yes
${}_{17}\text{Cl}$	$1s^2 2s^2 2p^6 3s^2 3p^5$	3	17	Cl^-	Non-metal	no
${}_{20}\text{Ca}$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	4	2	Ca^{2+}	metal	yes
${}_{16}\text{S}$	$1s^2 2s^2 2p^6 3s^2 3p^4$	3	16	S^{2-}	Non-metal	no

(3-6)

Iron(II) sulfide	Sodium silicate	Sodium sulfate	Iron(III) sulfate	Calcium hydroxide
FeS	Na_2SiO_3	Na_2SO_4	$\text{Fe}_2(\text{SO}_4)_3$	$\text{Ca}(\text{OH})_2$
Copper(II) fluoride	Aluminum carbonate	Potassium bromide	Magnesium bicarbonate	Zinc hydroxide
CuF_2	$\text{Al}_2(\text{CO}_3)_3$	KBr	$\text{Mg}(\text{HCO}_3)_2$	$\text{Zn}(\text{OH})_2$

(3-7)

Anions	OH^-	O^{2-}	PO_4^{3-}	NO_3^-	SO_4^{2-}	SO_3^{2-}
Cations						
Ca^{2+}	Ca(OH)_2 Calcium hydroxide	CaO Calcium oxide	$\text{Ca}_3(\text{PO}_4)_2$ Calcium phosphate	$\text{Ca(NO}_3)_2$ Calcium nitrate	CaSO_4 Calcium sulfate	CaSO_3 Calcium sulfite
K^+	KOH Potassium hydroxide	K_2O Potassium oxide	K_3PO_4 Potassium phosphate	KNO_3 Potassium nitrate	K_2SO_4 Potassium sulfate	K_2SO_3 Potassium sulfite
Al^{3+}	Al(OH)_3 Aluminum hydroxide	Al_2O_3 Aluminum oxide	AlPO_4 Aluminum phosphate	$\text{Al(NO}_3)_3$ Aluminum nitrate	$\text{Al}_2(\text{SO}_4)_3$ Aluminum sulfate	$\text{Al}_2(\text{SO}_3)_3$ Aluminum sulfite
Cu^{2+}	Cu(OH)_2 Copper(II) hydroxide	CuO Copper(II) oxide	$\text{Cu}_3(\text{PO}_4)_2$ Copper(II) phosphate	$\text{Cu(NO}_3)_2$ Copper(II) nitrate	CuSO_4 Copper(II) sulfate	CuSO_3 Copper(II) sulfite
Fe^{2+}	Fe(OH)_2 Iron(II) hydroxide	FeO Iron(II) oxide	$\text{Fe}_3(\text{PO}_4)_2$ Iron(II) phosphate	$\text{Fe(NO}_3)_2$ Iron(II) nitrate	FeSO_4 Iron(II) sulfate	FeSO_3 Iron(II) sulfite
Fe^{3+}	Fe(OH)_3 Iron(III) hydroxide	Fe_2O_3 Iron(III) oxide	FePO_4 Iron(III) phosphate	$\text{Fe(NO}_3)_3$ Iron(III) nitrate	$\text{Fe}_2(\text{SO}_4)_3$ Iron(III) sulfate	$\text{Fe}_2(\text{SO}_3)_3$ Iron(III) sulfite
NH_4^+	NH_4OH Ammonium hydroxide	$(\text{NH}_4)_2\text{O}$ Ammonium oxide	$(\text{NH}_4)_3\text{PO}_4$ Ammonium phosphate	NH_4NO_3 Ammonium nitrate	$(\text{NH}_4)_2\text{SO}_4$ Ammonium sulfate	$(\text{NH}_4)_2\text{SO}_3$ Ammonium sulfite

(3-8)

^{19}K	^{17}Cl	^8O	^7N	^{12}Mg	Element
K^+	Cl^-	O^{2-}	N^{3-}	Mg^{2+}	Symbol of Ion
1	1	2	3	2	Number of Bonds

- a) covalent
- b) ionic
- c) ionic
- d) Mg_3N_2
- e) O_2
- f) KCl

04

Covalent Bonds and Compounds

The Covalent Bonding Model

Covalent bonds form when atoms share electrons, and they appear in everything from simple molecules like H_2 to complex biological macromolecules and polyatomic ions. Electron sharing is the one of the main ways atoms bond.

Hydrogen exists as H_2 because two atoms are more stable together than alone. As they approach, the system's potential energy decreases—the attraction between electrons and nuclei outweighs repulsion. At the optimal distance, called the bond length, a covalent bond forms (Figure 4.1), creating the stable H_2 molecule instead of separate atoms.

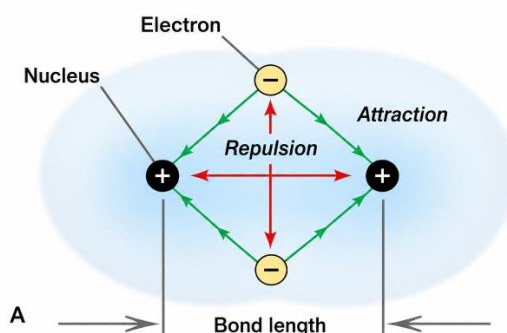


Figure 4.1 Covalent bond formation in H_2

The Formation of the Covalent Bond

Hydrogen gas exists as H_2 because two atoms are more stable bonded together. This stability can be seen on the potential energy curve (Figure 4.2).

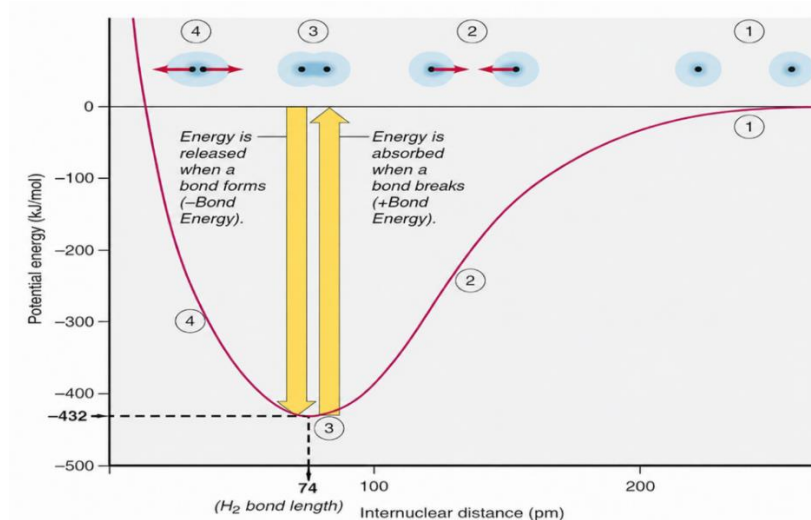


Figure 4.2 The potential energy curve of covalent bond formation in H_2

- **Point 1** – Atoms are far apart and act independently.
- **Point 2** – As atoms move closer, each nucleus attracts the other's electron, lowering potential energy. Attractions increase, but so do repulsions.
- **Point 3** – At the bottom of the curve (bond length), attraction balances repulsion. This is the most stable, lowest-energy state, where a covalent bond forms.
- **Point 4** – If atoms move too close, repulsions dominate, pushing them back toward the stable bond length.

Thus, a covalent bond results from the balance of attractive and repulsive forces, with greater electron density between the nuclei holding the atoms together (Figure 4.3).

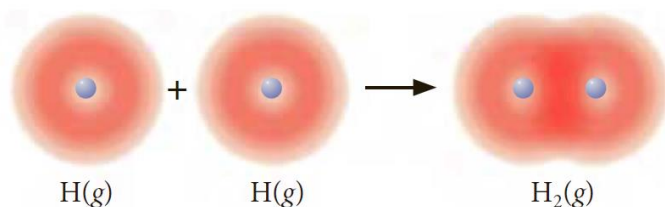
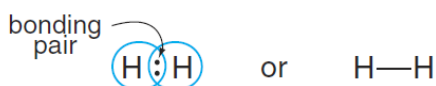


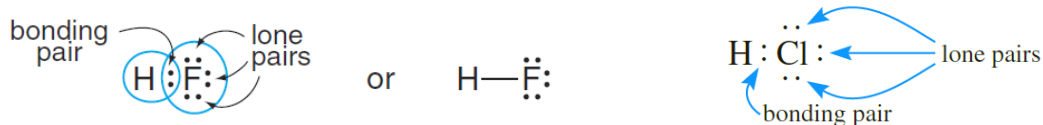
Figure 4.3 The electron probability distribution for the H_2 molecule

Bonding Pairs and Lone Pairs

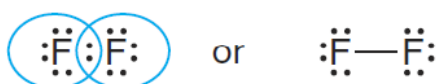
The **shared pair**, or **bonding pair**, is represented by a pair of dots or a line:



An outer-level electron pair that is not involved in bonding is called a **lone pair**, or **unshared pair**. The bonding pair in HF or HCl fills the outer level of the H atom and, together with three lone pairs, fills the outer level of the F atom as well.



In F_2 the bonding pair and three lone pairs fill the outer level of each F atom.



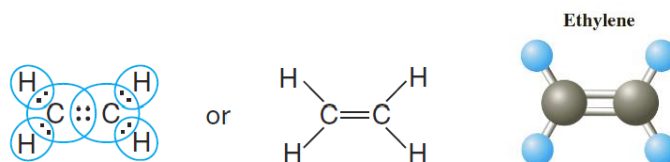
Properties of the Covalent Bond: Order, Energy, and Length

1. Bond Order

The number of electron pairs shared between two atoms.

- **Single bond:** 1 shared pair $\rightarrow$ bond order = 1 (e.g., H_2 , HF, F_2).

- **Double bond:** 2 shared pairs → bond order = 2 (e.g., C=C in ethylene, C₂H₄).



- **Triple bond:** 3 shared pairs → bond order = 3 (e.g., N≡N and C≡C in acetylene, C₂H₂).



2. Bond Energy

The energy required to break one mole of bonds in the gas phase.

- Higher bond order → stronger bond → higher bond energy.
- Triple bonds > double bonds > single bonds.

3. Bond Length

The average distance between the nuclei of bonded atoms.

- Higher bond order → shorter bond length.
C≡C (shortest), C=C (intermediate), C-C (longest)
- **Within a group:** Larger atoms → longer bonds → weaker bonds.
C-I > C-Br > C-Cl (length), but C-Cl > C-Br > C-I (energy)
- **Within a period:** Smaller atoms → shorter bonds → stronger bonds
C-N > C-O > C-F (length), but C-F > C-O > C-N (energy).

➤ Relationship Between the Three Properties:

- As bond order increases → bond length decreases and bond energy increases.
- Stronger, shorter bonds = greater stability, but often less reactive.

Bond	Bond Order	Average Bond Length (pm)	Average Bond Energy (kJ/mol)
C—O	1	143	358
C=O	2	123	745
C≡O	3	113	1070
C—C	1	154	347
C=C	2	134	614
C≡C	3	121	839
N—N	1	146	160
N=N	2	122	418
N≡N	3	110	945

Table 4.1 The Relation of Bond Order, Bond Length, and Bond Energy

Exercises:

(4-1)

Rank the bonds in each set in order of decreasing bond length and decreasing bond strength.



.....

.....

.....

.....

(4-2)

Rank the bonds in each set in order of increasing bond length and increasing bond strength.



.....

.....

.....

Naming Molecules

Binary molecular compounds are those that contain only two different elements. In naming such compounds, given their formulas, the rules below are followed.

1. Name the first element in the formula first, using its element name, unchanged.
2. Name the second element, using the root of its name and adding the suffix-ide.
3. Use the following prefixes, given the number of each type of atom present: 1, mono-; 2, di-; 3, tri-; 4, tetra-; 5, penta-; 6, hexa-; 7, hepta-. Mono- is not used as a prefix for the name of the first element.

Examples:

CO → carbon monoxide

CO₂ → carbon dioxide

NO → nitrogen monoxide

NO₂ → nitrogen dioxide

N₂O₅ → dinitrogen pentoxide

PCl₃ → phosphorus trichloride

SF₆ → sulfur hexafluoride

Cl₂O₇ → dichlorine heptoxide

Naming ionic compounds with polyatomic ions:

The name = name of the cation + name of the anion. Write the name of the polyatomic ion as it is without any change in the name. Do not use the suffix “-ide” as in the case of monatomic ions as shown in table 4.2.

Name	Formula	Name	Formula
Mercury(I)	Hg_2^{2+}	Permanganate	MnO_4^-
Ammonium	NH_4^+	Nitrite	NO_2^-
Cyanide	CN^-	Nitrate	NO_3^-
Carbonate	CO_3^{2-}	Hydroxide	OH^-
Hydrogen carbonate	HCO_3^-	Peroxide	O_2^{2-}
Acetate	$\text{C}_2\text{H}_3\text{O}_2^-$	Phosphate	PO_4^{3-}
Oxalate	$\text{C}_2\text{O}_4^{2-}$	Monohydrogen phosphate	HPO_4^{2-}
Hypochlorite	ClO^-	Dihydrogen phosphate	H_2PO_4^-
Chlorite	ClO_2^-	Sulfite	SO_3^{2-}
Chlorate	ClO_3^-	Sulfate	SO_4^{2-}
Perchlorate	ClO_4^-	Hydrogen sulfite	HSO_3^-
Chromate	CrO_4^{2-}	Hydrogen sulfate	HSO_4^-
Dichromate	$\text{Cr}_2\text{O}_7^{2-}$	Thiosulfate	$\text{S}_2\text{O}_3^{2-}$

Table 4.2 Common Polyatomic Ions

Example 1: Li_2CO_3 is lithium carbonate (not “dilithium carbonate”) because Li^+ is always +1 and CO_3^{2-} is always -2; the ratio is dictated by charge balance.

(4-3)

Write the proper formula and give the proper name for each ionic compound formed between the two listed ions.

(a) NH_4^+ and PO_4^{3-} (b) Al^{3+} and NO_2^-

.....

.....

.....

.....

Electronegativity (EN)

Electronegativity (EN) is the ability of atoms in a molecule to attract electrons to itself. On the periodic chart (Figure 4.4) , electronegativity increases as you go from left to right across a row. from the bottom to the top of a column.

1 H 2.20																		
3 Li 0.98	4 Be 1.57													5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98
11 Na 0.93	12 Mg 1.31													13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96		
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 2.10	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66		
55 Cs 0.79	56 Ba 0.89	57 La 1.10	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2		
87 Fr 0.7	88 Ra 0.9	89 Ac 1.1																
Metal Metalloid Nonmetal																		

Figure 4.4 The Pauling electronegativity

The Electronegativity Difference (ΔEN) and Bond Polarity

The difference between the electronegativity values of the bonded atoms is directly related to the polarity of the bond. It ranges from 0.0 non-polar covalent bond in a diatomic element, such as H_2 , O_2 , or Cl_2 , all the way up to 3.3, the difference between the most electronegative atom, F (4.0), and the most electropositive, Cs (0.7), in the ionic compound CsF.

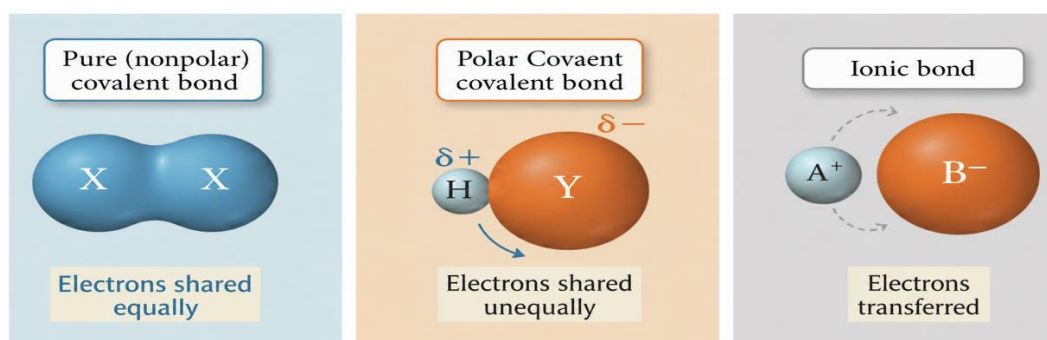
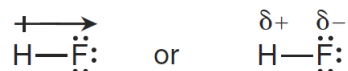


Figure 4.5 The Electronegativity difference (ΔEN) and types of bonds

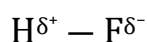
Nonpolar covalent bonds: When identical atoms bond (H_2 , F_2 , Cl_2), electron sharing is equal.

Polar covalent bond: Whenever atoms of different electronegativities form a bond, such as H (2.1) and F (4.0) in HF, the bonding pair is shared

unequally. This unequal configuration of electron density results in a polar covalent bond. It is depicted by a polar arrow ($\rightarrow$) pointing toward the partially negative atom or by the symbols δ^+ and δ^- . A greater ΔEN leads to larger partial charges and higher ionic character.



Example 2: HF



δ^+ and δ^- symbols show partial positive and negative charges.

Example 3: Use electronegativity values (Figure 4.4) to arrange the following bonds in order of increasing polarity.



(EN): P-H, 0.0; H-O, 1.4; C-Cl, 0.5:



Example 4: Consider ΔEN for the three Cl-containing gaseous molecules:

for LiCl (g) is $3.0 - 1.0 = 2.0$;

for HCl (g), it is $3.0 - 2.1 = 0.9$;

for Cl₂ (g), it is $3.0 - 3.0 = 0.0$.

Thus, the polarity of the bond,

LiCl: $\Delta EN = 2.0 \rightarrow$ highly ionic character.

HCl: $\Delta EN = 0.9 \rightarrow$ polar covalent.

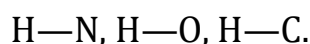
Cl₂: $\Delta EN = 0.0 \rightarrow$ nonpolar covalent.

(4-4)

(a) Use a polar arrow to indicate the polarity of each bond:



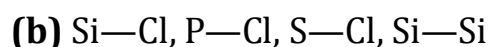
(b) Rank the following bonds in order of increasing polarity and decreasing percent ionic character.



.....
.....
.....
.....

(4-5)

Arrange each set of bonds in order of increasing polarity, and indicate bond polarity with δ^+ and δ^- symbols:



.....
.....
.....

The effect of electronegativity and bond type on the physical properties

For Period 3 chlorides (Figure 4.6), there is a gradation in bond type from ionic to polar covalent to nonpolar covalent.

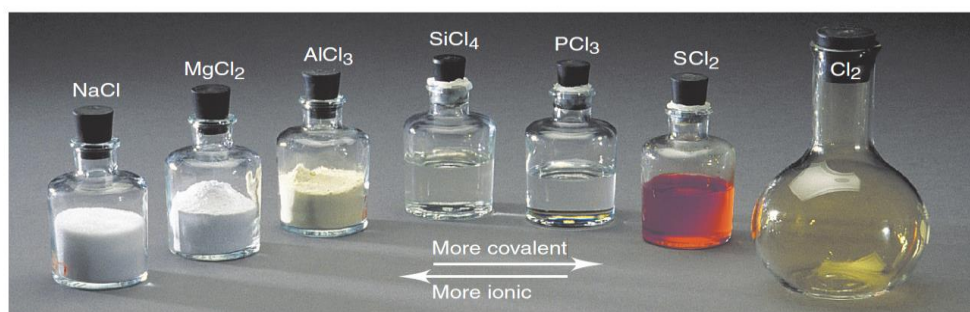


Figure 4.6 Properties of the Period 3 chlorides

As ΔEN decreases, melting point and electrical conductivity decrease because the bond type changes from ionic to polar covalent to nonpolar covalent.

The Shapes of Molecules

Depicting Molecules and Ions with Lewis Structures

The Lewis structure of a covalent compound or polyatomic ion shows how the valence electrons are arranged among the atoms in the molecule to show the connectivity of the atoms.

The valence electrons are then written as dots or (small cross marks) around the symbol of the atom. Represent the number of valence electrons as dots. Valence number is the same as the Periodic Table Group Number (Figure 4.7).

[illegible]

Figure 4.7 Lewis symbols of the main group elements

Applying the Octet Rule to Write Lewis Structures

Lewis summarized much of his theory of chemical bonding with the octet rule. According to the **octet rule**, atoms will lose, gain, or share electrons to achieve a noble gas electron configuration.

Steps:

- 1. Place atoms** – Central atom is usually the one needing more electrons (lower group number, lower electronegativity). H is never central.
- 2. Count valence electrons** – Add up valence e^- of atoms; adjust for ion charge (add e^- for negative charge, subtract for positive).
- 3. Form single bonds** – Connect each surrounding atom to the central atom with a bond (subtract $2 e^-$ per bond).
- 4. Distribute remaining e^-** – Complete octets (2 for H) with lone pairs. Put leftovers on central atom.

Bond rules:

- H $\rightarrow$ 1 bond (never central)
- C $\rightarrow$ 4 bonds
- N $\rightarrow$ 3 bonds
- O $\rightarrow$ 2 bond
- Halogens (F, Cl, Br, I) $\rightarrow$ 1 bond (F always terminal)

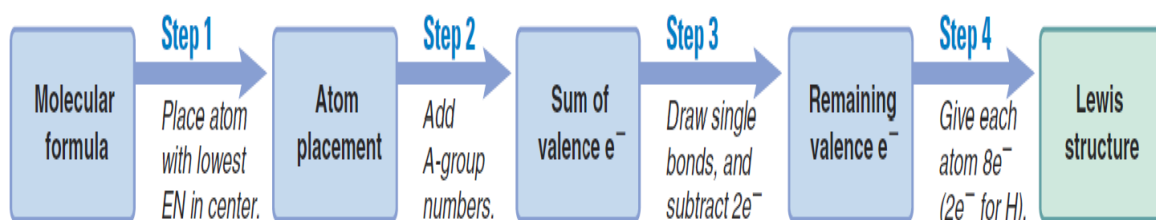
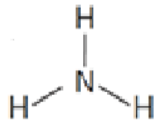
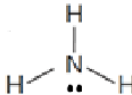
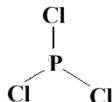
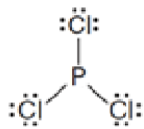
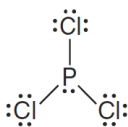


Figure 4.8 The steps in converting a molecular formula into a Lewis structure.

Example 5:	
Lewis structure for NF_3	
N in center, 3 Fs around.	
Total valence e^- N - has 5 valence electrons (5A) F - has 7 valence electrons (7A) NF_3 has $5 + 3(7) = 26 e^-$	
3 bonds use $6e^- \rightarrow 20 e^-$ left.	
Distribute: each F gets 3 lone pairs ($18 e^-$), N gets 1 lone pair ($2 e^-$). Final structure: N bonded to 3 Fs, each atom with an octet.	

Example 6: Lewis structure for NH_3	
Total valence e^- N - has 5 valence electrons (5A) H - has 1 valence electrons (1A) NH_3 has total $5+3(1) = 8 e^-$	
Draw in the bonds All 8 electrons are accounted Everything is full ($2 e^-$ for H)	

Example 7: Lewis structure for PCl_3 The central atom is the least electronegative element	
Total valence e^- P - has 5 valence electrons Cl - has 7 valence electrons PCl_3 has total $5+3(7) = 26 e^-$	
Draw in the bonds All 8 electrons are accounted Everything is full	

(4-6)

Sulfur dichloride, SCl_2 , is a red, fuming liquid used in the manufacture of insecticides. Write the Lewis formula for the molecule.

.....

.....

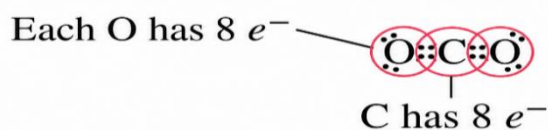
.....

.....

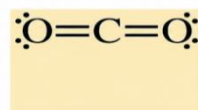
Double Bond

2 pairs of e^- 's shared between 2 atoms

Ex. CO_2



or

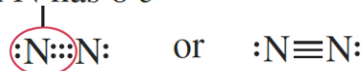


Triple bond

3 pairs of e⁻s shared between 2 atoms

Ex. N₂

Each N has 8 e⁻



Writing Lewis formulas for ions

Find the sum of valence electrons of all atoms in the polyatomic ion or molecule.

If it is an anion, add one electron for each negative charge.

If it is a cation, subtract one electron for each positive charge.

Example 8:

CO ₃ ²⁻	$\chi\text{C} = 2.5, \chi\text{O} = 3.5$	
$1\text{ C} = 1 \times 4\text{ e}^- = 4\text{ e}^-$ $3\text{ O} = 3 \times 6\text{ e}^- = 18\text{ e}^-$ $2\text{- charge} = 2\text{ e}^-$ Total = 24 e ⁻	Skeletal Structure	
single bonds - 6 e ⁻ 18 e ⁻	Complete terminal atom octets	
0 lone pairs -16 e ⁻ Charge 2 e ⁻	Place the structure in brackets and include the charge.	

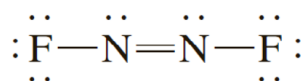
More Than One Central Atom

The molecules with more than one central atom are composites of the shapes around each of the central atoms.

Example 9:

N ₂ F ₂		
$2\text{ N} = 2 \times 5\text{ e}^- = 10\text{ e}^-$ $2\text{ F} = 2 \times 7\text{ e}^- = 14\text{ e}^-$ Total = 24 e ⁻	Skeletal Structure	$\text{F}-\text{N}-\text{N}-\text{F}$
F lone pairs - 12 e ⁻ 6 e ⁻	Complete terminal atom octets	$\text{:}\ddot{\text{F}}-\text{N}-\text{N}-\ddot{\text{F}}\text{:}$
N electrons - 6 e ⁻ 0 e ⁻	Put remaining e ⁻ s on central atoms	$\text{:}\ddot{\text{F}}-\ddot{\text{N}}-\ddot{\text{N}}-\ddot{\text{F}}\text{:}$

- Not enough electrons to complete N octets
- Must form double bond between N's to satisfy both octets.



(4-7)

Draw Lewis structure for ethylene (C₂H₄), the most important reactant in the manufacture of polymers.

.....

.....

.....

.....

(4-8)

Draw the Lewis structure for methanol (molecular formula CH₄O), an important industrial alcohol that can be used as a gasoline alternative in cars.

.....

.....

.....

Coordinate Covalent Bond

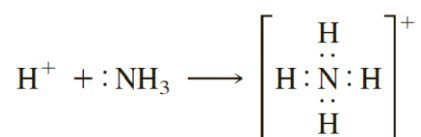
A coordinate bond (also called a dative covalent bond) is a covalent bond (a shared pair of electrons) in which **both** electrons come from the same atom.

Example 10: Ammonium Ion NH₄⁺

H⁺ has no electrons

N has lone pair

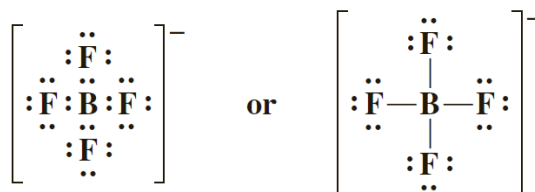
Can still get 2 e⁻s shared between them



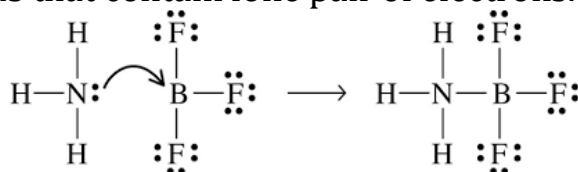
Example 11: Obtain the electron-dot formula (Lewis formula) of the BF₄⁻ ion.

$$n \text{ e}^- = 3 + (4 \times 7) = 31 \text{ e}^-$$

31+1 for charge = 32 (or 16 electron pairs).



Example 12: Especially boron (electron deficient molecule) reacts with nitrogen compounds that contain lone pair of electrons.



Exceptions to the Octet Rule in Lewis Structures

The octet rule applies to most molecules (especially with Period 2 central atoms), but important exceptions exist.

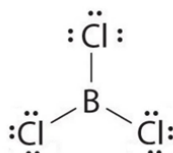
Three Types of Exceptions

1. Molecules with electron-deficient atoms.
2. Molecules with odd-electron atoms (free radicals).
3. Molecules with expanded valence shells (Period 3 or higher).

1. Electron-Deficient Atoms

- Found in molecules with beryllium (Be) or boron (B) as central atoms.
- These atoms often have fewer than 8 electrons around them.

Example 13: $\text{BCl}_3 \rightarrow$ only 6 electrons around B.



2. Molecules with Odd-Electron Atoms (Free Radicals)

- Species with an odd number of valence electrons.
- Central atom often comes from an odd-numbered group (e.g., N [Group 5A], Cl [Group 7A]).

Example 14: Nitrogen Dioxide (NO_2), Nitrogen monoxide (NO).



3. Expanded Valence Shells

Some molecules/ions have more than 8 valence electrons around the central atom.

- Expansion occurs when forming more bonds releases energy and the atom has available orbitals to accommodate them.
- Occurs only in nonmetals from Period 3 or higher (can use vacant 3d orbitals).

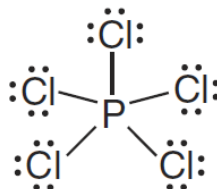
a) Central Atom Bonded to More Than Four Atoms

Example 15: PCl_5

P in PCl_3 has an octet.

Reacts with $\text{Cl}_2 \rightarrow$ expands to 10 electrons in PCl_5 .

Net effect: one extra bond formed.

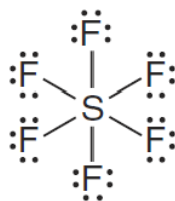


Example 16: SF_6

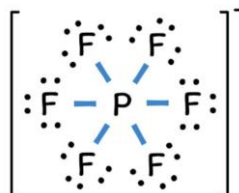
S bonded to six fluorine atoms.

Central S has 12 electrons (six bonds).

Stable, inert gas used as an electrical insulator.



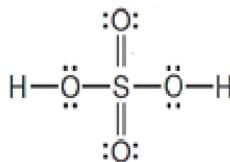
Example 17: PF_6^-



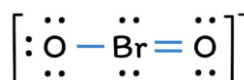
b) Central Atom Bonded to Four or Fewer Atoms

Expansion is also possible when bonded to ≤ 4 atoms.

Example 18: H_2SO_4 (sulfuric acid)

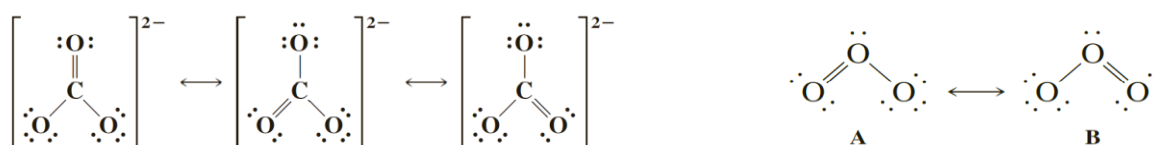


Example 19: BrO_2^-



Delocalized Bonding: Resonance

Resonance structures: Different valid Lewis structures can be drawn for the same molecule (e.g., O_3 , CO_3^{2-}), placing a double bond in different positions. Resonance structures are shown with a double-headed arrow ($\longleftrightarrow$). Resonance means the actual molecule is a **blend (hybrid)** of all valid structures.



Lewis Structures and Formal Charge

formal charge, which can be used to determine the most plausible Lewis structures when more than one possibility exists for a compound.

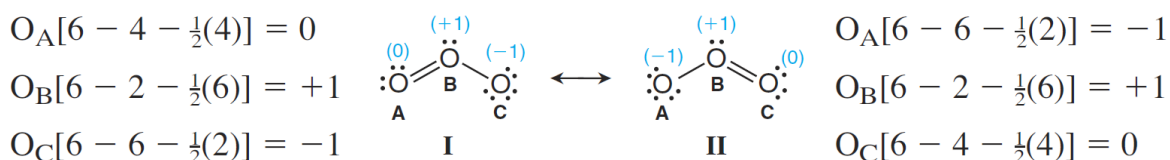
$$\text{Formal charge} = (\text{valence } e^-) - (\text{nonbonding } e^- + \frac{1}{2} \text{ bonding } e^-)$$

The sum of formal charges = total charge of the species (0 for neutral molecules, equals the ion charge otherwise).

For example, in O_3 , the formal charge of oxygen A in resonance form I is:

$$6 \text{ valence } e^- - (4 \text{ unshared } e^- + \frac{1}{2} \text{ of } 4 \text{ shared } e^-) = 6 - 4 - 2 = 0$$

The formal charges of all the atoms in the two O_3 resonance forms are:

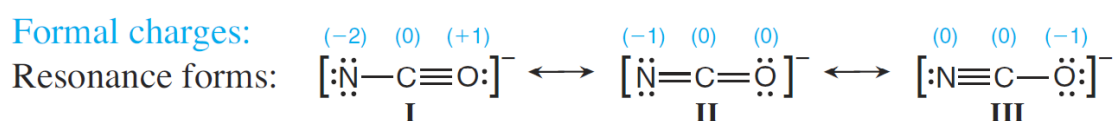


Choosing the More Important Resonance

the more important resonance structure:

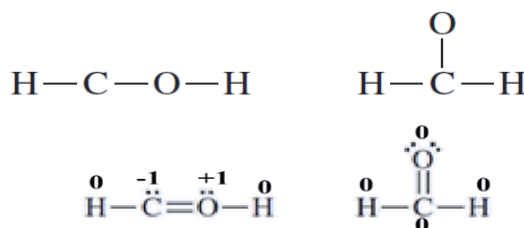
- Formal charges on atoms are preferably small (close to zero) rather than large.
- The same nonzero formal charges on adjacent atoms are not preferred.
- A more negative formal charge should reside on a more electronegative atom.

For NCO^- , which has two different atoms around the central one, Three resonance forms with formal charges:

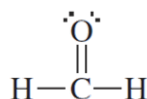


Form I is not an important contributor to the hybrid because it has a larger formal charge on N and a positive formal charge on the more electronegative O. Forms II and III have the same magnitude of charges, but III has a -1 charge on O, the more electronegative atom, Therefore III is more important than II.

Example 20: In formaldehyde (CH_2O) Use formal charges to determine which skeletal arrangement of atoms shown here is the best choice for the Lewis structure of CH_2O .



The form with formal charges all zero and is the better choice for the Lewis structure of CH_2O .



(4-9)

Draw a Lewis structure and identify the octet-rule exception for H_3PO_4 (draw two resonance forms and select the more important).

.....

.....

.....

.....

Valence-Shell Electron-Pair Repulsion (VSEPR) Theory and Molecular Shape

The shape of molecules is crucial in processes like medicine, smell, and taste, as one molecule often must fit with another. To predict molecular shape, chemists use the Lewis structure and apply VSEPR theory. Its main idea is that electron groups around a central atom arrange themselves as far apart as possible to minimize repulsion. An electron group can be a bond (single, double, or triple) or a lone pair, but each count as one group. The molecular shape is the 3D arrangement of the nuclei connected by these groups.

Electron-Group Arrangements and Molecular Shapes

When 2–6 groups are placed around a central atom, they arrange to maximize space and minimize repulsion. This leads to **five main electron-group arrangements**, found in most molecules and polyatomic ions (Figure 4.9).

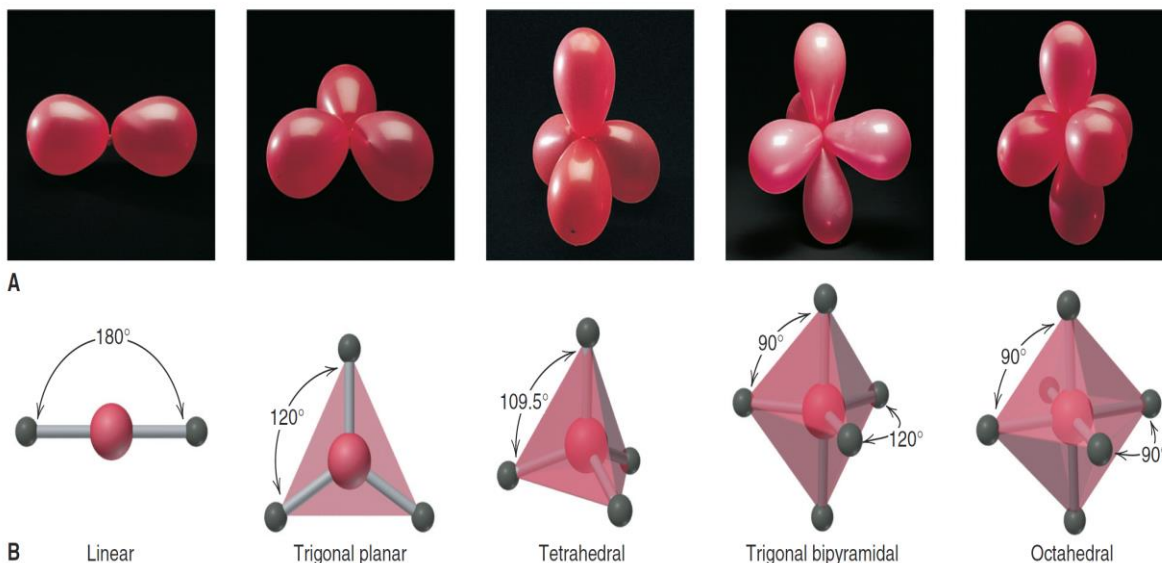


Figure 4.9 Electron-group repulsions and molecular shapes

The Molecular Shape with Two Electron Groups (Linear Arrangement)

Two electron groups attached to a central atom point in opposite directions. This linear arrangement of electron groups results in a molecule with a linear shape and a bond angle of 180° .

Example 21: Gaseous beryllium chloride (BeCl_2) is a linear molecule. Recall that the central Be atom is electron deficient, with two electron pairs around it.



Example 22: In carbon dioxide, the central C atom forms two double bonds with the O atoms.



Each double bond acts as one electron group and is 180° away from the other. The lone pairs on the O atoms of CO_2 or on the Cl atoms of BeCl_2 are not involved in the molecular shape: only electron groups around the central atom affect shape.

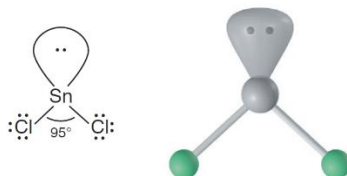
Molecular Shapes with Three Electron Groups (Trigonal Planar Arrangement)

1. All bonding groups: trigonal planar shape (AX_3).

Example 23: BF_3 , with three single bonds and 120° bond angles. NO_3^- also has a trigonal planar shape.

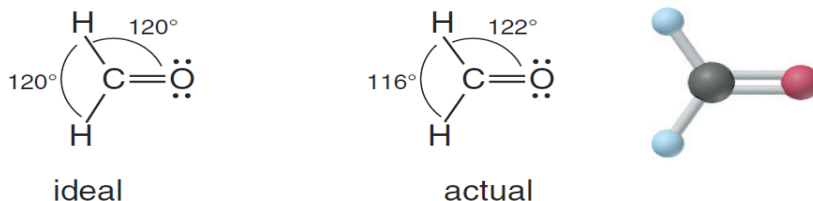


2. One lone pair: bent or V shape (AX_2E), a lone pair repels bonding pairs more than bonding pairs repel each other, so it decreases the angle between bonding pairs. Note the 95° bond angle in SnCl_2 , which is considerably less than the ideal 120° .



Effect of Double Bonds on Bond Angle

In molecules with different atoms or electron groups, bond angles shift. In formaldehyde (CH_2O), the trigonal planar arrangement ideally has 120° angles. But the double bond has greater electron density, repelling single bonds more strongly. As a result, the $\text{H}-\text{C}-\text{H}$ angle shrinks to 116° , while the $\text{O}-\text{C}-\text{H}$ angle widens to 122° .

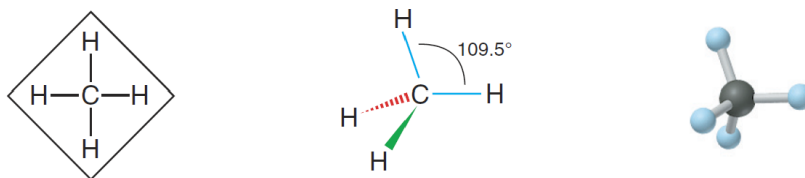


Molecular Shapes with Four Electron Groups (Tetrahedral Arrangement)

Shapes based on two or three electron groups lie in a plane, but four electron groups require three dimensions to maximize separation.

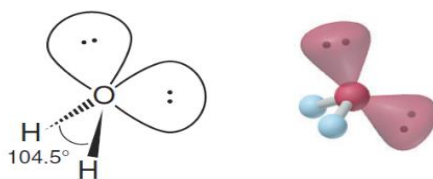
1. All bonding groups: tetrahedral shape (AX_4).

Example 24: CH_4 (methane), common in organic molecules.



2. One lone pair: trigonal pyramidal shape (AX_3E). Ammonia (NH_3) is an example, where its shape is a trigonal pyramid, and the $\text{H}-\text{N}-\text{H}$ bond angle decreases from 109.5° to 107.3° .

3. Two lone pairs: bent or V shape (AX_2E_2). Water is the most important V-shaped molecule with the tetrahedral arrangement.

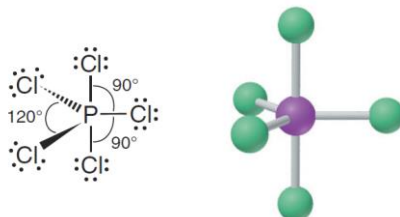


Molecular Shapes with Five Electron Groups (Trigonal Bipyramidal)

Molecules with five or six electron groups have a central atom from Period 3 or higher because only those atoms have d orbitals available to expand the valence shell.

When there are five groups (AX_5), the molecule has a trigonal bipyramidal arrangement. Three bonds are arranged in the trigonal plane at 120° (in the plane of the central atom), and two axial bonds are positioned above and below the plane at 90° to the equatorial bonds. The repulsions between the equatorial groups are weaker, so lone pairs and double bonds prefer to occupy equatorial positions.

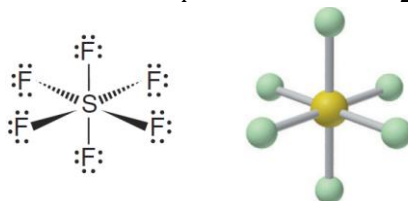
Example 25: PCl_5 , with ideal 90° and 120° bond angles.



Molecular Shapes with Six Electron Groups (Octahedral Arrangement)

Six electron groups form the **octahedral arrangement**. An octahedron is a polyhedron with eight equilateral triangles for faces and six identical vertices. Each of the six groups points to a corner, which gives a 90° ideal bond angle.

Example 26: SF_6 , formed when SF_4 reacts with F_2 .



(4-10)

Draw the molecular shapes and predict the bond angles (relative to the ideal angles) $COCl_2$.

.....

.....

Exercise answers

1. (a) Bond length: S-Br > S-Cl > S-F

Bond strength: S-F > S-Cl > S-Br

(b) Bond length: C-O > C=O > C≡O

Bond strength: C≡O > C=O > C-O

2. (a) Bond length: Si-F < Si-O < Si-C

Bond strength: Si-C < Si-O < Si-F

(b) Bond length: N≡N < N=N, N-N

Bond strength: N-N < N=N < N≡N

3. (a) (NH₄)₃PO₄ ammonium phosphate

(b) Al(NO₂)₃, Aluminum nitrite

4. (a) $\overset{+}{\text{N}}-\text{H}$, $\text{F}-\overset{+}{\text{N}}$, $\overset{+}{\text{I}}-\text{Cl}$

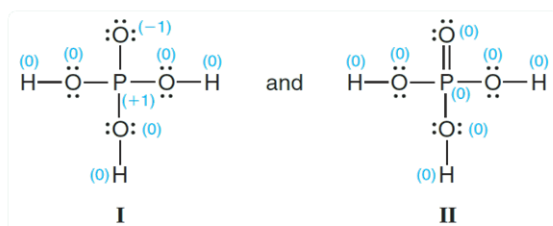
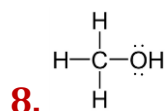
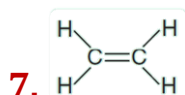
(b) H-C < H-N < H-O

H-O > H-N > H-C

5. (a) Cl-Cl < Br^{δ+}-Cl^{δ-} < Cl^{δ+}-F^{δ-}

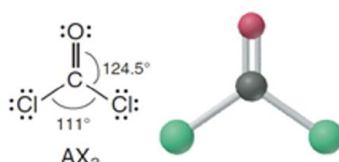
(b) Si-Si < S^{δ+}-Cl^{δ-} < P^{δ+}-Cl^{δ-} < Si^{δ+}-Cl^{δ-}

6. $\text{:}\ddot{\text{Cl}}\text{:}\ddot{\text{S}}\text{:}\ddot{\text{Cl}}\text{:}$ or $\text{:}\ddot{\text{Cl}}-\ddot{\text{S}}-\ddot{\text{Cl}}\text{:}$



Structure II is the more important form.

10.



1.	2.	3.	4.	5.	6.	7.	8.	9.	10.	11.	12.	13.	14.	15.	16.	17.	18.
PERIODIC TABLE																	
I. A																	
1.	2.	3.	4.	5.	6.	7.	8.	9.	10.	11.	12.	13.	14.	15.	16.	17.	18.
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
H	He																
1.0	4.0																
II. A																	
3	4																
Li	Be																
6.9	9.0																
III. A																	
11	12																
Na	Mg																
23.0	24.3																
IV. A																	
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
39.1	40.1	45.0	47.9	50.9	52.0	54.9	55.8	58.9	58.7	63.5	65.4	69.7	72.6	74.9	79.0	79.9	83.8
V. A																	
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
85.5	87.6	88.9	91.2	92.9	95.9	[98]	101.1	102.9	106.4	107.9	112.4	114.8	118.7	121.8	127.6	126.9	131.3
VI. A																	
55	56	57	58	59	60	61	62	63	64	65	66	67	68	69	70	71	72
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
132.9	137.3	138.9	178.5	180.9	183.8	186.2	190.2	192.2	195.1	197.0	200.6	204.4	207.2	209.0	[209]	[210]	[222]
VII. A																	
87	88	89	90	91	92	93	94	95	96	97	98	99	100	101	102	103	104
Fr	Ra	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	
[223]	[226]	[227]	232.0	231.0	238.0	[237]	[244]	[243]	[247]	[247]	[251]	[252]	[257]	[258]	[259]	[262]	[294]
VIII. A																	
2																	
He																	
4.0																	

1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1	1.1
58	59	60	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75
Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu				
140.1	140.9	144.2	[145]	150.4	152.0	157.3	158.9	162.5	164.9	167.3	168.9	173.0	175.0				
90	91	92	93	94	95	96	97	98	99	100	101	102	103	104	105	106	107
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr				
232.0	231.0	238.0	[237]	[244]	[243]	[247]	[247]	[251]	[252]	[257]	[258]	[259]	[262]				

* EN: electronegativity

Scientific Sources and References

- Ministry of Education. Chemistry Textbook Series for the Secondary Stage (Pathways System). Kingdom of Saudi Arabia, 2025 .
- Burdge , Julia and Jason Overby . Chemistry - atoms first , Third Edition . USA : McGraw-Hill , 2018 .
- Chang, Raymond. Chemistry, Tenth Edition. USA: McGraw-Hill, 2010.
- Ebbing , Darrell D and Steven D. Gammon . General Chemistry , Eleventh Edition . USA : Cengage Learning , 2015 .
- Rosenberg , Jerome L and Others . Schaum's Outline Theory and Problems of College Chemistry , Ninth Edition . USA : McGraw-Hill , 2007 .
- Silberberg , Martin S . Principles of General Chemistry , Third Edition . USA : McGraw-Hill , 2013 .
- Villanyi , Attila . How to Get an A in Chemistry ? , Budapest : Muszaki KonyvKiado , 2021 .
- Zumdahl , Steven S and Susan A. Zumdahl . Chemistry , Ninth Edition . USA : Brooks Cole , 2014 .

