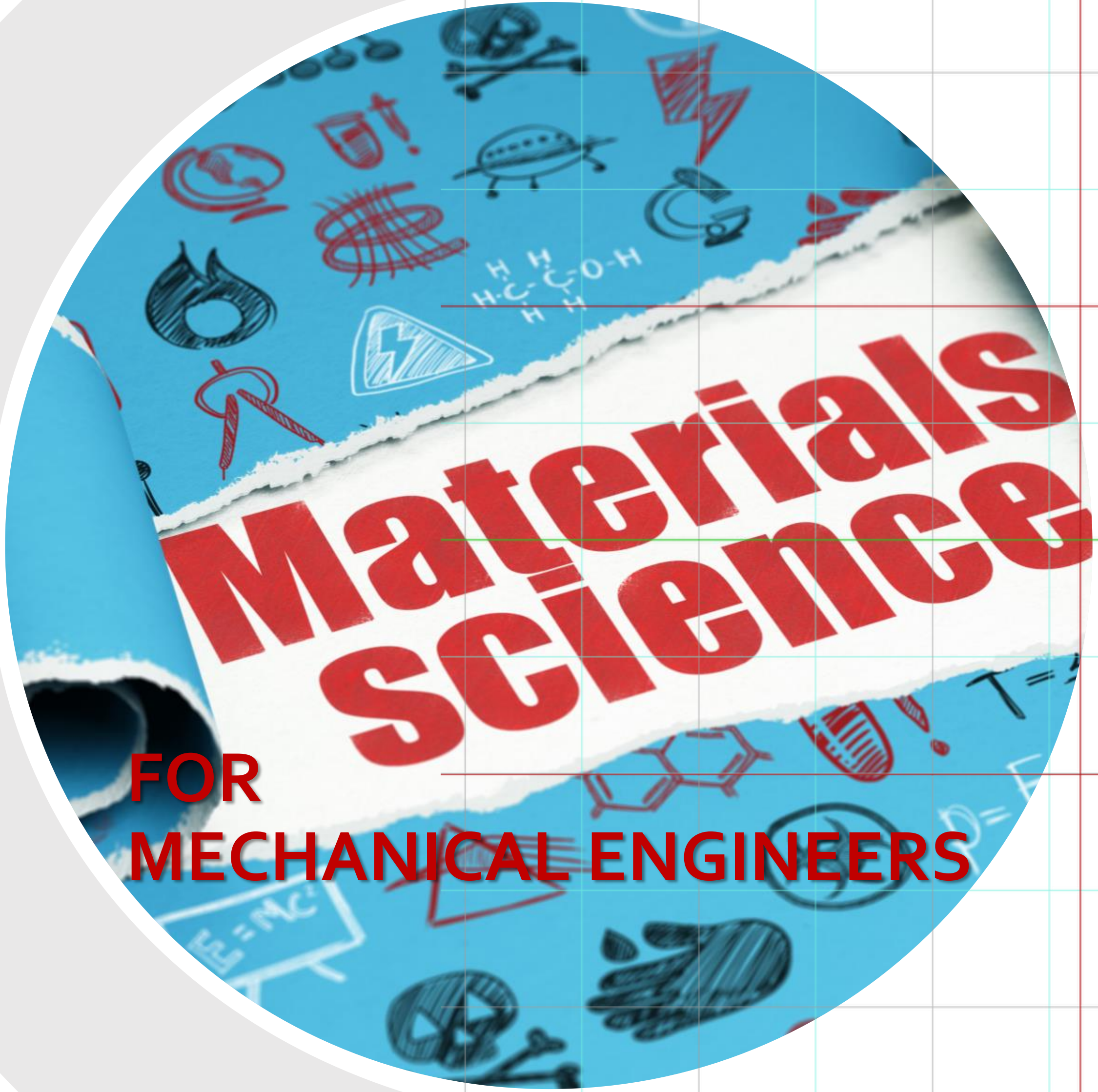


FOR
MECHANICAL ENGINEERS

Atomic Structure and Bonding

SPRING 2022-2023

Introduction

Learning Objectives

1. Name the two atomic models cited, and note the differences between them.

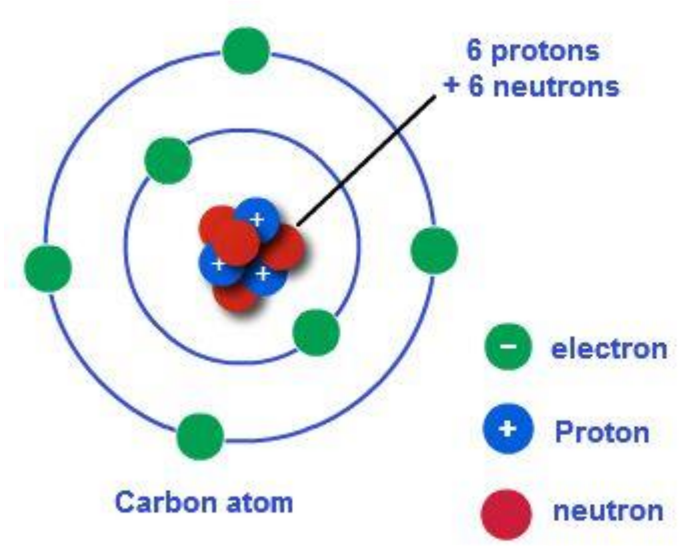
2. Describe the important quantum-mechanical principle that relates to electron energies.

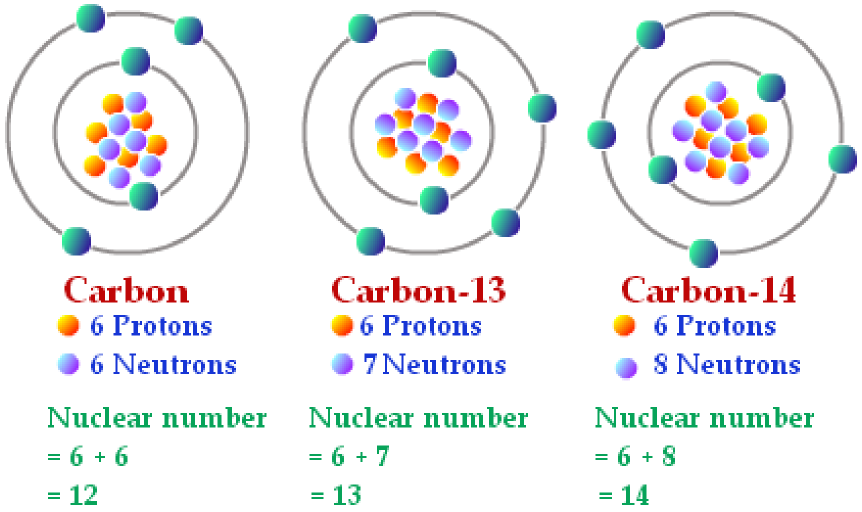
3. (a) Schematically plot attractive, repulsive, and net energies versus interatomic separation for two atoms or ions.

(b) Note on this plot the equilibrium separation and the bonding energy.

4. (a) Briefly describe ionic, covalent, metallic, hydrogen, and van der Waals bonds.

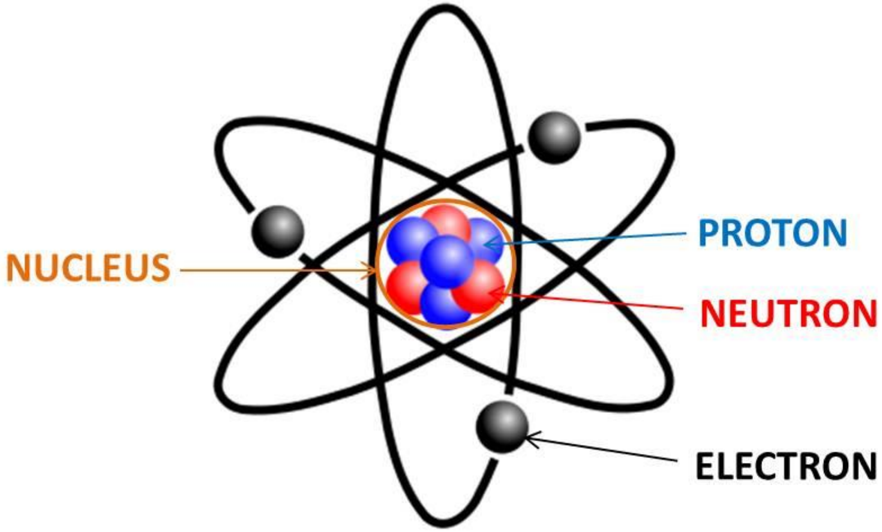
(b) Note which materials exhibit each of these bonding types.





Fundamental Concepts

Each atom consists of a very small **nucleus** composed of **protons** and **neutrons** and is encircled by moving electrons. Both **electrons** and protons are electrically charged, the charge magnitude being 1.602×10^{-19} C, which is negative in sign for electrons and positive for protons; neutrons are electrically neutral.



Masses for these subatomic particles are extremely small; protons and neutrons have approximately the same mass, 1.67×10^{-27} kg, which is significantly larger than that of an electron, 9.11×10^{-31} kg.

Each chemical element is characterized by the number of protons in the nucleus, or the **atomic number (*Z*).**

For an electrically neutral or complete atom, the atomic number also equals the number of electrons.

This atomic number ranges in integral units from 1 for hydrogen to 92 for uranium, the highest of the naturally occurring elements.

The *atomic mass* (A) of a specific atom may be expressed as the *sum of the masses of protons and neutrons within the nucleus*.

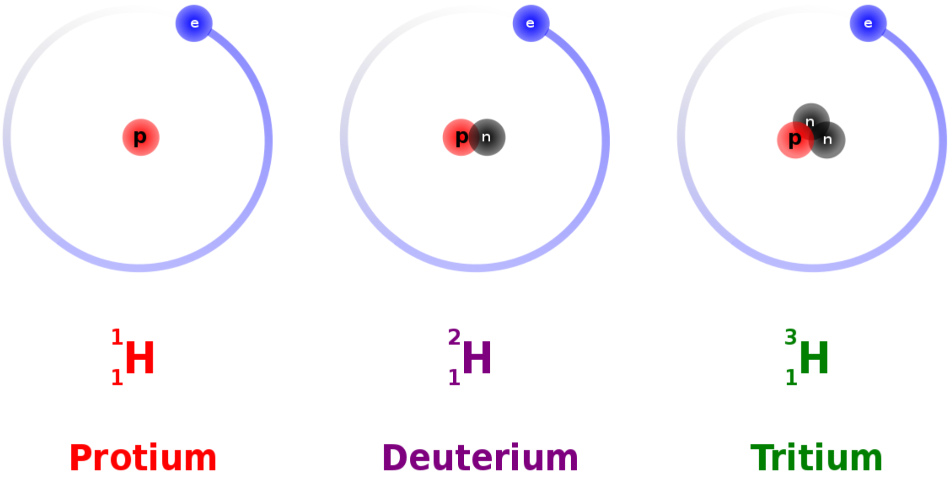
Although the number of protons is the same for all atoms of a given element, the number of neutrons (N) may be variable.

Fundamental Concepts

Atoms of some elements have two or more different atomic masses, which are called **isotopes**.

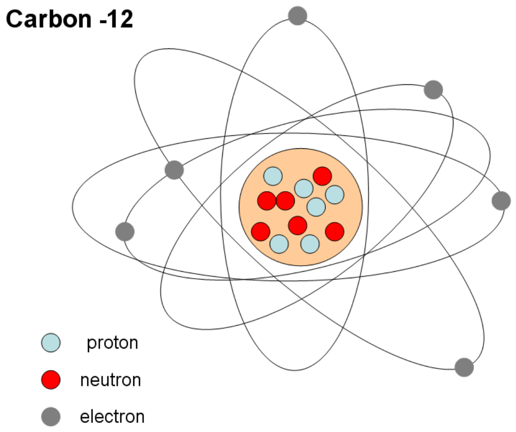
The **atomic weight** of an element corresponds to the weighted average of the atomic masses of the atom's naturally occurring isotopes.

The **atomic mass unit** (amu) may be used to compute atomic weight.



Isotopes of Hydrogen

A scale has been established whereby 1 amu is defined as 1/12 of the atomic mass of the most common isotope of carbon, carbon 12 (^{12}C) (A = 12.00000).



Each chemical element is characterized by the number of protons in the nucleus, or the **atomic number (*Z*).**

The masses of protons and neutrons are slightly greater than unity:

$$A \cong Z + N$$

The atomic weight of an element or the molecular weight of a compound may be specified on the basis of amu per atom (molecule) or mass per mole of material.

Fundamental Concepts

In one **mole** of a substance, there are 6.022 × 10^{23} (Avogadro's number) atoms or molecules.

These two atomic weight schemes are related through the following equation:

$$1 \text{ amu/atom (or molecule)} = 1 \text{ g/mol}$$

For example, the atomic weight of iron is 55.85 amu/atom, or 55.85 g/mol.

Sometimes use of amu per atom or molecule is convenient; on other occasions, grams (or kilograms) per mole is preferred.

ELECTRONS IN ATOMS

Many phenomena involving electrons in solids could not be explained in terms of classical mechanics.

Thus the establishment of a set of principles and laws that govern systems of atomic and subatomic entities that came to be known as **quantum mechanics.**

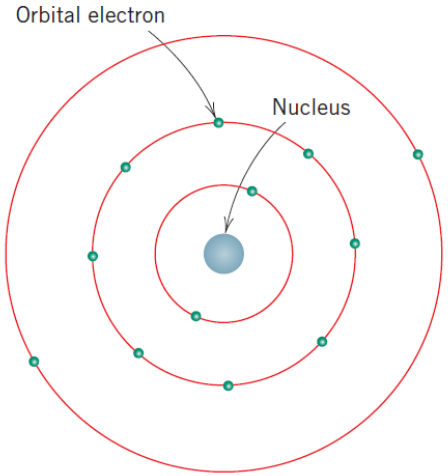
An understanding of the behaviour of electrons in atoms and crystalline solids necessarily involves the discussion of quantum-mechanical concepts.

One early outgrowth of quantum mechanics was the simplified **Bohr atomic model,** in which electrons are assumed to revolve around the atomic nucleus in discrete orbitals, and the position of any particular electron is more or less well defined in terms of its orbital.

Another important quantum-mechanical principle stipulates that the energies of electrons are quantized:

ELECTRONS IN ATOMS

electrons are permitted to have only specific values of energy.



Schematic representation of the Bohr atom

An electron may change energy, but in doing so, it must make a quantum jump either to an allowed higher energy (with absorption of energy) or to a lower energy (with emission of energy).

Often, it is convenient to think of these allowed electron energies as being associated with **energy levels or states**.

These energies are taken to be negative, whereas the zero reference is the unbound or free electron.

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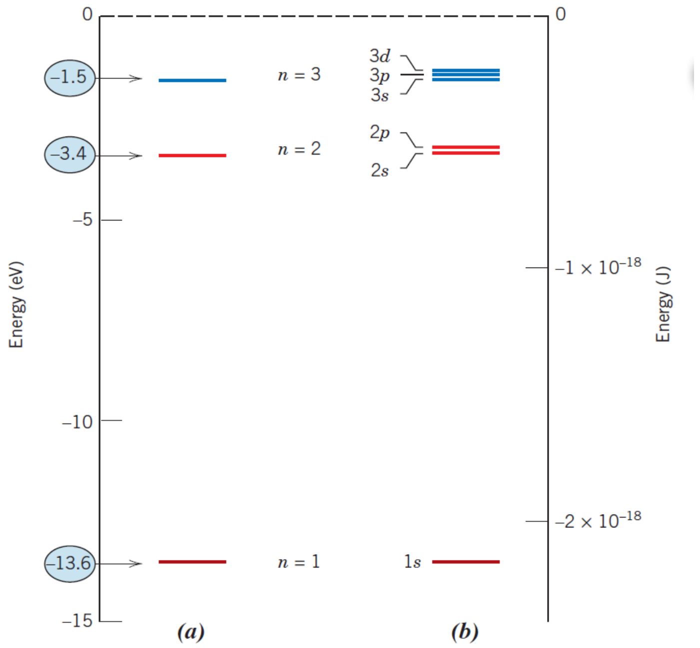
These energies are taken to be negative, whereas the zero reference is the unbound or free electron.

For example, allowed states for the Bohr hydrogen atom are represented in the figure below (a).



ELECTRONS IN ATOMS

These energies are taken to be negative, whereas the zero reference is the unbound or free electron. Of course, the single electron associated with the hydrogen atom fills only one of these states.

Thus, the Bohr model represents an early attempt to describe electrons in atoms, in terms of both position (electron orbitals) and energy (quantized energy levels).

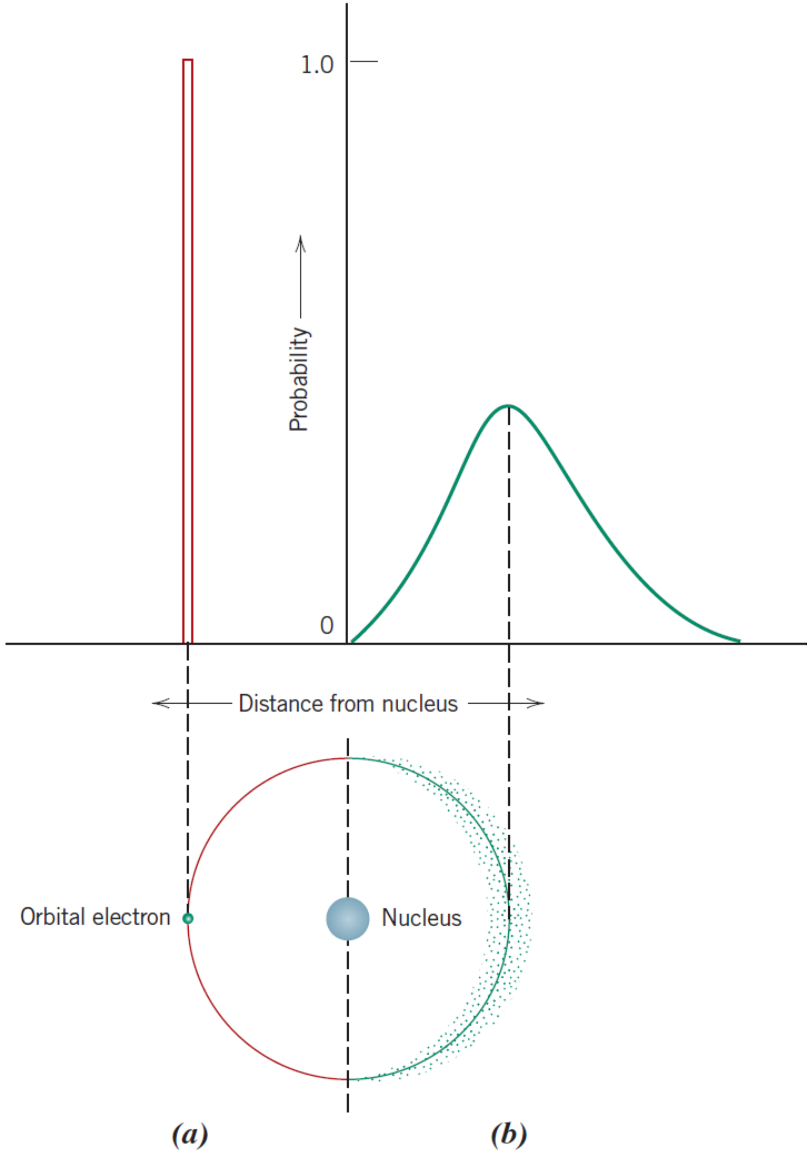
This Bohr model was eventually found to have some significant limitations because of its inability to explain several phenomena involving electrons.

A resolution was reached with a **wave-mechanical model**, in which the electron is considered to exhibit both wavelike and particle-like characteristics.

With this model, an electron is no longer treated as a particle moving in a discrete orbital; rather, position is considered to be the probability of an electron being at various locations around the nucleus.

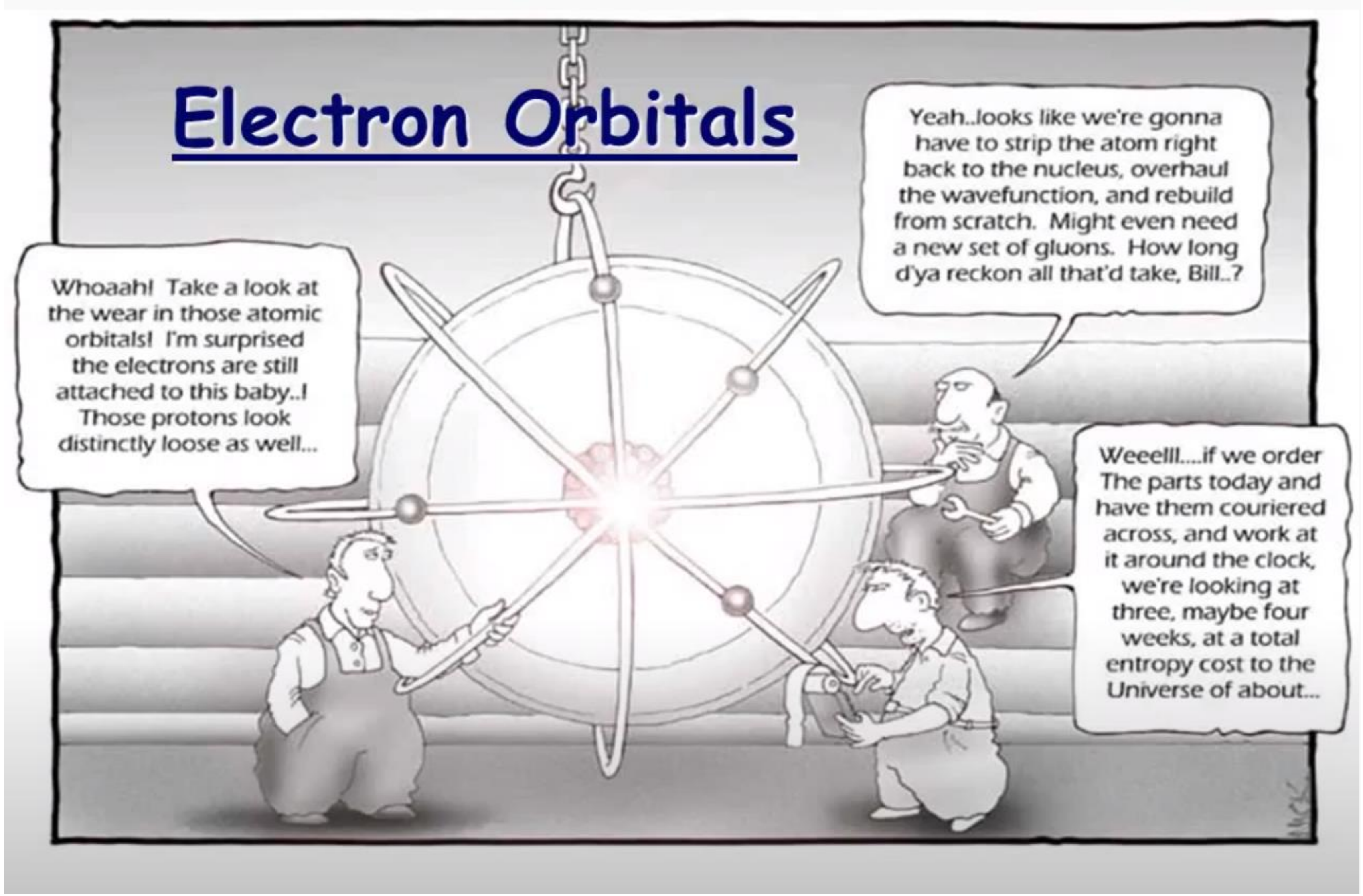
Position is described by a probability distribution or electron cloud. Here is a comparison between Bohr and wave-mechanical models for the hydrogen atom.

Both models are used throughout the course of this text; the choice depends on which model allows the simplest explanation.



ELECTRONS IN ATOMS

Quantum Numbers



ELECTRONS IN ATOMS

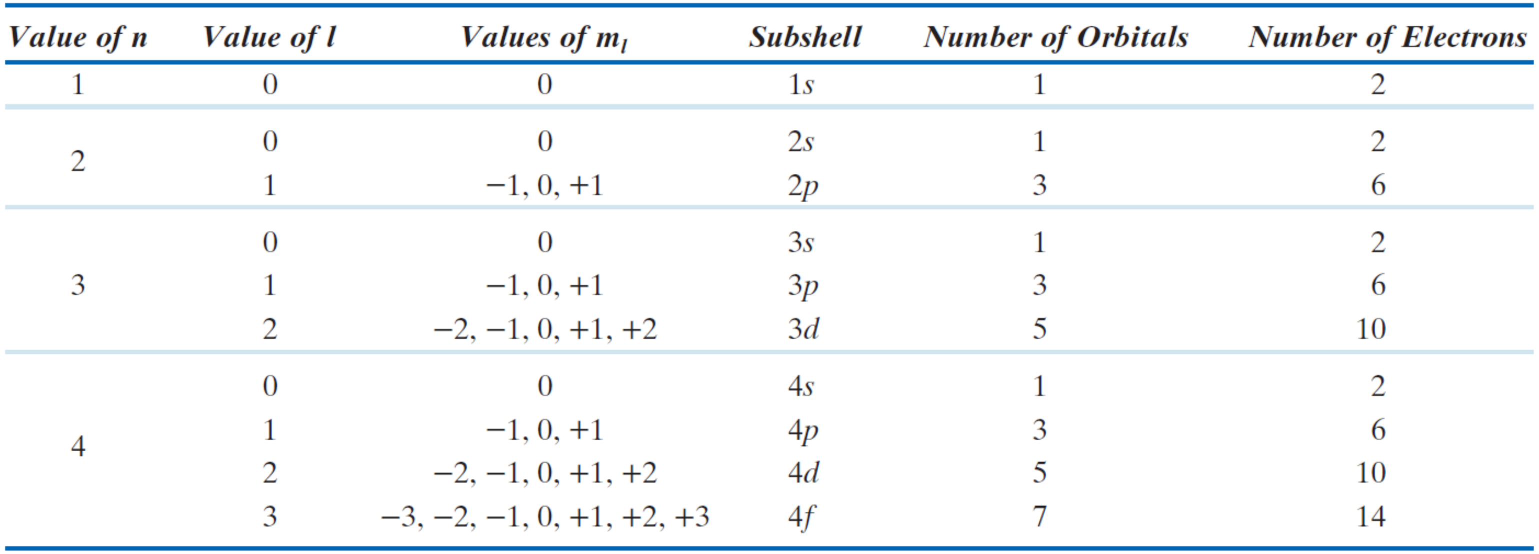
Quantum Numbers

In wave mechanics, every electron in an atom is characterized by four parameters called quantum numbers.

The size, shape, and spatial orientation of an electron’s probability density (or orbital) are specified by three of these quantum numbers.

Furthermore, Bohr energy levels separate into electron subshells, and quantum numbers dictate the number of states within each subshell.

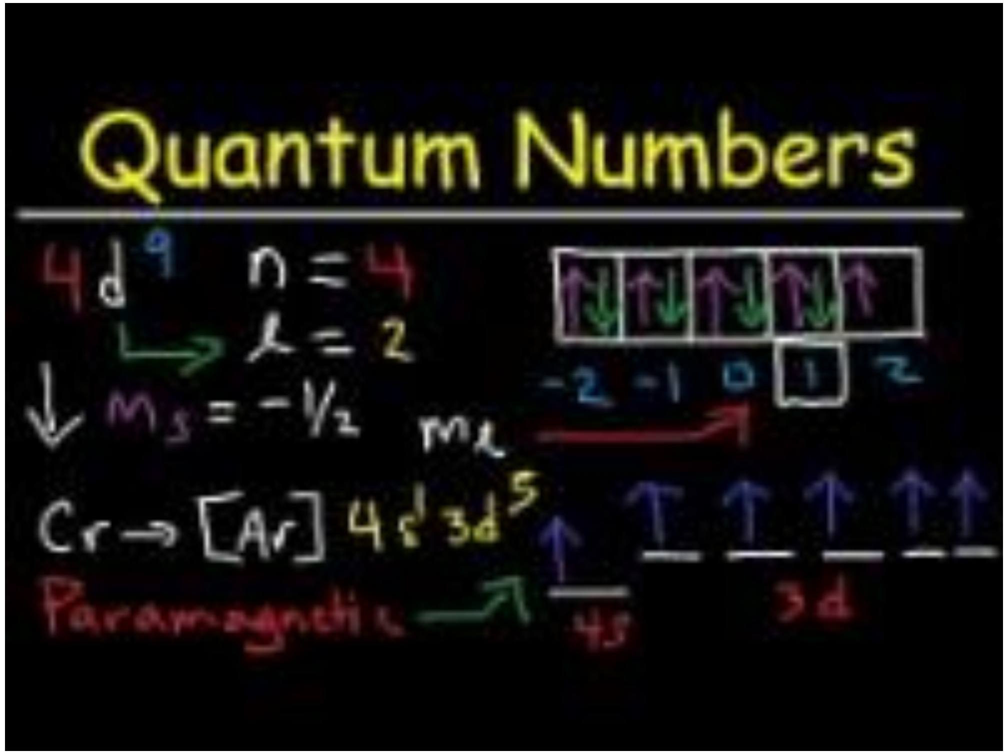
Shells are specified by a principal quantum number n, which may take on integral values beginning with unity; sometimes these shells are designated by the letters K, L, M, N, O, and so on, which correspond, respectively, to n = 1, 2, 3, 4, 5, . . . , as indicated in the Table.



Value of n	*Value of l*	*Values of* m_l	*Subshell*	*Number of Orbitals*	*Number of Electrons*
1	0	0	1*s*	1	2
2	0	0	2*s*	1	2
	1	−1, 0, +1	2*p*	3	6
3	0	0	3*s*	1	2
	1	−1, 0, +1	3*p*	3	6
	2	−2, −1, 0, +1, +2	3*d*	5	10
4	0	0	4*s*	1	2
	1	−1, 0, +1	4*p*	3	6
	2	−2, −1, 0, +1, +2	4*d*	5	10
	3	−3, −2, −1, 0, +1, +2, +3	4*f*	7	14

ELECTRONS IN ATOMS

Quantum Numbers



Quantum Numbers - The Easy Way! - YouTube

ELECTRONS IN ATOMS

Quantum Numbers

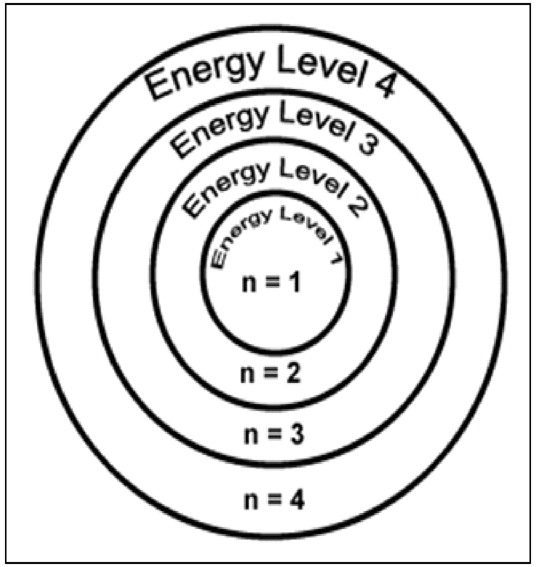
Each electron in an atom has a unique set of 4 quantum numbers which describe it.

- Principal quantum number
- Angular momentum quantum number
- Magnetic quantum number
- Spin quantum number

ELECTRONS IN ATOMS

Quantum Numbers

Electron Energy Level - Shell
Generally symbolized by n, it denotes the probable distance of the electron from the nucleus. “n” is also known as the **Principal Quantum number**



Number of electrons that can fit in a shell: $2n^2$

ELECTRONS IN ATOMS

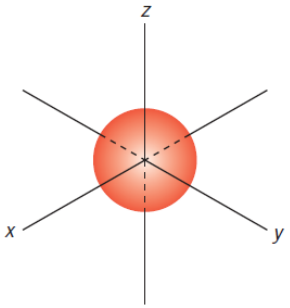
Quantum Numbers

Electron Orbitals - ℓ

An orbital is a region within an energy level where there is a probability of finding an electron.

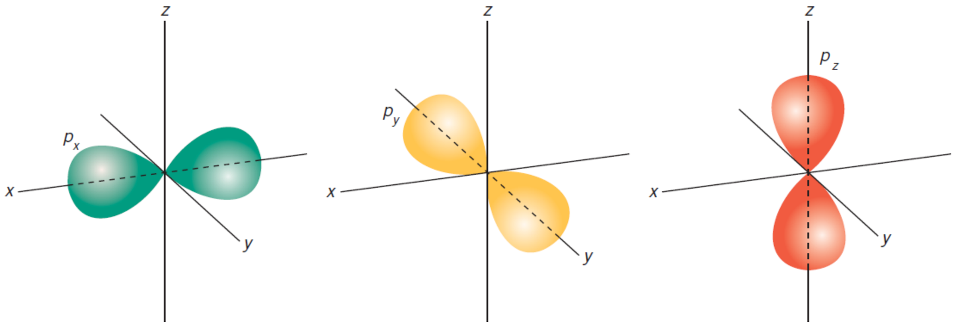
Orbital shapes are defined as the surface that contains 90% of the total electron probability.

The angular momentum quantum number, generally symbolized by ℓ, denotes the orbital (subshell) in which the electron is located.



s - Orbital Shape: The *s* orbital (ℓ = 0) has a spherical shape centered around the origin of the three axes in space.

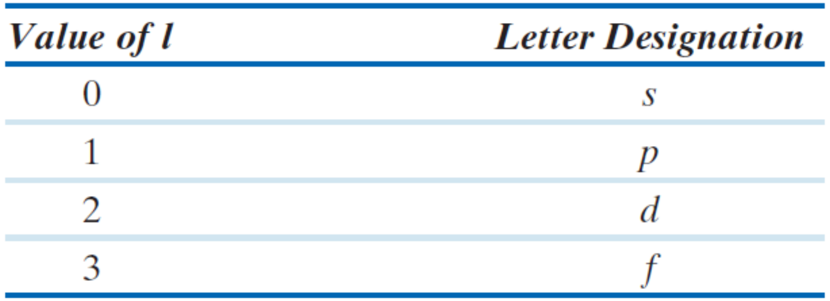
There are three dumbbell-shaped p orbitals (ℓ = 1) in each energy level above n = 1, each assigned to its own axis (x, y and z) in space.



ELECTRONS IN ATOMS

The second (or azimuthal) quantum number, *l*, designates the subshell. Values of *l* are restricted by the magnitude of n and can take on integer values that range from $l = 0$ to $l = (n - 1)$.

Each subshell is denoted by a lowercase letter—an s, p, d, or f—related to l values as follows:



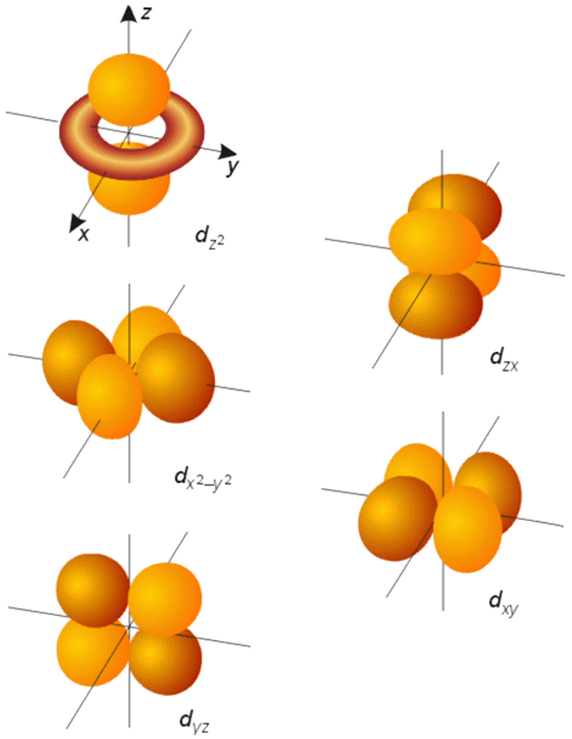
Value of l	*Letter Designation*
0	*s*
1	*p*
2	*d*
3	*f*

Furthermore, electron orbital shapes depend on *l*. For example *s* orbitals are spherical and centred on the nucleus.

There are three orbitals for a *p* subshell; each has a nodal surface in the shape of a dumbbell.

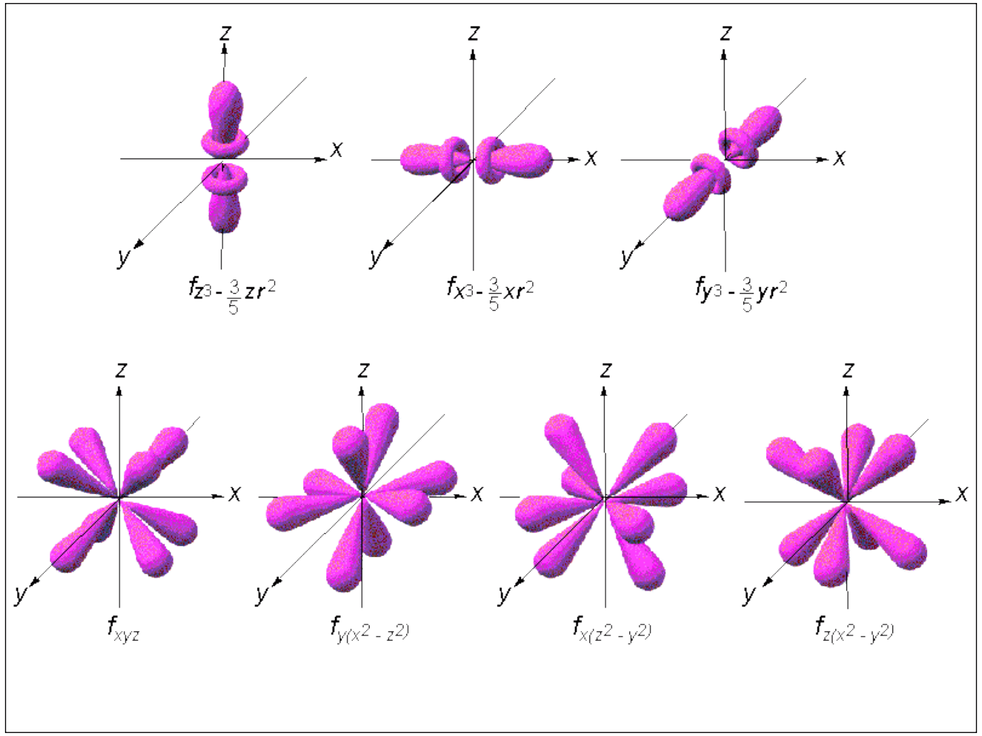
Axes for these three orbitals are mutually perpendicular to one another like those of an x-y-z coordinate system; thus, it is convenient to label these orbitals px, py, and pz.

Orbital configurations for d subshells are more complex.



Things get a bit more complicated with the five d orbitals ($\ell = 2$) that are found in the d sublevels beginning with n = 3. To remember the shapes, think of “double dumbbells”, and a “dumbbell with a donut”!

ELECTRONS IN ATOMS



Things are even more complicated with the seven *f* orbitals ($\ell = 3$)

ELECTRONS IN ATOMS

Energy Level (n)	Sublevels in main energy level (n sublevels)	Number of orbitals per sublevel	Number of Electrons per sublevel	Number of electrons per main energy level ($2n^2$)
1	s	1	2	2
2	s p	1 3	2 6	8
3	s p d	1 3 5	2 6 10	18
4	s p d f	1 3 5 7	2 6 10 14	32

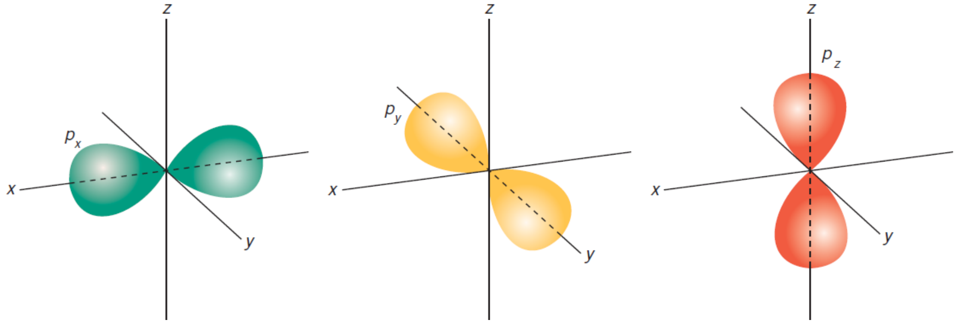
ELECTRONS IN ATOMS

The number of electron orbitals for each subshell is determined by the third (or magnetic) quantum number, m_l; m_l can take on integer values between $-l$ and $+l$, including 0. It denotes the orientation of the electron's orbital with respect to the three axes in space.

When $l = 0$, m_l can only have a value of 0 because +0 and −0 are the same.

This corresponds to an *s* subshell, which can have only one orbital.

Furthermore, for $l = 1$, m_l can take on values of −1, 0, and +1, and three *p* orbitals are possible.



Similarly, it can be shown that *d* subshells have five orbitals, and *f* subshells have seven.

In the absence of an external magnetic field, all orbitals within each subshell are identical in energy.

However, when a magnetic field is applied, these subshell states split, with each orbital assuming a slightly different energy.

Associated with each electron is a *spin moment*, which must be oriented either up or down. Related to this spin moment is the fourth quantum number, m_s (quantum spin), for which two values are possible: +1/2 (for spin up) and −1/2 (for spin down).

Thus, the Bohr model was further refined by wave mechanics, in which the introduction of three new quantum numbers gives rise to electron subshells within each shell.

PAULI EXCLUSION PRINCIPLE

Wolfgang Pauli

Two electrons occupying the same orbital must have *opposite* spins

ELECTRONS IN ATOMS

Element	Configuration notation	Orbital notation	Noble gas notation
Lithium	$1s^22s^1$	_____ _____ _____ _____ _____ 1s 2s 2p	$[He]2s^1$
Beryllium	$1s^22s^2$	_____ _____ _____ _____ _____ 1s 2s 2p	$[He]2s^2$
Boron	$1s^22s^2p^1$	_____ _____ _____ _____ _____ 1s 2s 2p	$[He]2s^2p^1$
Carbon	$1s^22s^2p^2$	_____ _____ _____ _____ _____ 1s 2s 2p	$[He]2s^2p^2$
Nitrogen	$1s^22s^2p^3$	_____ _____ _____ _____ _____ 1s 2s 2p	$[He]2s^2p^3$
Oxygen	$1s^22s^2p^4$	_____ _____ _____ _____ _____ 1s 2s 2p	$[He]2s^2p^4$
Fluorine	$1s^22s^2p^5$	_____ _____ _____ _____ _____ 1s 2s 2p	$[He]2s^2p^5$
Neon	$1s^22s^2p^6$	_____ _____ _____ _____ _____ 1s 2s 2p	$[He]2s^2p^6$

ELECTRONS IN ATOMS

- The three quantum numbers (n, ℓ, and m) are integers.
- The principal quantum number (n) cannot be zero.
- n must be 1, 2, 3, etc.
- The angular momentum quantum number (ℓ) can be any integer between 0 and n - 1.
- For $n = 3$, ℓ can be either 0, 1, or 2.
- The magnetic quantum number (m_l) can be any integer between $-\ell$ and $+\ell$.
- For $\ell = 2$, m can be either -2, -1, 0, +1, +2.

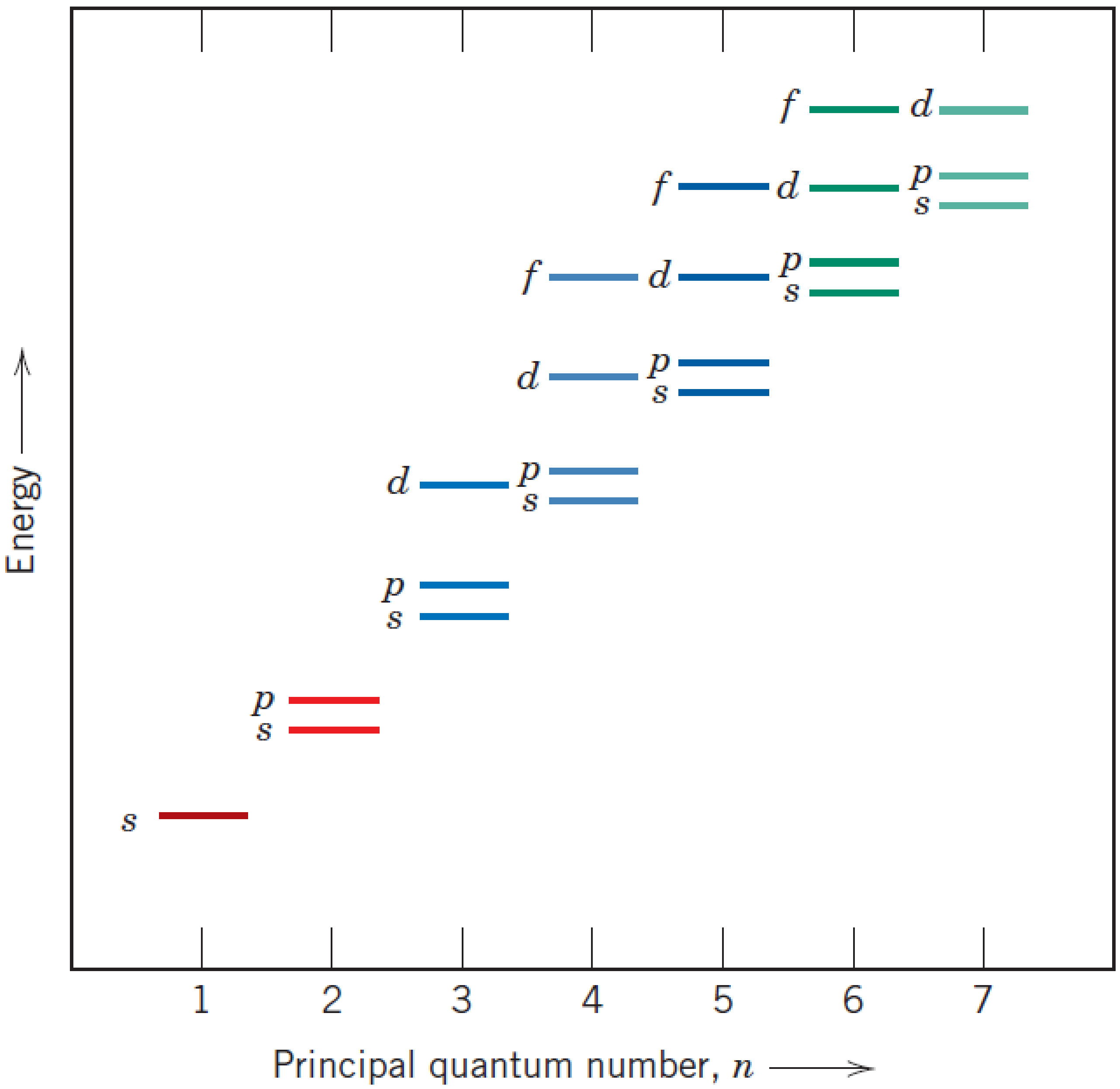
ELECTRONS IN ATOMS

A complete energy level diagram for the various shells and subshells using the wave-mechanical model is shown.

Several features of the diagram are worth noting. First, the smaller the principal quantum number, the lower the energy level; for example, the energy of a 1*s* state is less than that of a 2*s* state, which in turn is lower than that of the 3*s*.

Second, within each shell, the energy of a subshell level increases with the value of the *l* quantum number. For example, the energy of a 3*d* state is greater than that of a 3*p*, which is larger than 3*s*.

Finally, there may be overlap in energy of a state in one shell with states in an adjacent shell, which is especially true of *d* and *f* states; for example, the energy of a 3*d* state is generally greater than that of a 4*s*.



ELECTRONS IN ATOMS

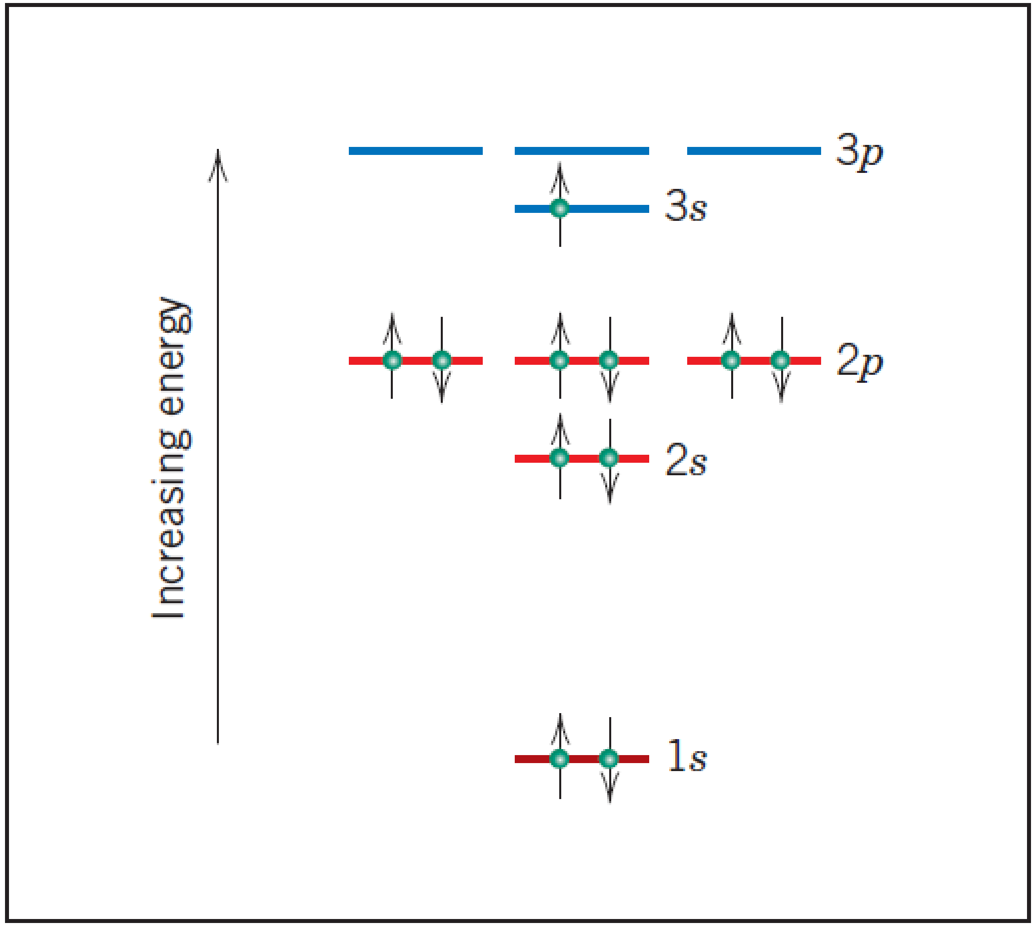
Electron Configurations

To determine the manner in which electron states (values of energy that are permitted for electrons) are filled with electrons, we use the **Pauli exclusion principle,** another quantum-mechanical
concept, which stipulates that each electron state can hold no more than two electrons that must have opposite spins.

Thus, *s*, *p*, *d*, and *f* subshells may each accommodate, respectively, a total of 2, 6, 10, and 14 electrons; For most atoms, the electrons fill up the lowest possible energy states in the electron shells and subshells, two electrons (having opposite spins) per state.

The energy structure for a sodium atom is represented schematically.

When all the electrons occupy the lowest possible energies in accord with the foregoing restrictions, an atom is said to be in its **ground state.**



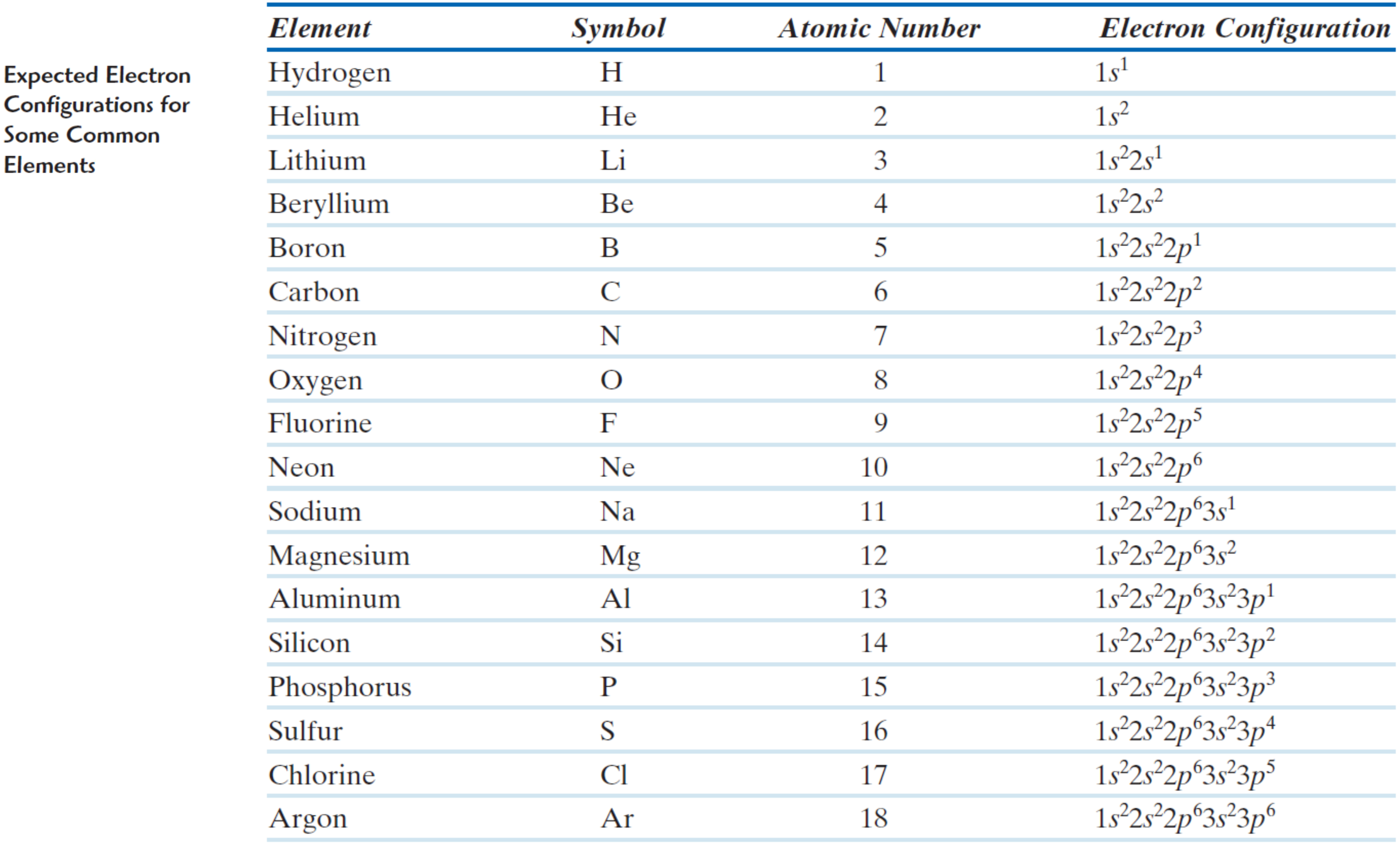
ELECTRONS IN ATOMS

The **electron configuration** or structure of an atom represents the manner in which these states are occupied.

In the conventional notation, the number of electrons in each subshell is indicated by a superscript after the shell–subshell designation.

the electron configurations for hydrogen, helium, and sodium are, respectively, $1s^1$, $1s^2$, and $1s^2 2s^2 2p^6 3s^1$. Electron configurations for some of the more common elements are listed in the following slides.

ELECTRONS IN ATOMS

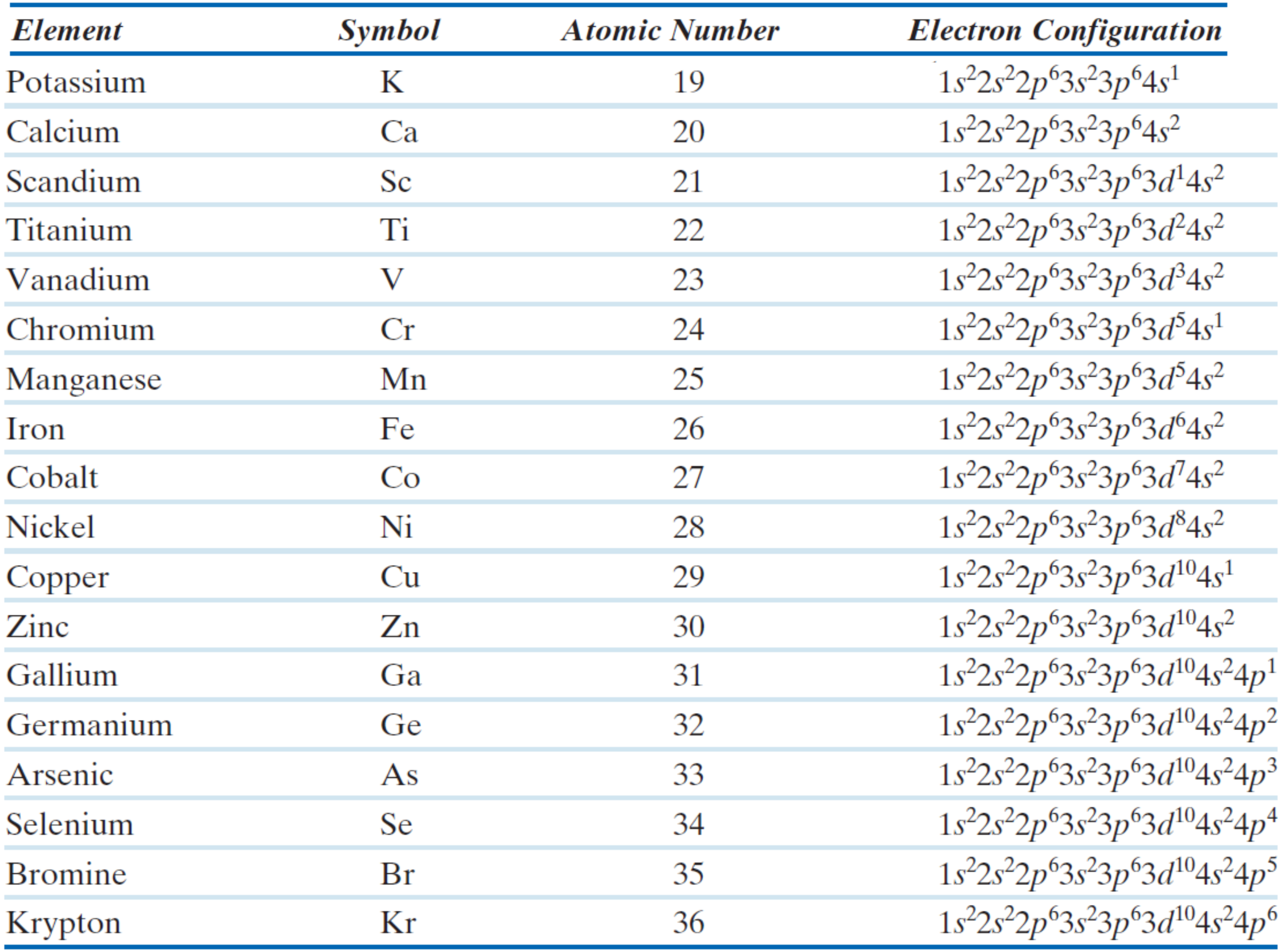


Expected Electron Configurations for Some Common Elements

Element	*Symbol*	*Atomic Number*	*Electron Configuration*
Hydrogen	H	1	$1s^1$
Helium	He	2	$1s^2$
Lithium	Li	3	$1s^2 2s^1$
Beryllium	Be	4	$1s^2 2s^2$
Boron	B	5	$1s^2 2s^2 2p^1$
Carbon	C	6	$1s^2 2s^2 2p^2$
Nitrogen	N	7	$1s^2 2s^2 2p^3$
Oxygen	O	8	$1s^2 2s^2 2p^4$
Fluorine	F	9	$1s^2 2s^2 2p^5$
Neon	Ne	10	$1s^2 2s^2 2p^6$
Sodium	Na	11	$1s^2 2s^2 2p^6 3s^1$
Magnesium	Mg	12	$1s^2 2s^2 2p^6 3s^2$
Aluminum	Al	13	$1s^2 2s^2 2p^6 3s^2 3p^1$
Silicon	Si	14	$1s^2 2s^2 2p^6 3s^2 3p^2$
Phosphorus	P	15	$1s^2 2s^2 2p^6 3s^2 3p^3$
Sulfur	S	16	$1s^2 2s^2 2p^6 3s^2 3p^4$
Chlorine	Cl	17	$1s^2 2s^2 2p^6 3s^2 3p^5$
Argon	Ar	18	$1s^2 2s^2 2p^6 3s^2 3p^6$

ELECTRONS IN ATOMS

Expected Electron Configurations for Some Common Elements

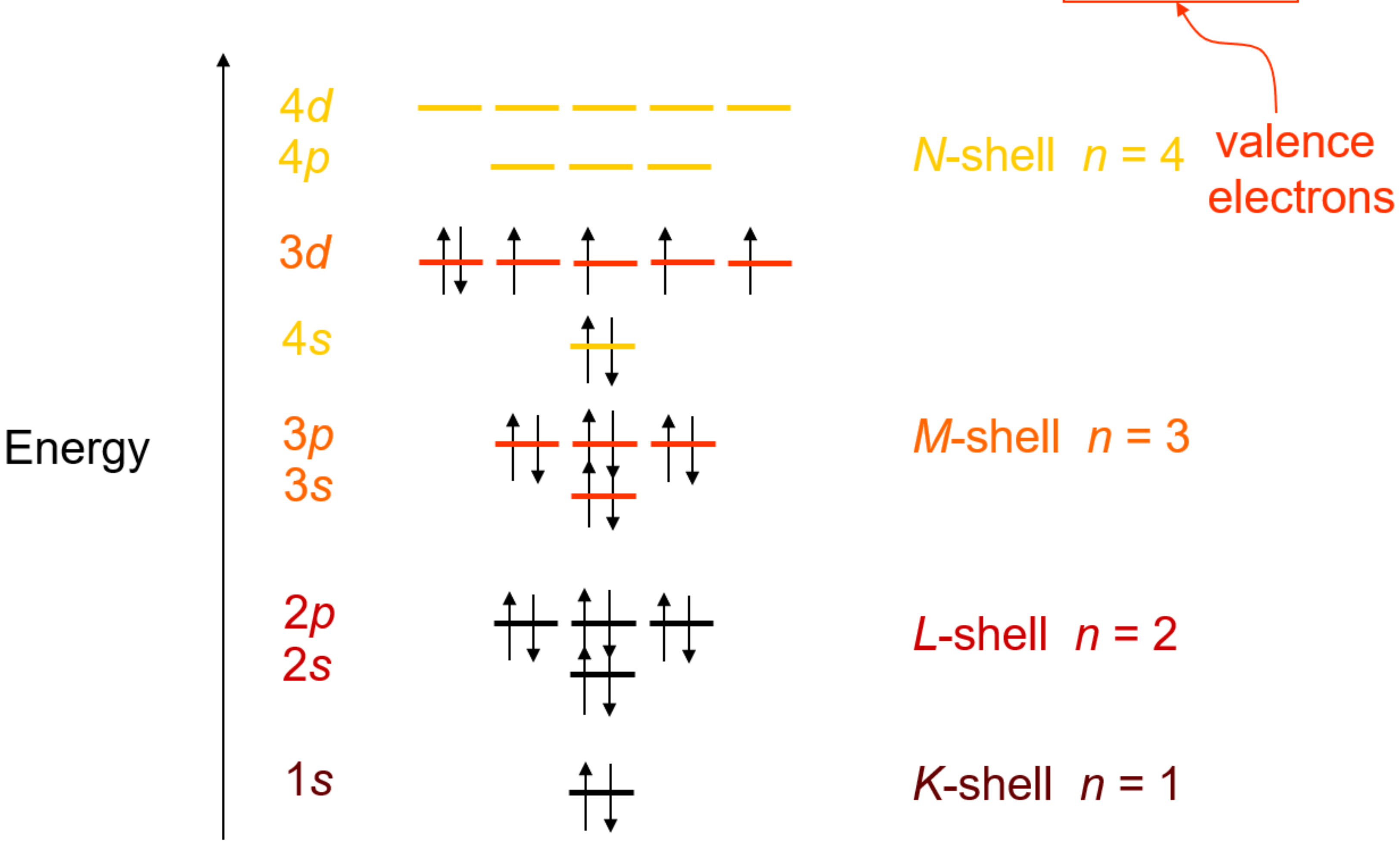


Element	*Symbol*	*Atomic Number*	*Electron Configuration*
Potassium	K	19	$1s^22s^22p^63s^23p^64s^1$
Calcium	Ca	20	$1s^22s^22p^63s^23p^64s^2$
Scandium	Sc	21	$1s^22s^22p^63s^23p^63d^14s^2$
Titanium	Ti	22	$1s^22s^22p^63s^23p^63d^24s^2$
Vanadium	V	23	$1s^22s^22p^63s^23p^63d^34s^2$
Chromium	Cr	24	$1s^22s^22p^63s^23p^63d^54s^1$
Manganese	Mn	25	$1s^22s^22p^63s^23p^63d^54s^2$
Iron	Fe	26	$1s^22s^22p^63s^23p^63d^64s^2$
Cobalt	Co	27	$1s^22s^22p^63s^23p^63d^74s^2$
Nickel	Ni	28	$1s^22s^22p^63s^23p^63d^84s^2$
Copper	Cu	29	$1s^22s^22p^63s^23p^63d^{10}4s^1$
Zinc	Zn	30	$1s^22s^22p^63s^23p^63d^{10}4s^2$
Gallium	Ga	31	$1s^22s^22p^63s^23p^63d^{10}4s^24p^1$
Germanium	Ge	32	$1s^22s^22p^63s^23p^63d^{10}4s^24p^2$
Arsenic	As	33	$1s^22s^22p^63s^23p^63d^{10}4s^24p^3$
Selenium	Se	34	$1s^22s^22p^63s^23p^63d^{10}4s^24p^4$
Bromine	Br	35	$1s^22s^22p^63s^23p^63d^{10}4s^24p^5$
Krypton	Kr	36	$1s^22s^22p^63s^23p^63d^{10}4s^24p^6$

ELECTRONS IN ATOMS

Iron	Fe	26	$1s^22s^22p^63s^23p^63d^64s^2$

Fe - atomic # = 26 $1s^2$ $2s^2\,2p^6$ $3s^2\,3p^6$ $\boxed{3d^6\ \ 4s^2}$



ELECTRONS IN ATOMS

At this point, comments regarding these electron configurations are necessary.

First, the valence electrons are those that occupy the outermost shell. These electrons are extremely important; as will be seen, they participate in the bonding between atoms to form atomic and molecular aggregates.

Furthermore, many of the physical and chemical properties of solids are based on these valence electrons.

In addition, some atoms have what are termed ***stable electron configurations***—that is, the states within the outermost or valence electron shell are completely filled.

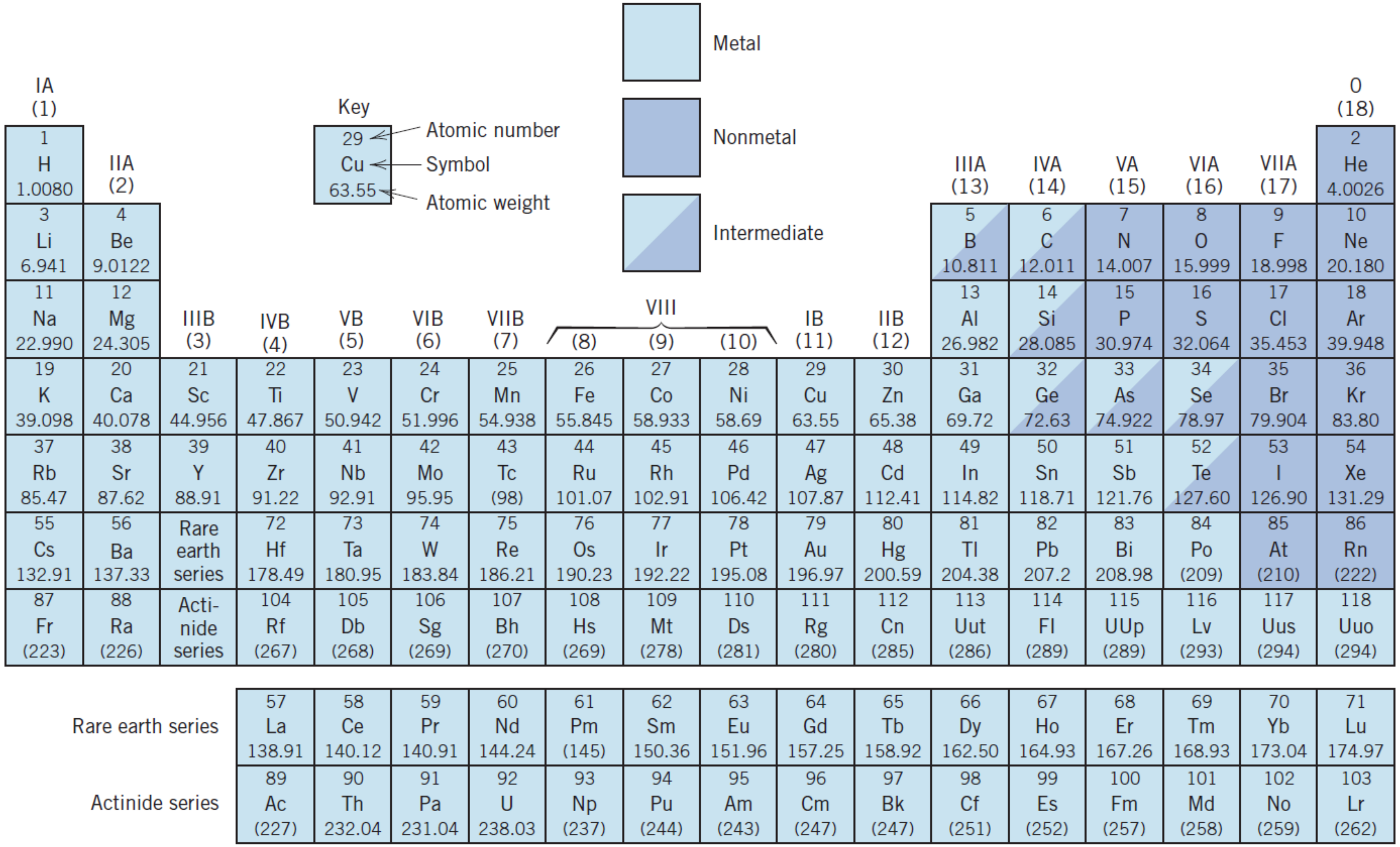
Normally, this corresponds to the occupation of just the *s* and *p* states for the outermost shell by a total of eight electrons, as in neon, argon, and krypton; one exception is helium, which contains only two 1*s* electrons.

These elements (Ne, Ar, Kr, and He) are the inert, or noble, gases, which are virtually unreactive chemically.

Some atoms of the elements that have unfilled valence shells assume stable electron configurations by gaining or losing electrons to form charged ions or by sharing electrons with other atoms.

THE PERIODIC TABLE

All the elements have been classified according to electron configuration in the **periodic table.** Here, the elements are situated, with increasing atomic number, in seven horizontal rows called *periods*.



Key

29 ← Atomic number
Cu ← Symbol
63.55 ← Atomic weight

Metal

Nonmetal

Intermediate

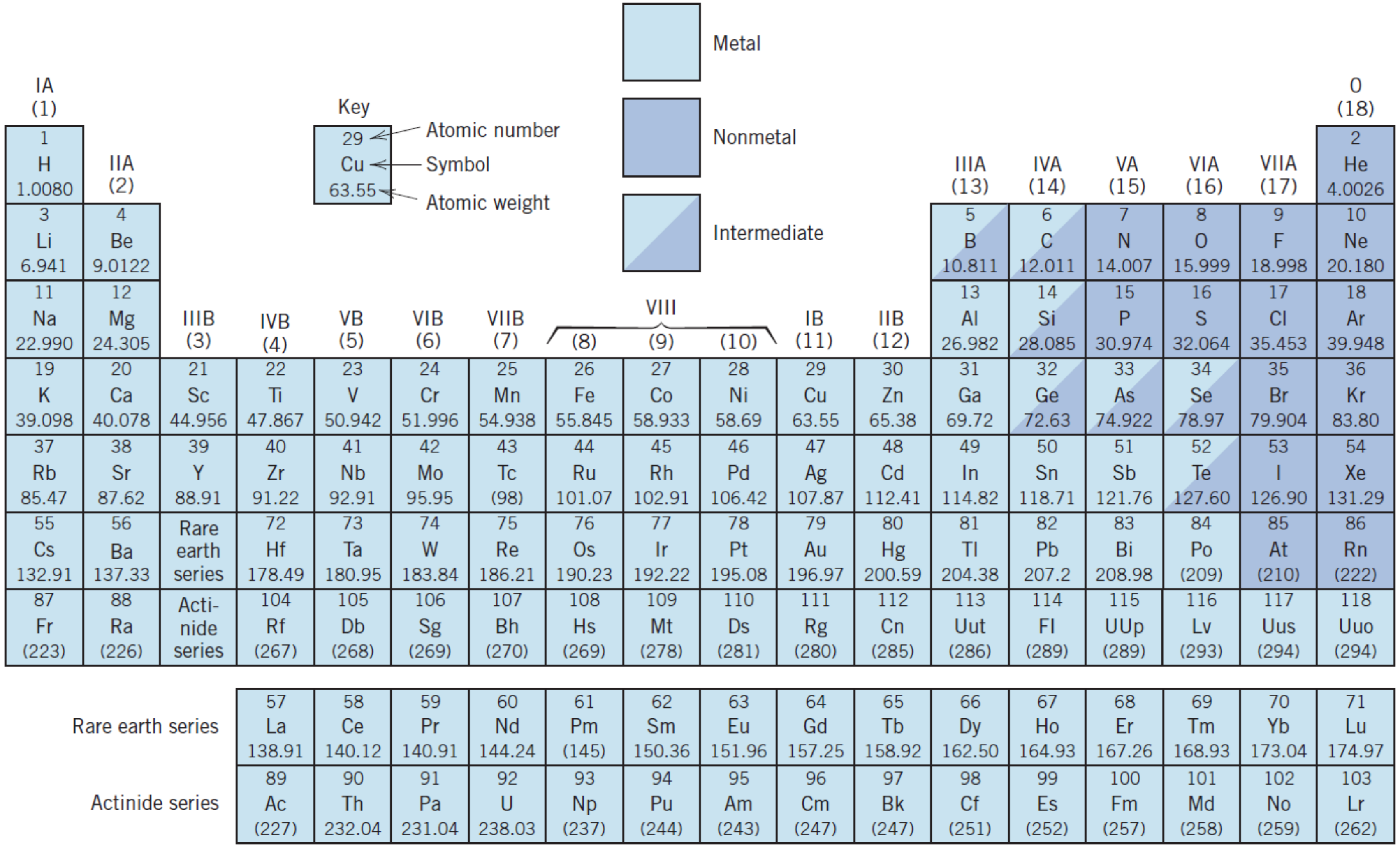
IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	(9)	(10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
1 H 1.0080																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.085	15 P 30.974	16 S 32.064	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.922	34 Se 78.97	35 Br 79.904	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	Rare earth series	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	Acti-nide series	104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (269)	109 Mt (278)	110 Ds (281)	111 Rg (280)	112 Cn (285)	113 Uut (286)	114 Fl (289)	115 UUp (289)	116 Lv (293)	117 Uus (294)	118 Uuo (294)

Rare earth series	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.92	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
Actinide series	89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

THE PERIODIC TABLE

The arrangement is such that all elements arrayed in a given column or group have similar valence electron structures, as well as chemical and physical properties.

These properties change gradually, moving horizontally across each period and vertically down each column.



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Nonmetal

Intermediate

IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	VIII (9)	VIII (10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
1 H 1.0080																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.085	15 P 30.974	16 S 32.064	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.922	34 Se 78.97	35 Br 79.904	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	Rare earth series	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	Acti-nide series	104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (269)	109 Mt (278)	110 Ds (281)	111 Rg (280)	112 Cn (285)	113 Uut (286)	114 Fl (289)	115 UUp (289)	116 Lv (293)	117 Uus (294)	118 Uuo (294)

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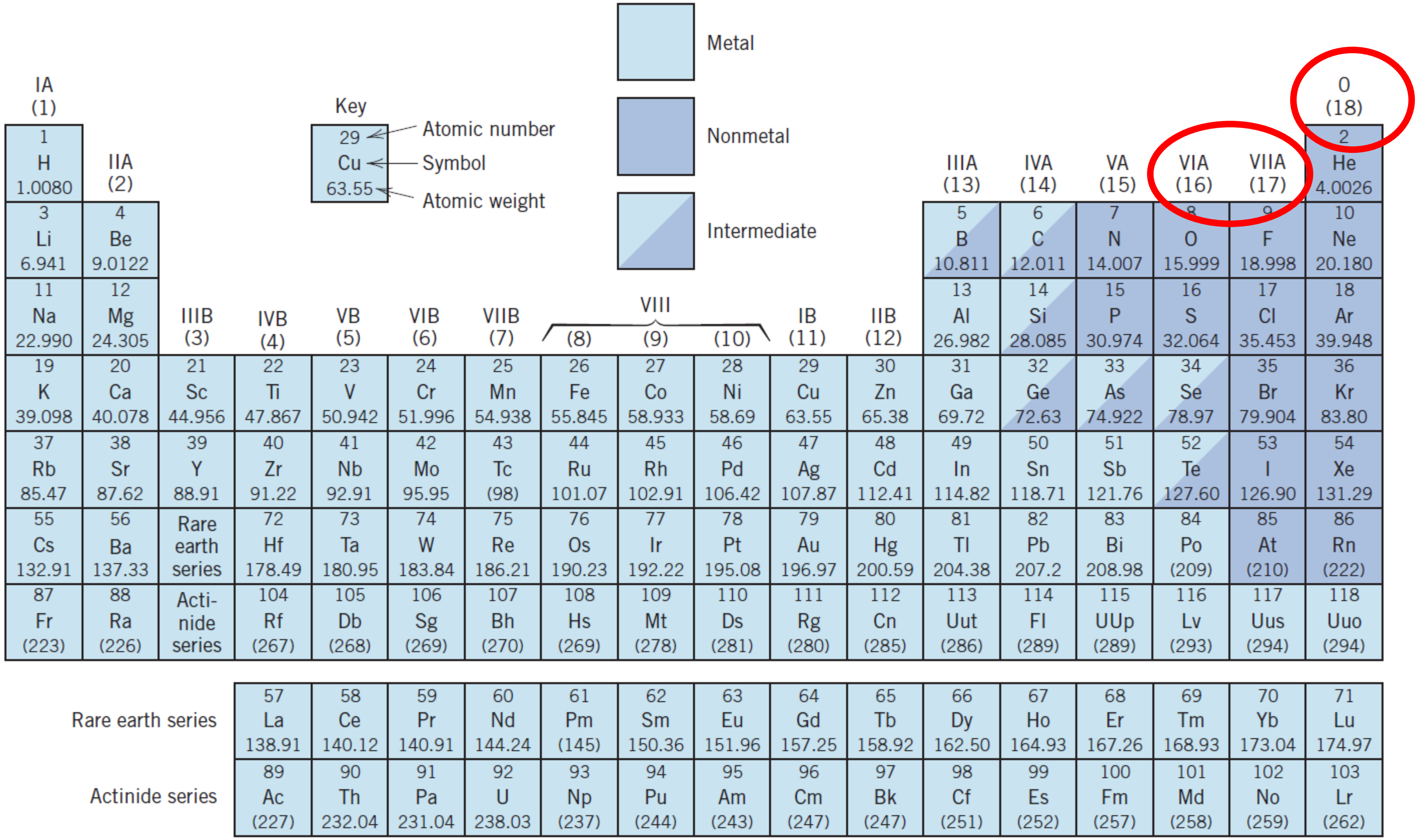
Actinide series	89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

THE PERIODIC TABLE

The elements positioned in Group 0, the rightmost group, are the ***inert gases***, which have filled electron shells and stable electron configurations.

Group VIIA and VIA elements are one and two electrons deficient, respectively, from having stable structures.



Key

29 ← Atomic number
Cu ← Symbol
63.55 ← Atomic weight

🟦 Metal
🟦 Nonmetal
◩ Intermediate

IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	(9)	(10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
1 H 1.0080																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
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37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
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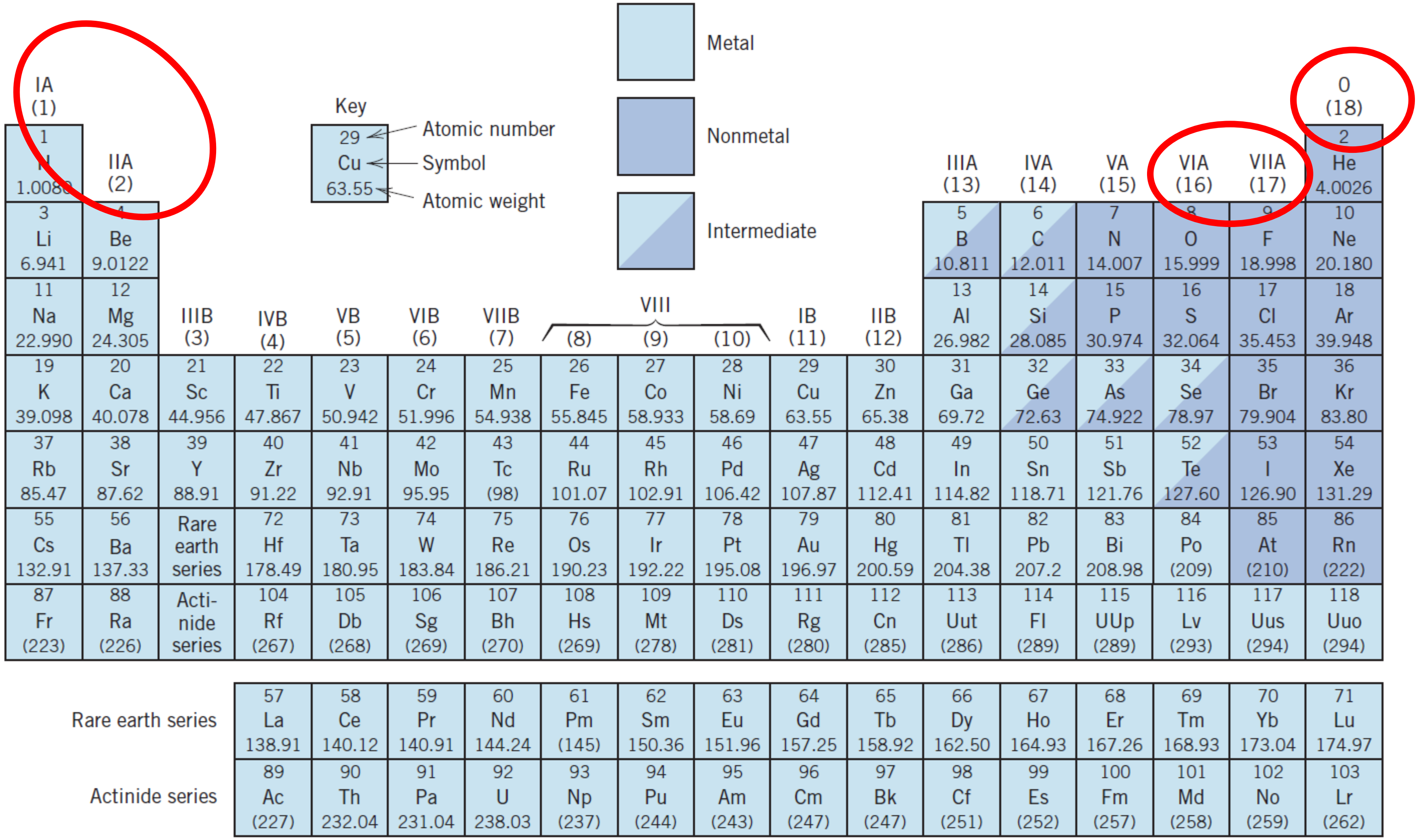
Actinide series	89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

THE PERIODIC TABLE

The Group VIIA elements (F, Cl, Br, I, and At) are sometimes termed the ***halogens***.

The alkali and the alkaline earth metals (Li, Na, K, Be, Mg, Ca, etc.) are labelled as Groups IA and IIA, having, respectively, one and two electrons in excess of stable structures.



Key

29 – Atomic number

Cu – Symbol

63.55 – Atomic weight

☐ Metal

▨ Nonmetal

◪ Intermediate

IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	VIII (9)	VIII (10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
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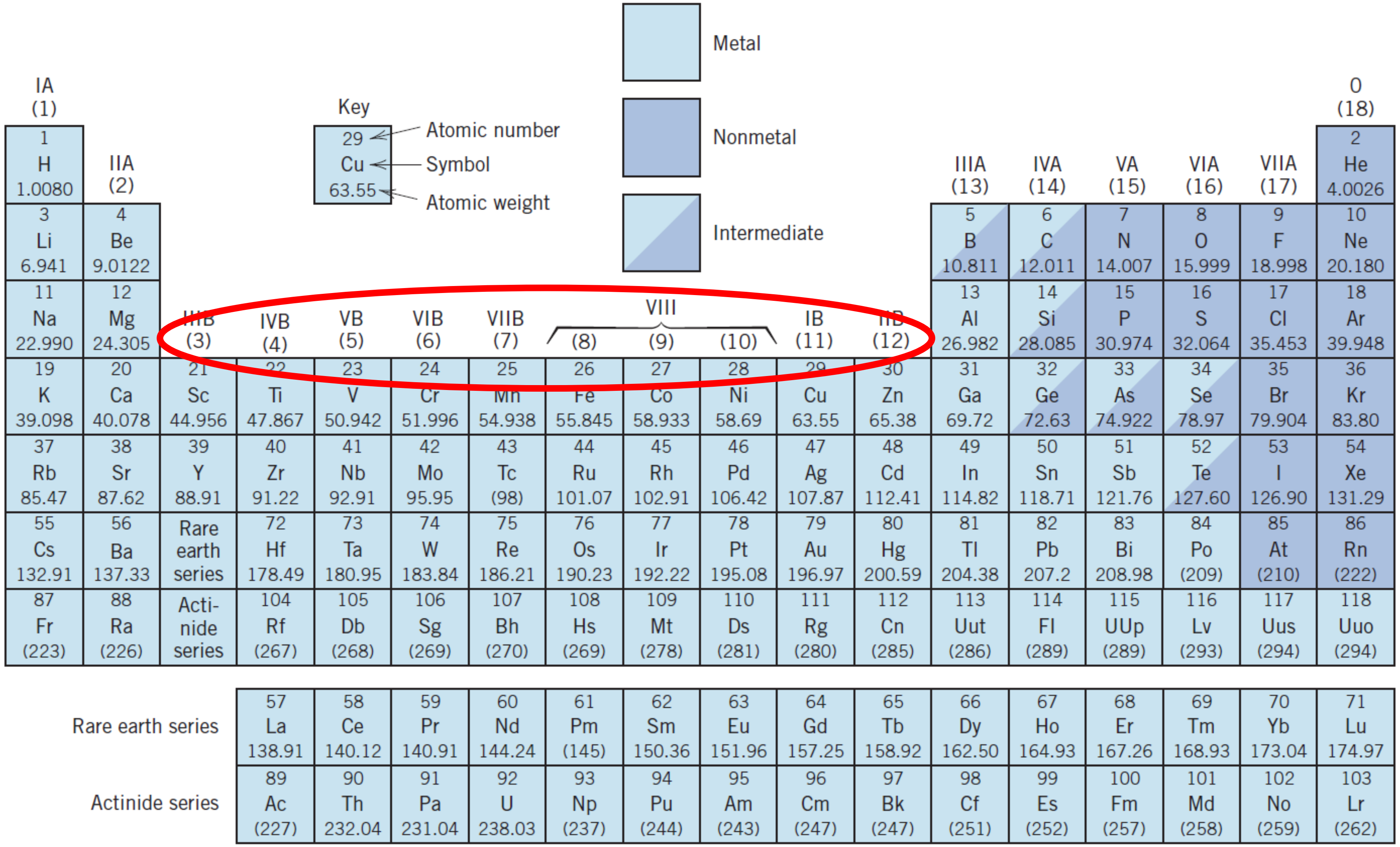
THE PERIODIC TABLE

The elements in the three long periods, Groups IIIB through IIB, are termed the ***transition metals***, which have **partially filled *d* electron** states and in some cases one or two electrons in the next higher energy shell.



Key

29 Atomic number

Cu Symbol

63.55 Atomic weight

Metal

Nonmetal

Intermediate

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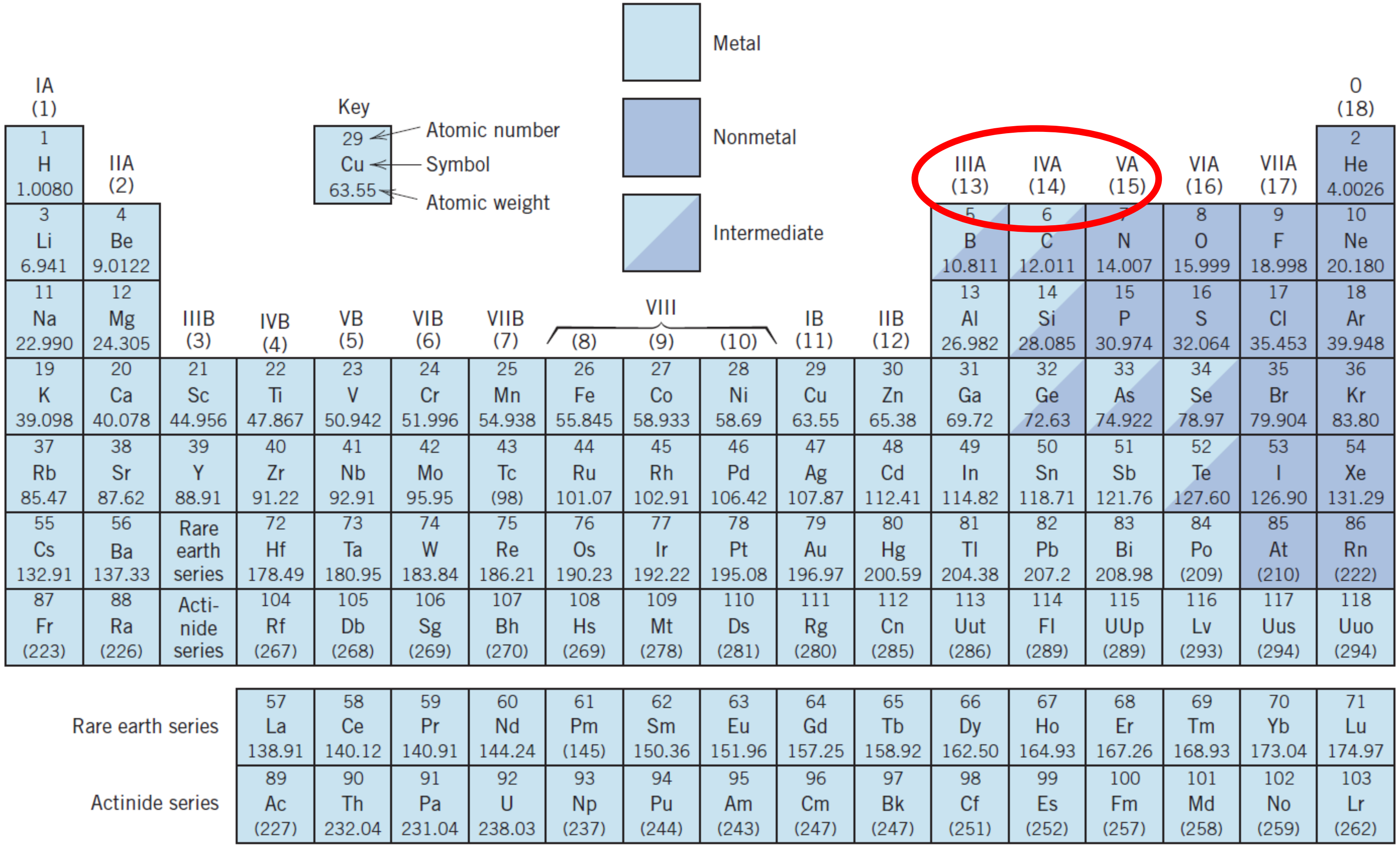
THE PERIODIC TABLE

Groups IIIA, IVA, and VA (B, Si, Ge, As, etc.) display characteristics that are intermediate between the metals and nonmetals by virtue of their valence electron structures.



Key: 29 – Atomic number; Cu – Symbol; 63.55 – Atomic weight

Metal

Nonmetal

Intermediate

IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	VIII (9)	VIII (10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
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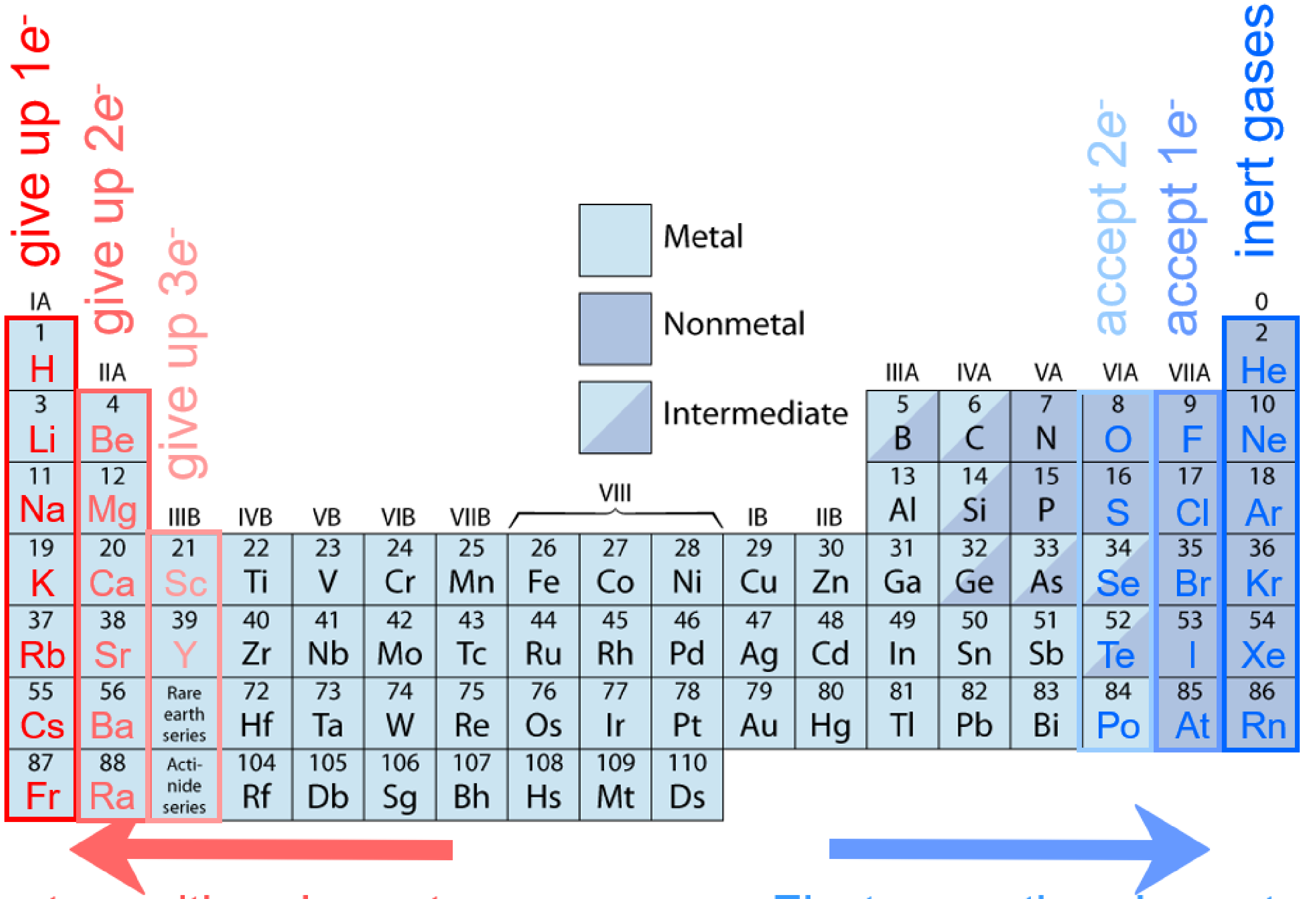
THE PERIODIC TABLE

As a general rule, **electronegativity increases in moving from left to right and from bottom to top.** Atoms are more likely to accept electrons if their outer shells are almost full and if they are less “shielded” from (i.e., closer to) the nucleus.



IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	Rare earth series	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	Acti-nide series	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds								

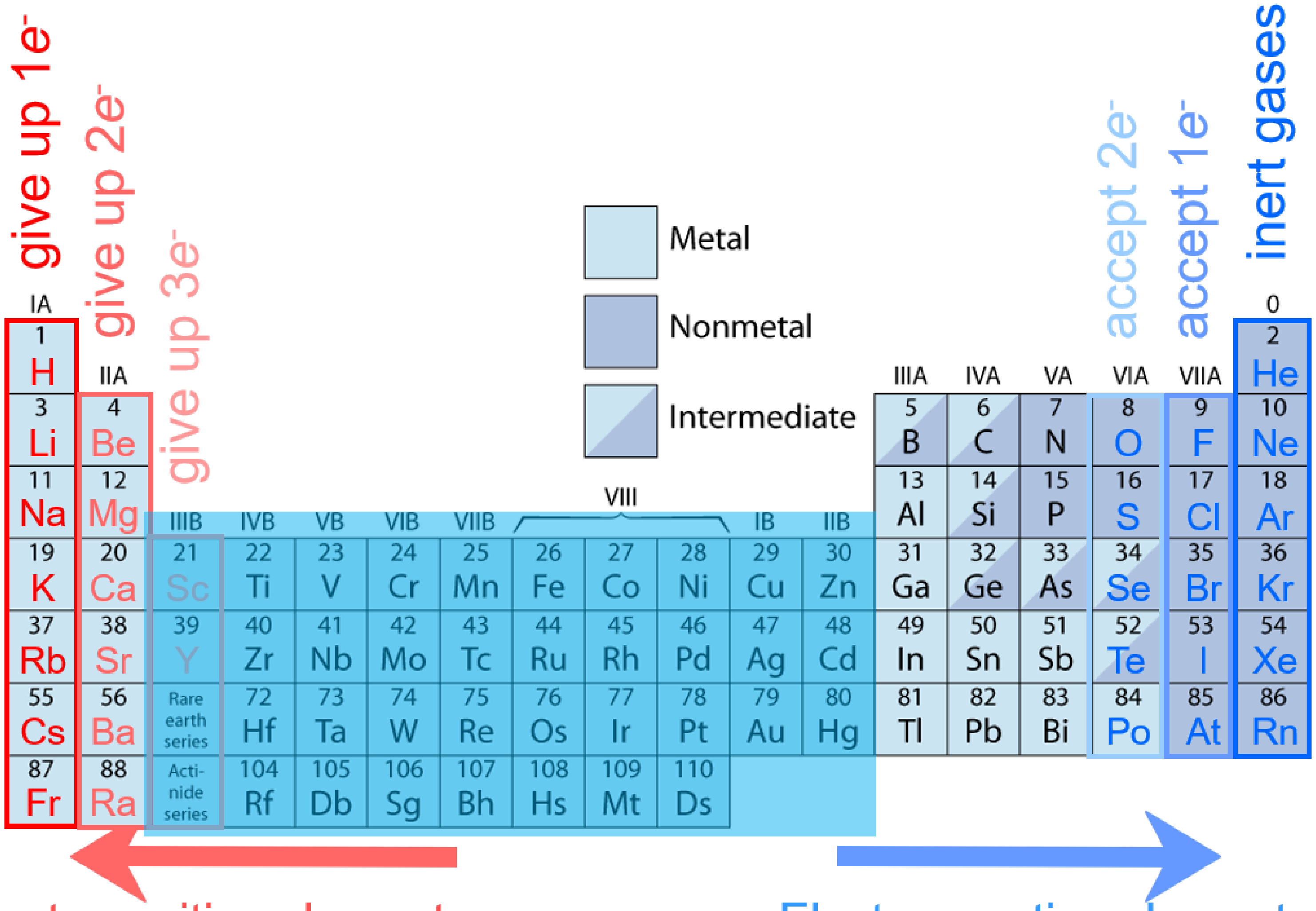
▢	Metal
▨	Nonmetal
◩	Intermediate

Electropositive elements: Readily give up electrons to become + ions.

Electronegative elements: Readily acquire electrons to become - ions.

THE PERIODIC TABLE

In addition to chemical behaviour, physical properties of the elements also tend to vary systematically with position in the periodic table. For example, most metals that reside in the centre of the table (Groups IIIB through IIB) are relatively good conductors of electricity and heat; nonmetals are typically electrical and thermal insulators.



IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	Rare earth series	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	Actinide series	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds								

Electropositive elements:
Readily give up electrons
to become + ions.

Electronegative elements:
Readily acquire electrons
to become - ions.

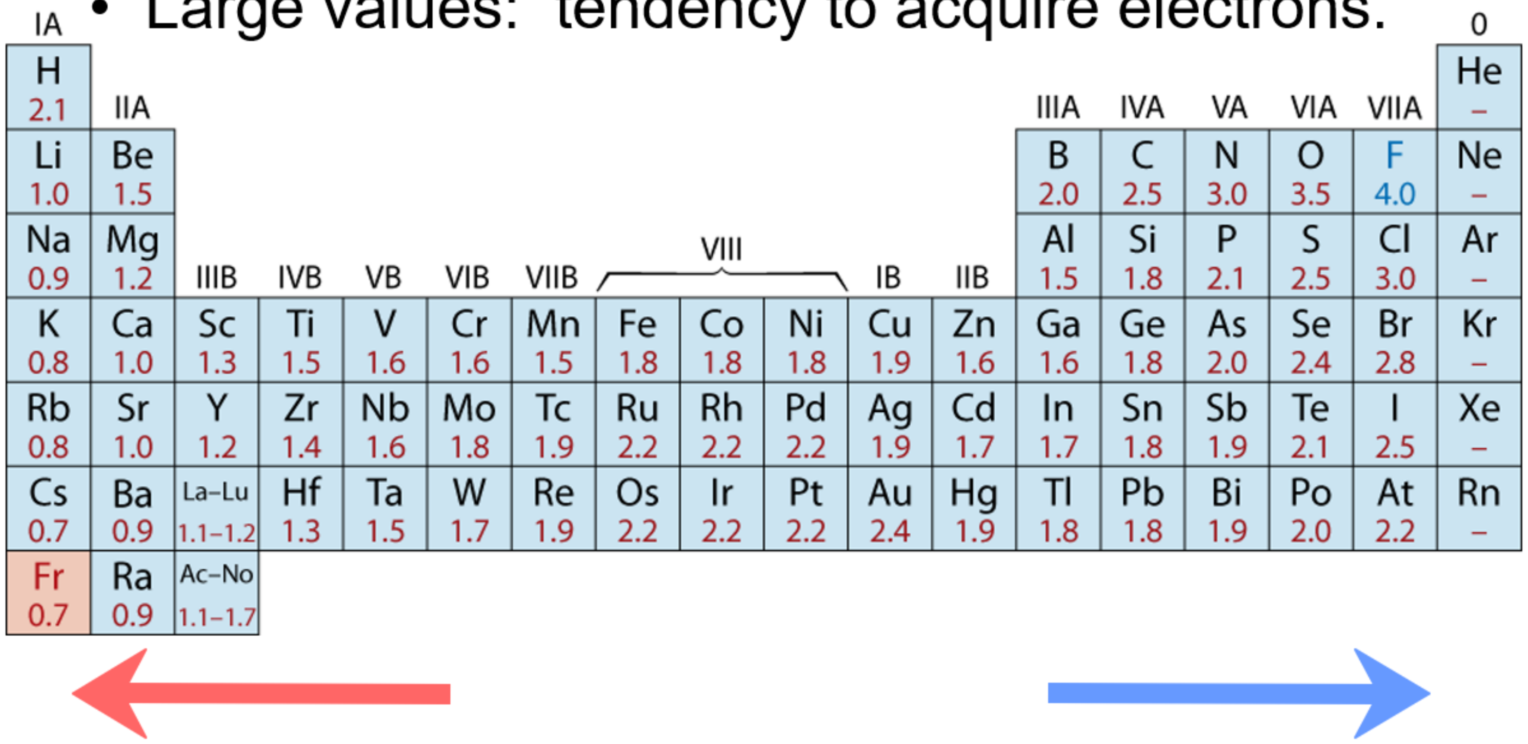
THE PERIODIC TABLE

As may be noted from the periodic table, most of the elements really come under the metal classification.

These are sometimes termed **electropositive** elements, indicating that they are capable of giving up their few valence electrons to become positively charged ions.

The electronegativity values for the elements

- Ranges from 0.7 to 4.0,
- Large values: tendency to acquire electrons.



IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
H 2.1																	He –
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne –
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar –
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr –
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe –
Cs 0.7	Ba 0.9	La–Lu 1.1–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn –
Fr 0.7	Ra 0.9	Ac–No 1.1–1.7															

Smaller electronegativity

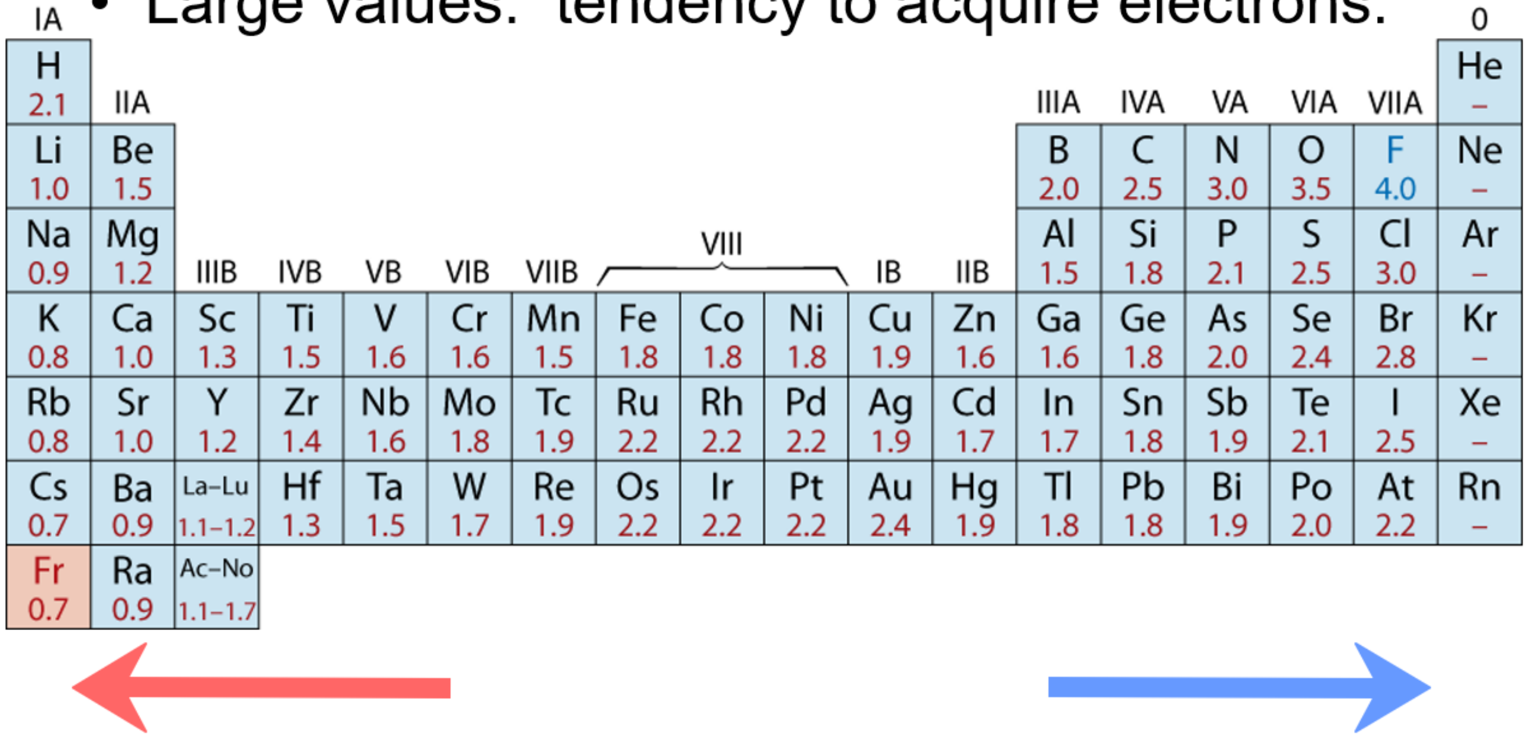
Larger electronegativity

THE PERIODIC TABLE

Furthermore, the elements situated on the right side of the table are **electronegative**—that is, they readily accept electrons to form negatively charged ions, or sometimes they share electrons with other atoms.

The electronegativity values for the elements

- Ranges from 0.7 to 4.0,
- Large values: tendency to acquire electrons.



IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
H 2.1																	He –
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne –
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar –
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr –
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe –
Cs 0.7	Ba 0.9	La–Lu 1.1–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn –
Fr 0.7	Ra 0.9	Ac–No 1.1–1.7															

Smaller electronegativity

Larger electronegativity

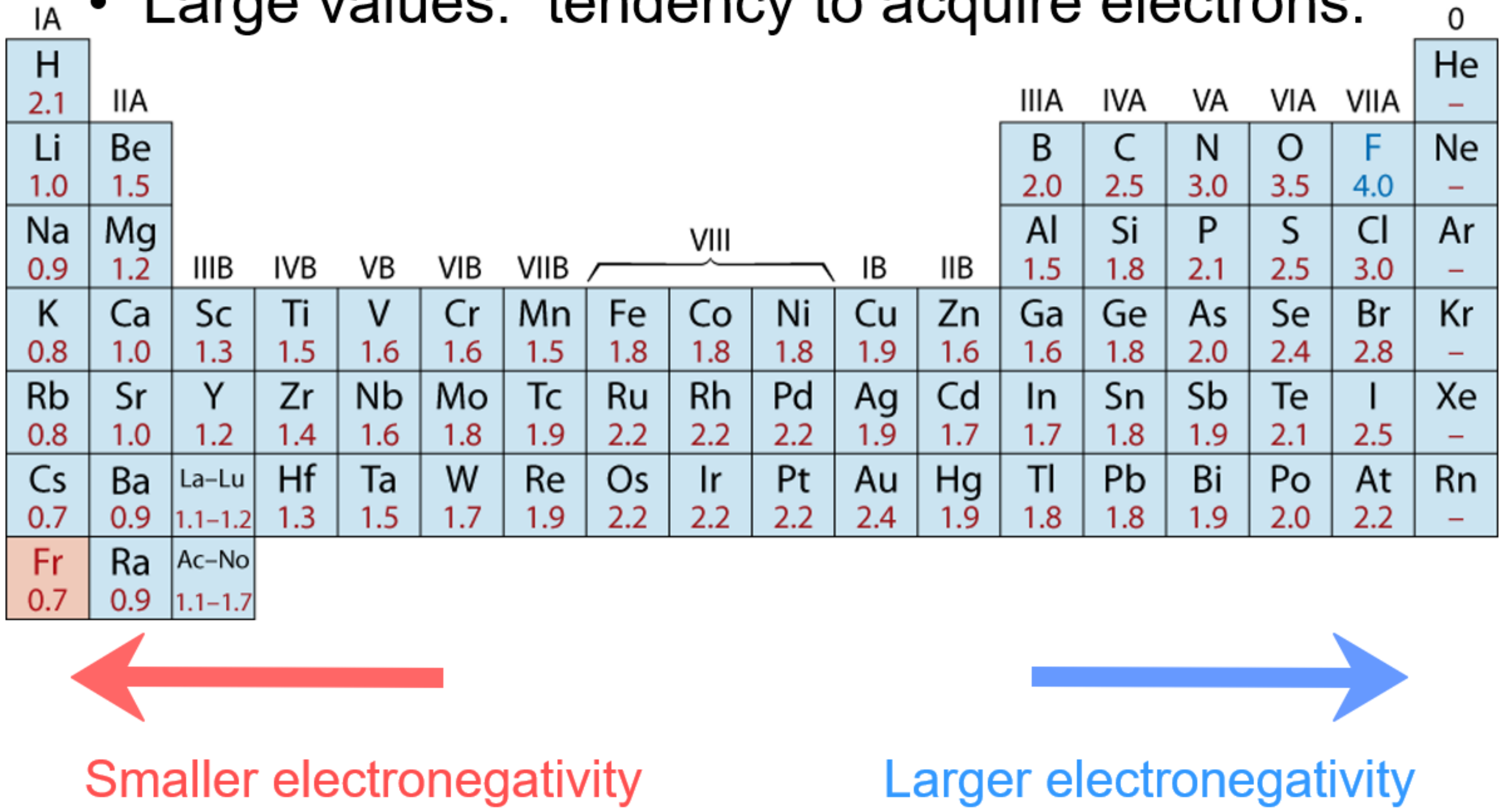
THE PERIODIC TABLE

Mechanically, the metallic elements exhibit varying degrees of *ductility*—the ability to be plastically deformed without fracturing (e.g., the ability to be rolled into thin sheets).

Most of the nonmetals are either gases or liquids, or in the solid state are brittle in nature.

The electronegativity values for the elements

- Ranges from 0.7 to 4.0,
- Large values: tendency to acquire electrons.



IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
H 2.1																	He –
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne –
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar –
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr –
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe –
Cs 0.7	Ba 0.9	La–Lu 1.1–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn –
Fr 0.7	Ra 0.9	Ac–No 1.1–1.7															

Smaller electronegativity

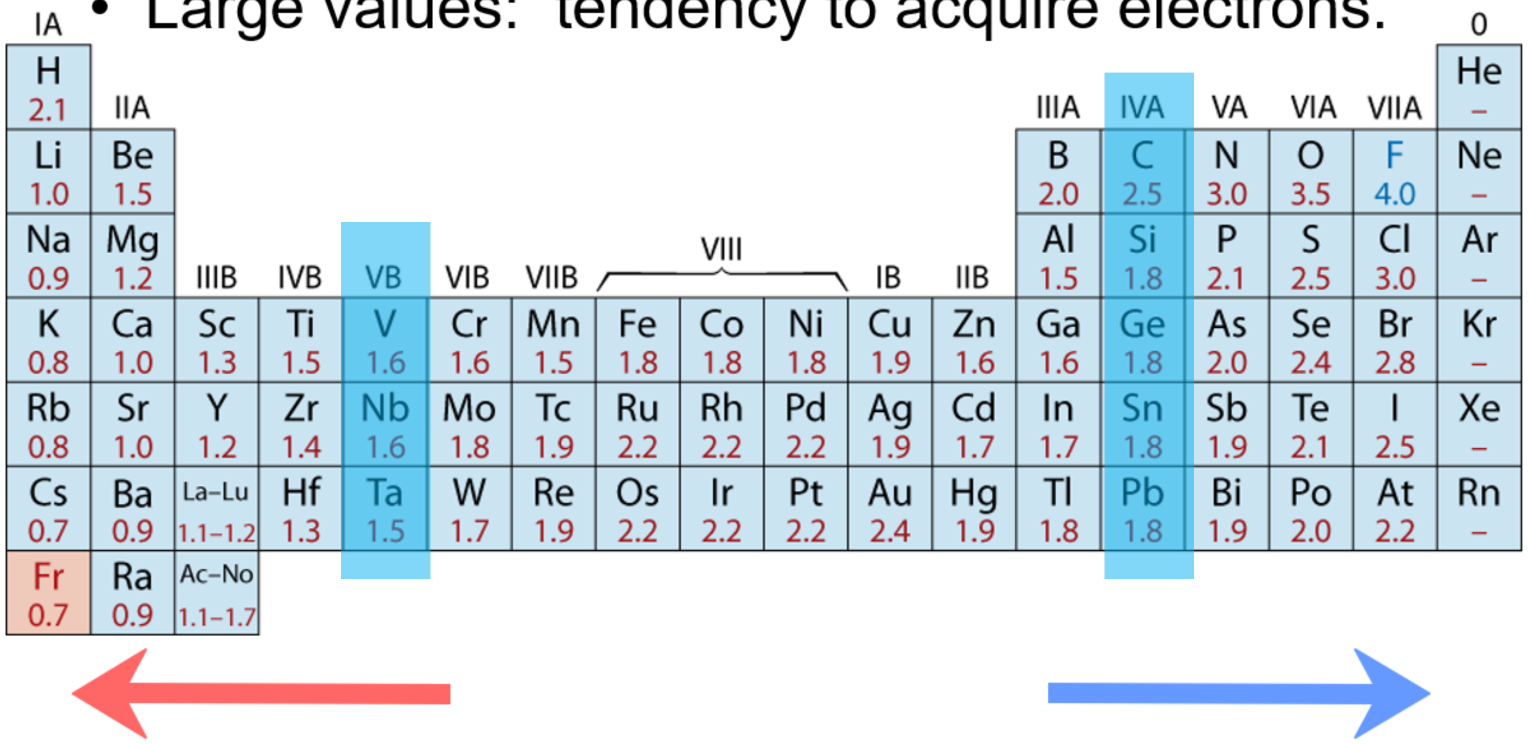
Larger electronegativity

THE PERIODIC TABLE

Furthermore, for the Group IVA elements [C (diamond), Si, Ge, Sn, and Pb], electrical conductivity increases as we move down this column. The Group VB metals (V, Nb, and Ta) have very high melting temperatures, which increase in going down this column.

The electronegativity values for the elements

- Ranges from 0.7 to 4.0,
- Large values: tendency to acquire electrons.



IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII	VIII	VIII	IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
H 2.1																	He –
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne –
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar –
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr –
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe –
Cs 0.7	Ba 0.9	La–Lu 1.1–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn –
Fr 0.7	Ra 0.9	Ac–No 1.1–1.7															

Smaller electronegativity

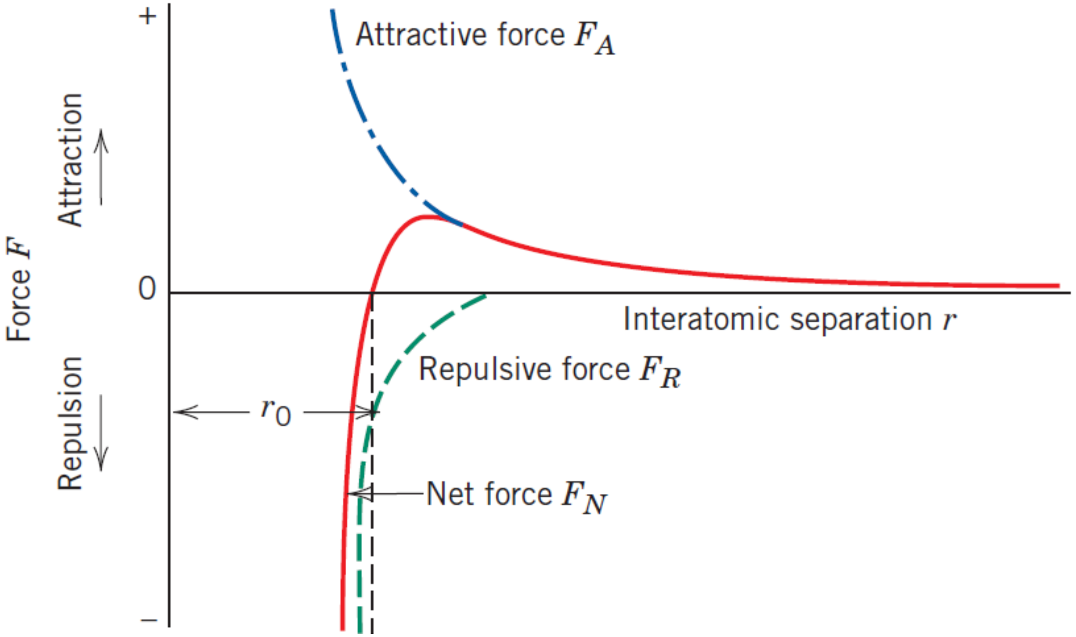
Larger electronegativity

Atomic Bonding in Solids

knowledge of the interatomic forces that bind the atoms together. The principles of atomic bonding is illustrated by considering how two isolated atoms interact as they are brought close together from an infinite separation.

At large distances, interactions are negligible because the atoms are too far apart to have an influence on each other; however, at small separation distances, each atom exerts forces on the others.

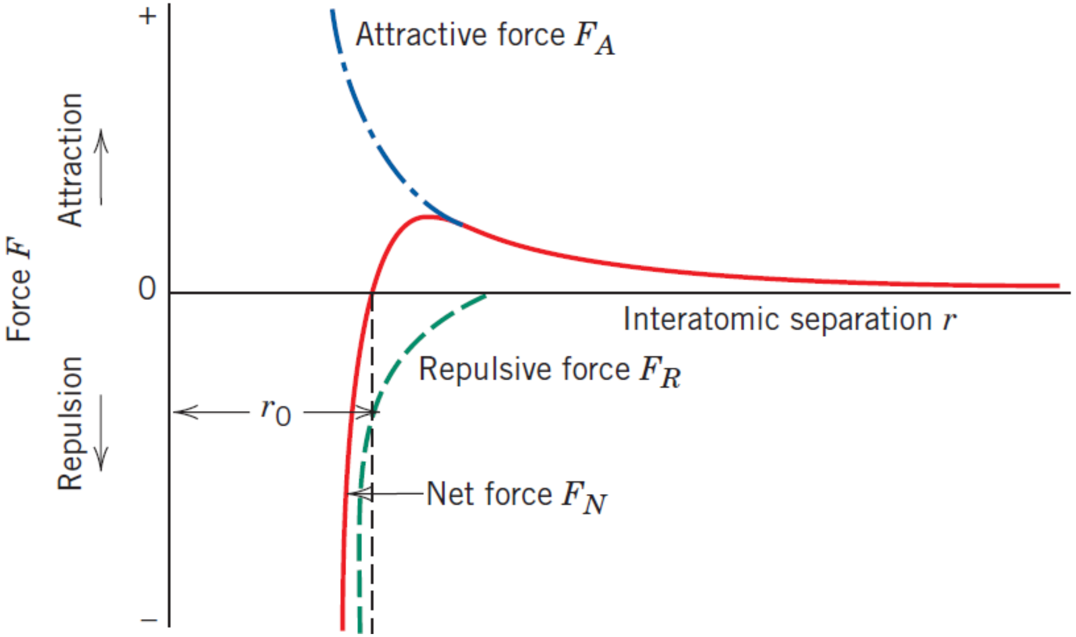
These forces are of two types, attractive (F_A) and repulsive (F_R), and the magnitude of each depends on the separation or interatomic distance (r);



Atomic Bonding in Solids

The figure is a schematic plot of F_A and F_R versus r. The origin of an attractive force F_A depends on the particular type of bonding that exists between the two atoms.

Repulsive forces arise from interactions between t negatively charged electron clouds for the two atoms and are important only at small values of r a the outer electron shells of the two atoms begin to overlap.

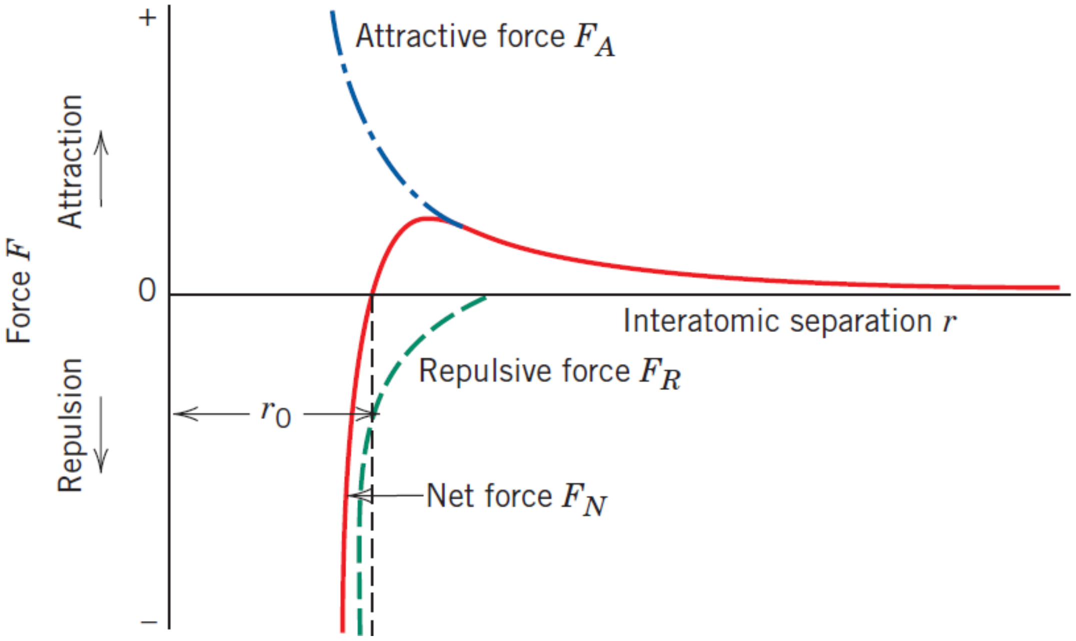


Atomic Bonding in Solids

The net force F_N between the two atoms is just the sum of both attractive and repulsive components; that is, $F_N = F_A + F_R$

which is also a function of the interatomic separation, as also plotted in the figure.

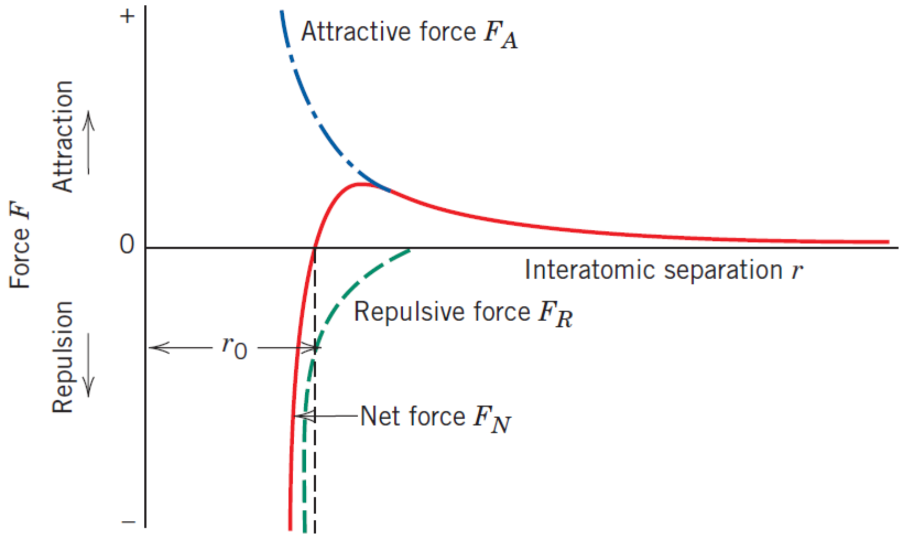
When F_A and F_R are equal in magnitude but opposite in sign, there is no net force, that is, $F_A + F_R = 0$



and a state of equilibrium exists. The centres of the two atoms remain separated by the equilibrium spacing r_0, as indicated in the figure. For many atoms, r_0 is approximately 0.3 nm.

Atomic Bonding in Solids

Once in this position, any attempt to move the two atoms farther apart is counteracted by the attractive force, while pushing them closer together is resisted by the increasing repulsive force.



Sometimes it is more convenient to work with the potential energies between two atoms instead of forces. Mathematically, energy (*E*) and force (*F*) are related as:

$$F = \frac{dE}{dr}$$

$$E = \int F\,dr$$

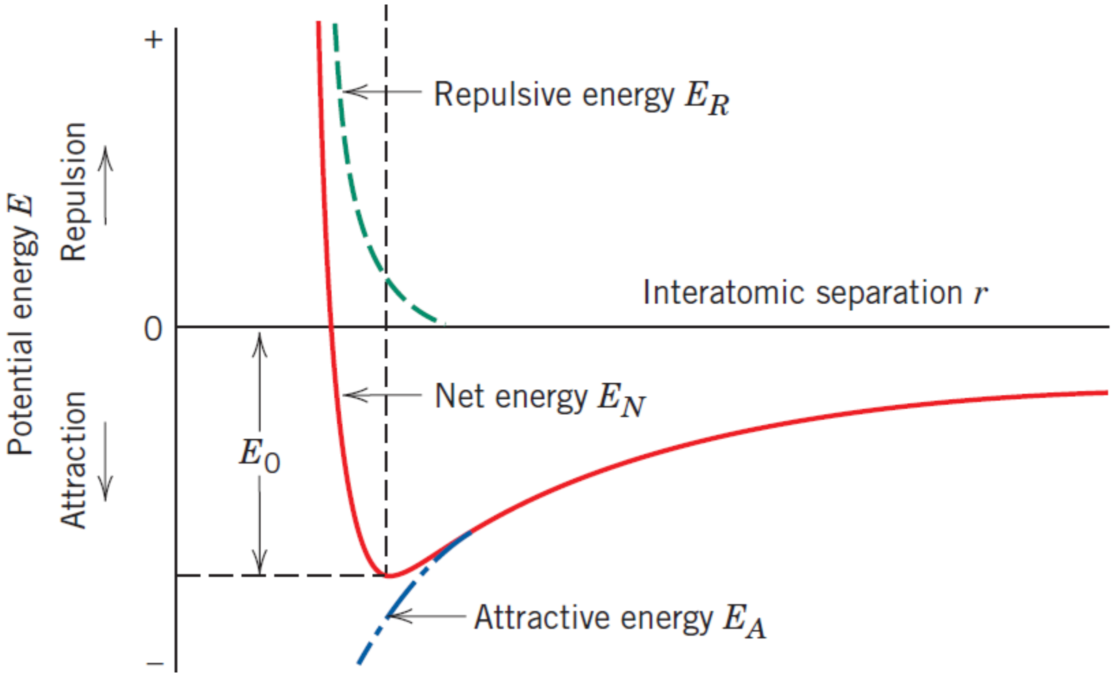
$$E_N = \int_r^{\infty} F_N\,dr \quad = \int_r^{\infty} F_A\,dr + \int_r^{\infty} F_R\,dr \quad = E_A + E_R$$

$$F_N = \frac{dE_A}{dr} + \frac{dE_R}{dr}$$

Atomic Bonding in Solids

Here, the figure plots attractive, repulsive, and net potential energies as a function of interatomic separation for two atoms.

The net curve is the sum of the attractive and repulsive curves. The minimum in the net energy curve corresponds to the equilibrium spacing, r_0. Furthermore, the **bonding energy** for these two atoms, E_0, corresponds to the energy at this minimum point (also shown in Figure); it represents the energy required to separate these two atoms to an infinite separation.

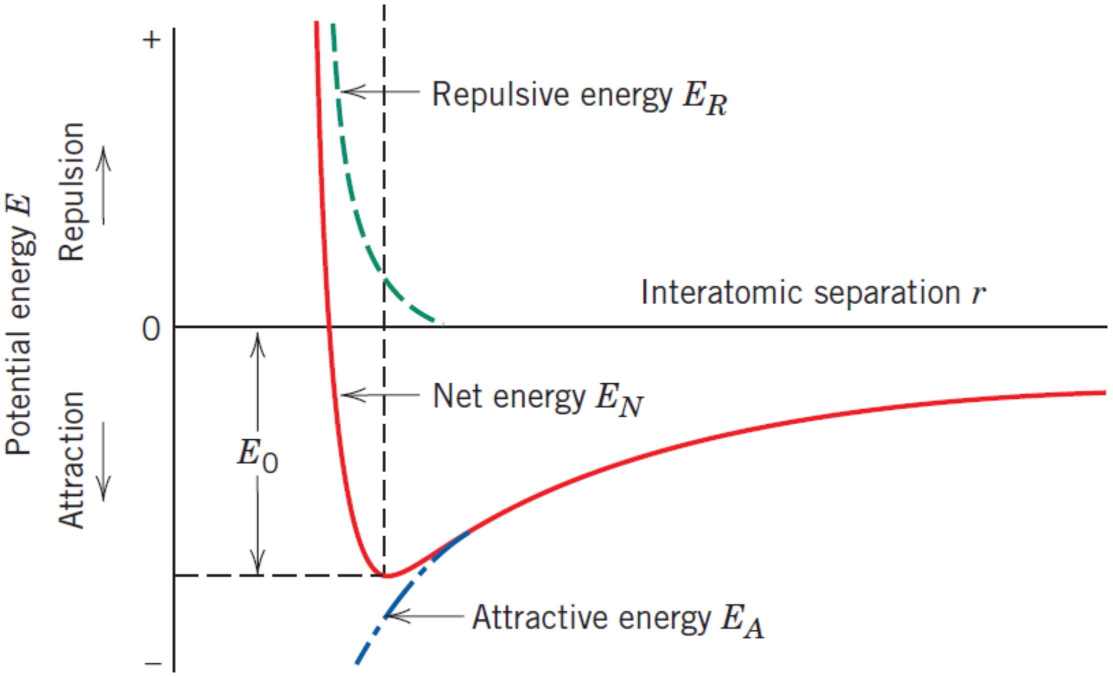


Atomic Bonding in Solids

Although the preceding treatment deals with an ideal situation involving only two atoms, a similar yet more complex condition exists for solid materials because force and energy interactions among atoms must be considered.

Nevertheless, a bonding energy, analogous to E_0 above, may be associated with each atom. The magnitude of this bonding energy and the shape of the energy–versus–interatomic separation curve vary from material to material, and they both depend on the type of atomic bonding.

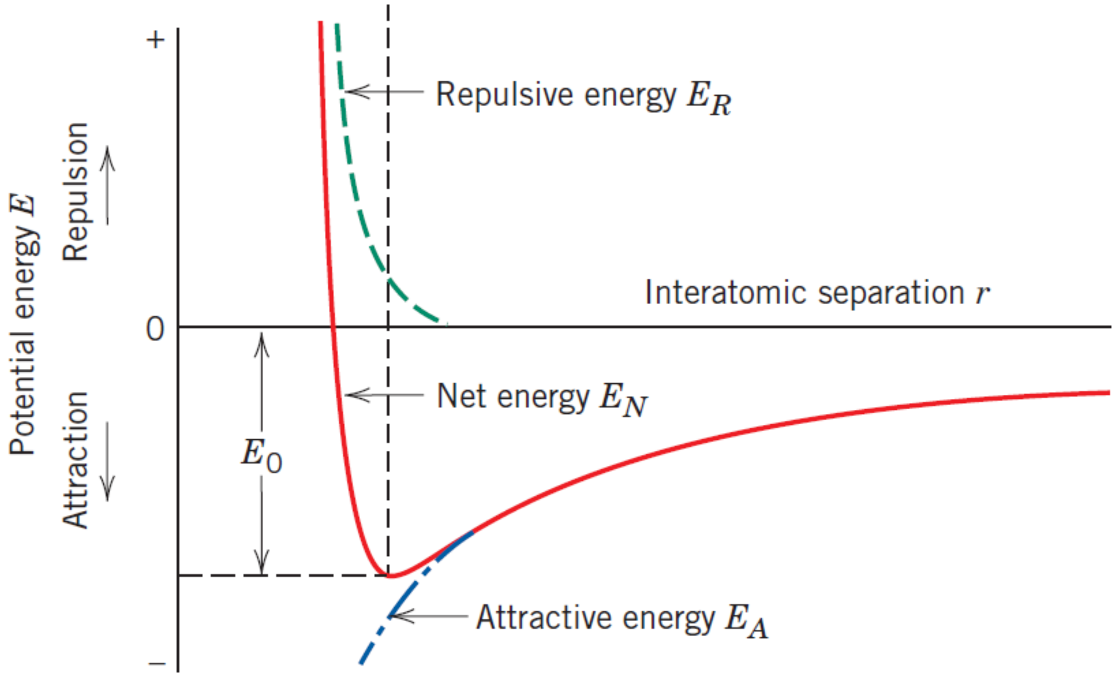
Furthermore, a number of material properties depend on E_0, the curve shape, and bonding type.



Atomic Bonding in Solids

- *Materials having large bonding energies typically also have high melting temperatures; at room temperature, solid substances are formed for large bonding energies, whereas for small energies, the gaseous state is favoured; liquids prevail when the energies are of intermediate magnitude.*

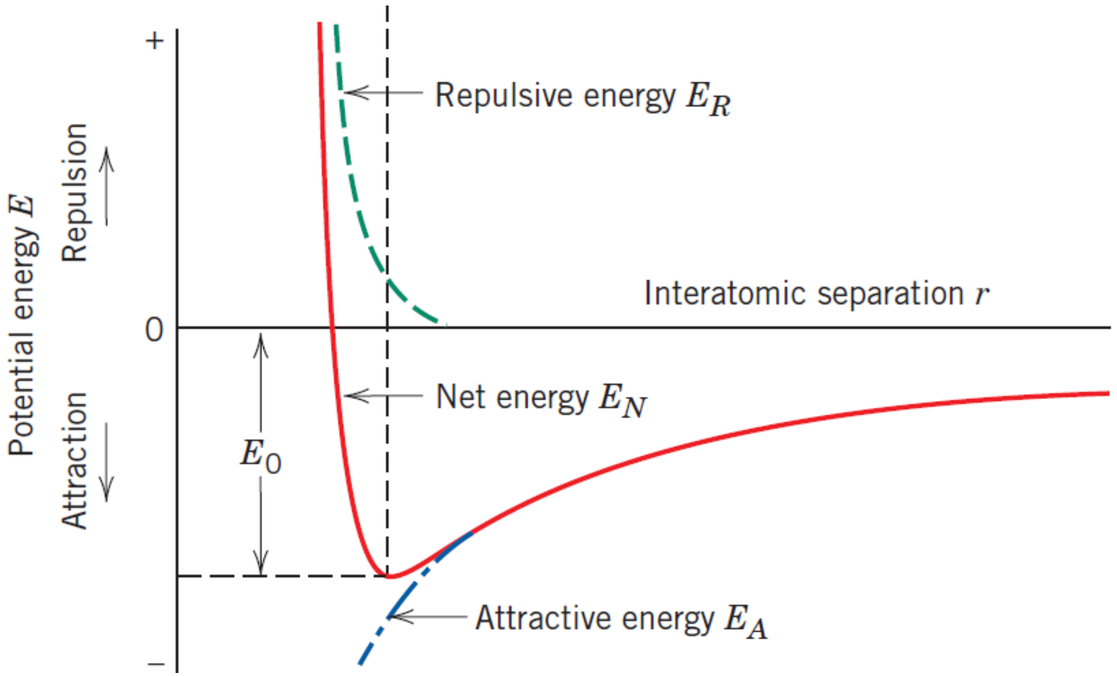
- *The mechanical stiffness (or modulus of elasticity) of a material is dependent on the shape of its force–versus–interatomic separation curve.*



Atomic Bonding in Solids

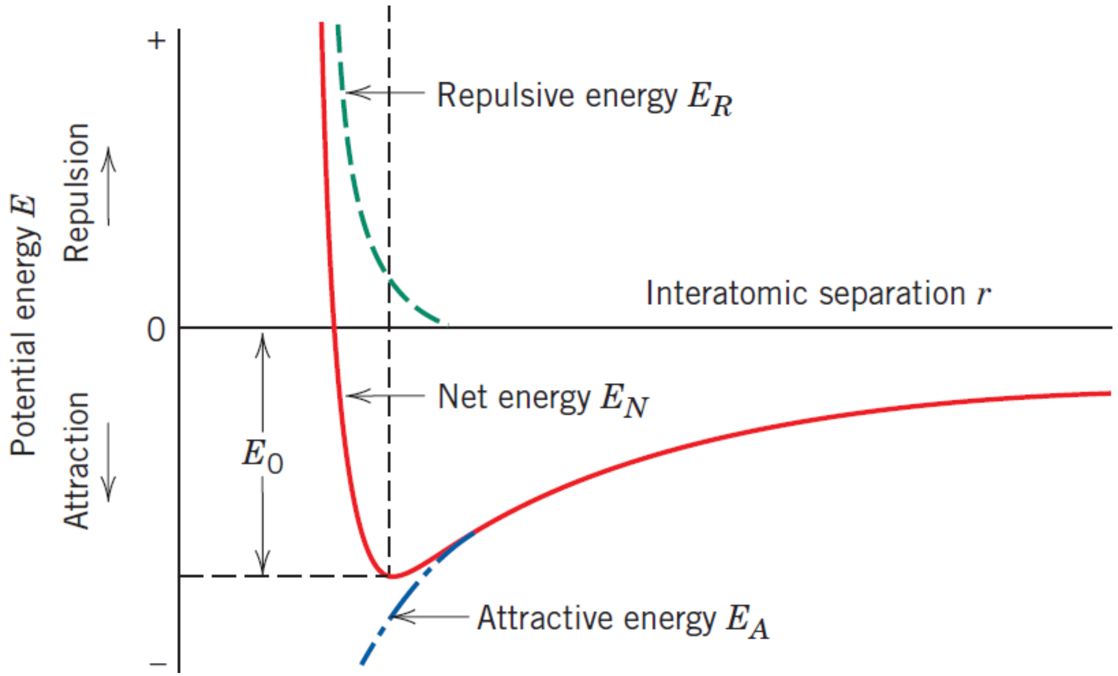
- *Furthermore, how much a material expands upon heating or contracts upon cooling (i.e., its linear coefficient of thermal expansion) is related to the shape of its E–versus–r curve.*

Three different types of **primary** or chemical **bond** are found in solids—ionic, covalent, and metallic. For each type, the bonding necessarily involves the *valence electrons*; furthermore, the nature of the bond depends on the electron structures of the constituent atoms. In general, each of these three types of bonding arises from the tendency of the atoms to assume stable electron structures, like those of the inert gases, by completely filling the outermost electron shell.



Atomic Bonding in Solids

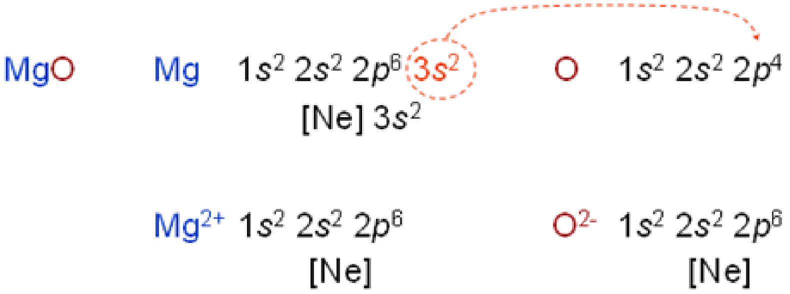
Secondary or physical forces and energies are also found in many solid materials; they are weaker than the primary ones but nonetheless influence the physical properties of some materials.



PRIMARY INTERATOMIC BONDS - Ionic Bonding

Ionic bonding is always found in compounds composed of both metallic and non-metallic elements, elements situated at the horizontal extremities of the periodic table.

Atoms of a metallic element easily give up their valence electrons to the non-metallic atoms. In the process, all the atoms acquire stable or inert gas configurations (i.e., completely filled orbital shells) and, in addition, an electrical charge—that is, they become ions.



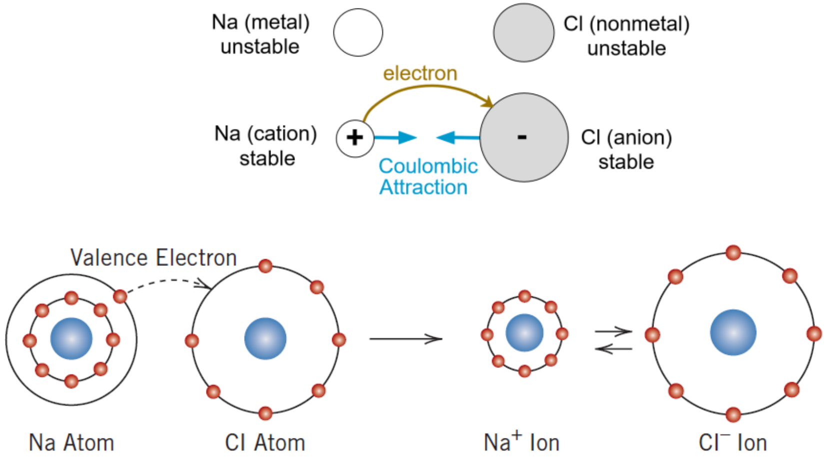
MgO

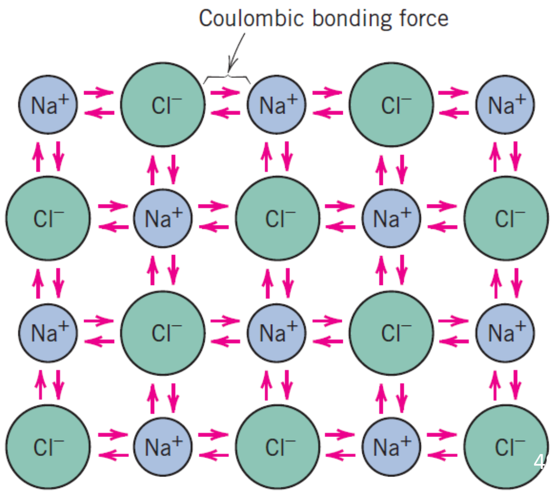
Mg $1s^2\ 2s^2\ 2p^6\ 3s^2$
[Ne] $3s^2$

O $1s^2\ 2s^2\ 2p^4$

Mg^{2+} $1s^2\ 2s^2\ 2p^6$
[Ne]

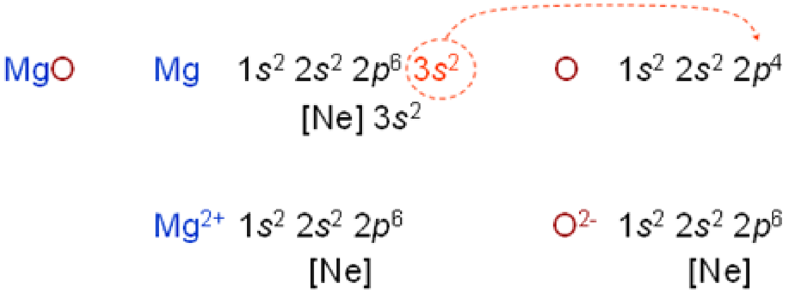
O^{2-} $1s^2\ 2s^2\ 2p^6$
[Ne]

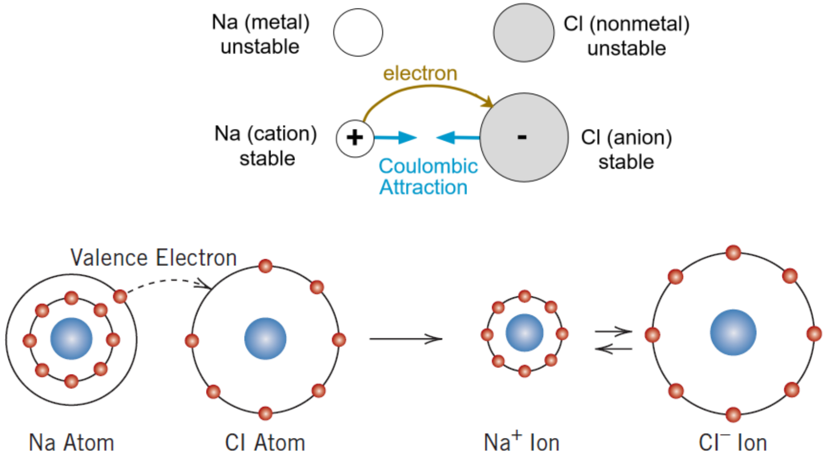


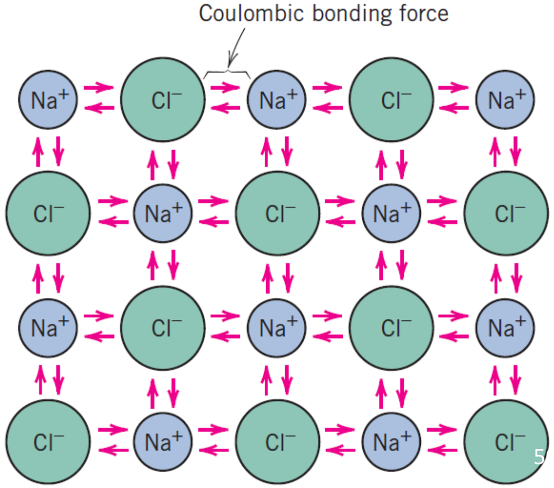


PRIMARY INTERATOMIC BONDS - Ionic Bonding

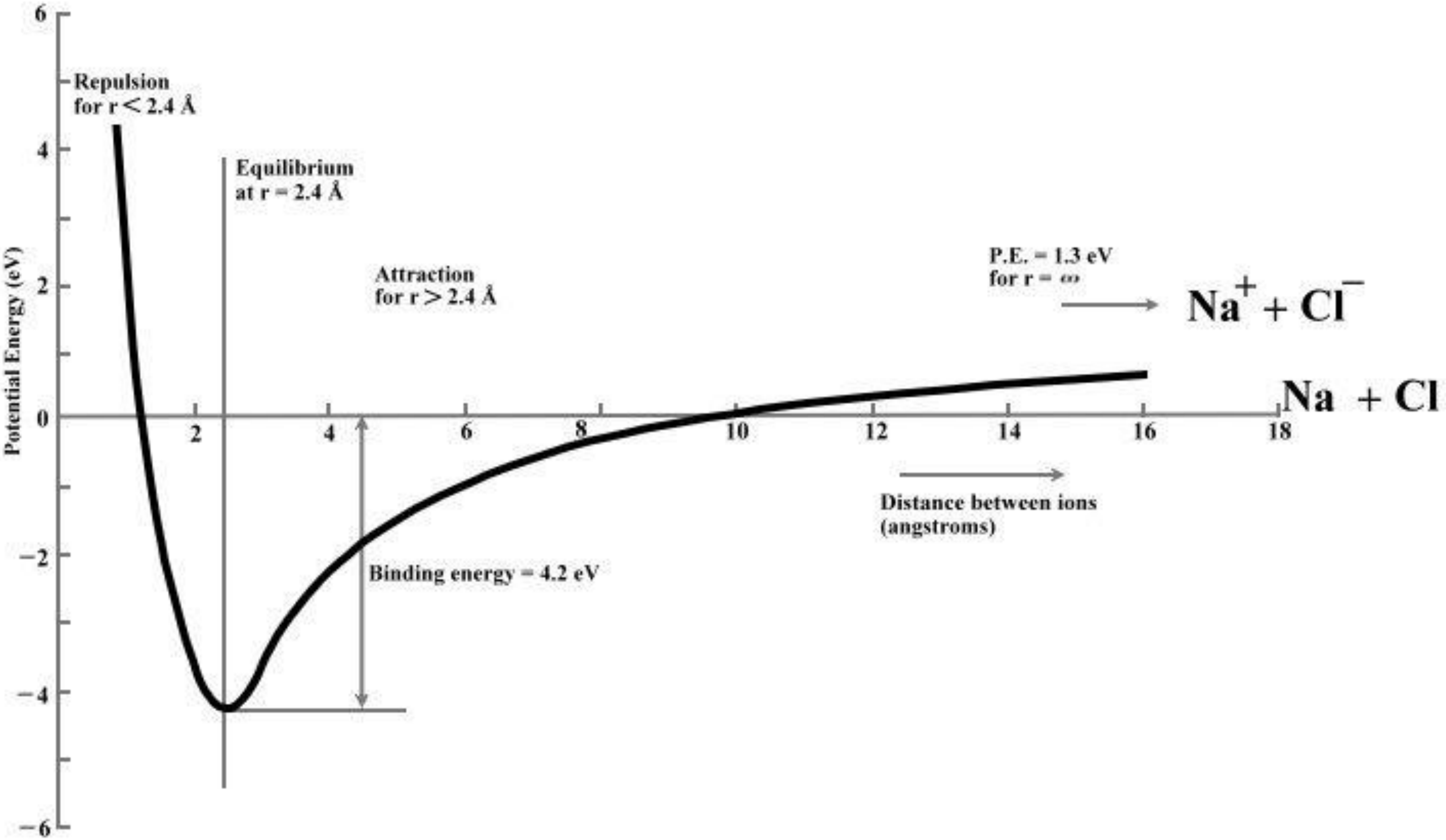
Sodium chloride (NaCl) is the classic ionic material. A sodium atom can assume the electron structure of neon (and a net single positive charge with a reduction in size) by a transfer of its one valence 3*s* electron to a chlorine atom. After such a transfer, the chlorine ion acquires a net negative charge, an electron configuration identical to that of argon; it is also larger than the chlorine atom. Ionic bonding is illustrated schematically.







PRIMARY INTERATOMIC BONDS - Ionic Bonding



PRIMARY INTERATOMIC BONDS - Ionic Bonding

The attractive bonding forces are **coulombic**—that is, positive and negative ions, by virtue of their net electrical charge, attract one another.

For two isolated ions, the attractive energy E_A is a function of the interatomic distance according to:

$$E_A = -\frac{A}{r}$$

Theoretically, the constant A is equal to: $A = \frac{1}{4\pi\varepsilon_0}(|Z_1|e)(|Z_2|e)$

Here ε_0 is the permittivity of a vacuum (8.85×10^{-12} F/m), $|Z_1|$ and $|Z_2|$ are absolute values of the valences for the two ion types, and e is the electronic charge (1.602×10^{-19} C).

PRIMARY INTERATOMIC BONDS - Ionic Bonding

The value of *A* assumes the bond between ions 1 and 2 is totally ionic. Inasmuch as bonds in most of these materials are not 100% ionic, the value of *A* is normally determined from experimental data rather than computed.

An analogous equation for the repulsive energy is:

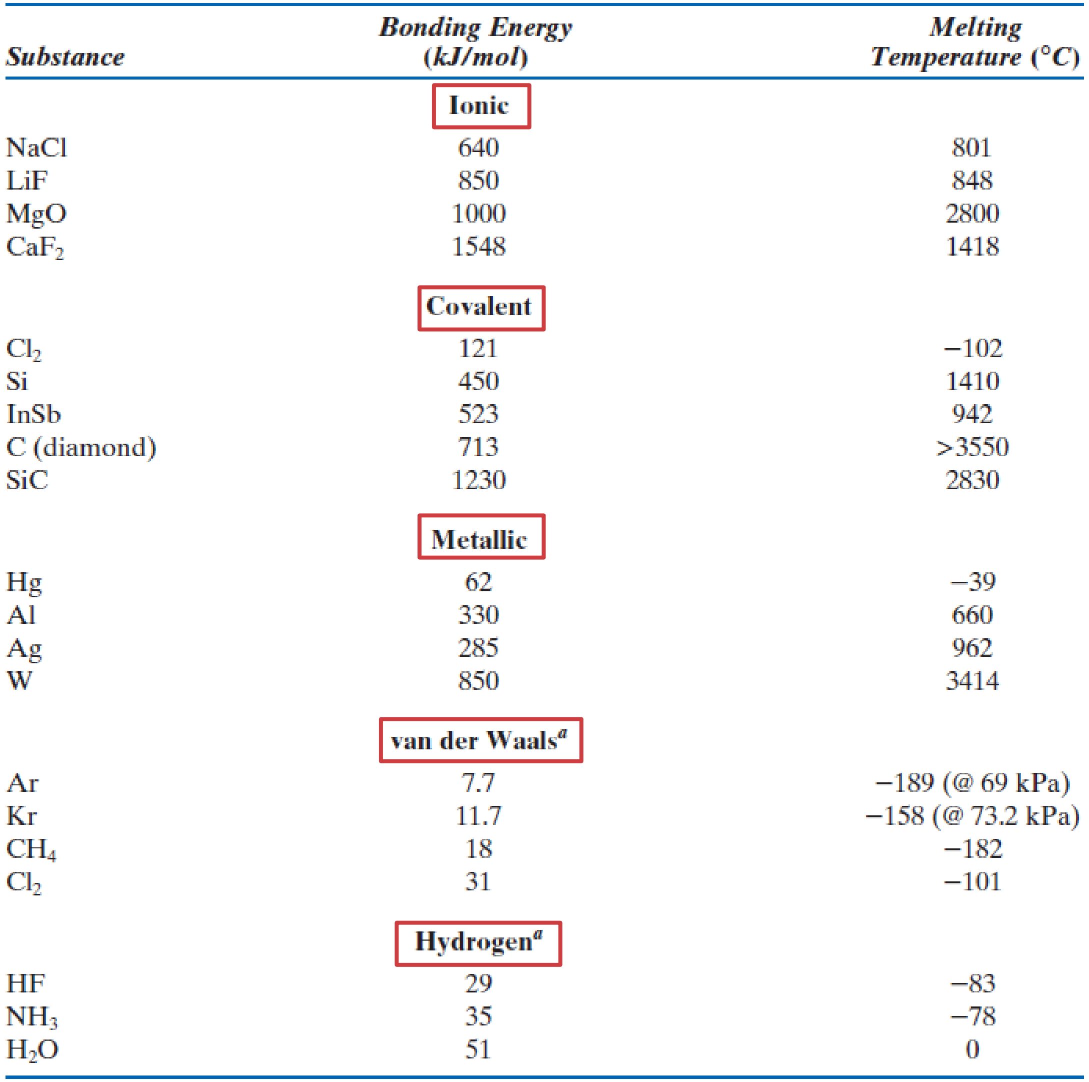
$$E_R = \frac{B}{r^n}$$

In this expression, *B* and *n* are constants whose values depend on the particular ionic system. The value of *n* is approximately 8.

PRIMARY INTERATOMIC BONDS - Ionic Bonding

Ionic bonding is termed *nondirectional*—that is, the magnitude of the bond is equal in all directions around an ion. It follows that for ionic materials to be stable, all positive ions must have as nearest neighbours negatively charged ions in a three-dimensional scheme, and vice versa.

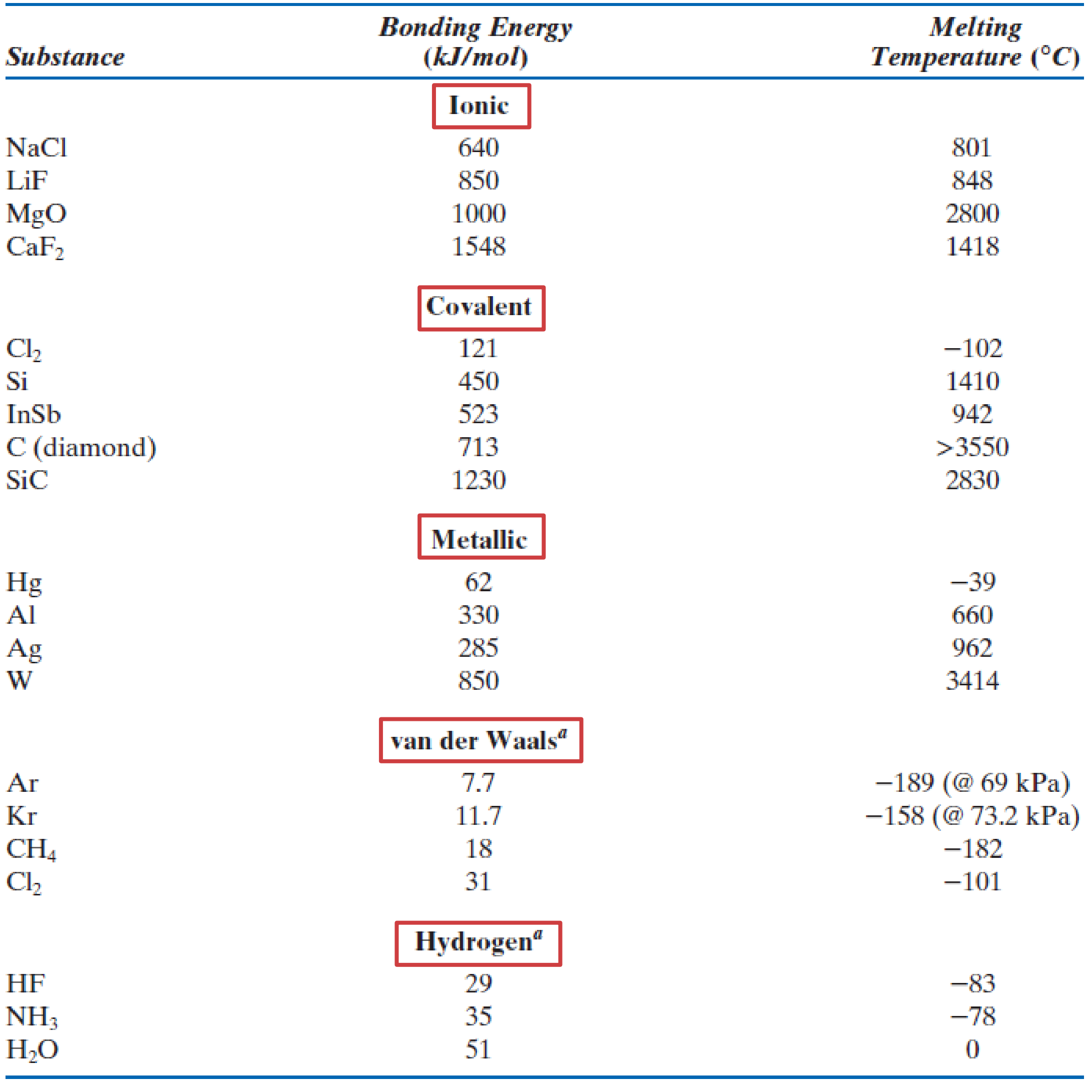
Bonding energies, which generally range between 600 and 1500 kJ/mol, are relatively large, as reflected in high melting temperatures. The table contains bonding energies and melting temperatures for several ionic materials.



Substance	*Bonding Energy (kJ/mol)*	*Melting Temperature (°C)*
	Ionic	
NaCl	640	801
LiF	850	848
MgO	1000	2800
CaF_2	1548	1418
	Covalent	
Cl_2	121	−102
Si	450	1410
InSb	523	942
C (diamond)	713	>3550
SiC	1230	2830
	Metallic	
Hg	62	−39
Al	330	660
Ag	285	962
W	850	3414
	van der Waals[a]	
Ar	7.7	−189 (@ 69 kPa)
Kr	11.7	−158 (@ 73.2 kPa)
CH_4	18	−182
Cl_2	31	−101
	Hydrogen[a]	
HF	29	−83
NH_3	35	−78
H_2O	51	0

PRIMARY INTERATOMIC BONDS - Ionic Bonding

Interatomic bonding is typified by ceramic materials, which are characteristically hard and brittle and, furthermore, electrically and thermally insulative. These properties are a direct consequence of electron configurations and/or the nature of the ionic bond.



Substance	*Bonding Energy (kJ/mol)*	*Melting Temperature (°C)*
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CH_4	18	−182
Cl_2	31	−101
	Hydrogen[a]	
HF	29	−83
NH_3	35	−78
H_2O	51	0

PRIMARY INTERATOMIC BONDS - Ionic Bonding

Example-Computation of Attractive and Repulsive Forces between Two Ions

The atomic radii of K+ and Br− ions are 0.138 and 0.196 nm, respectively.

1. Calculate the force of attraction between these two ions at their equilibrium interionic separation (i.e., when the ions just touch one another).
2. What is the force of repulsion at this same separation distance?

$$F_A = \frac{dE_A}{dr} \quad \text{and} \quad E_A = -\frac{A}{r}$$

Now, taking the derivation of *EA* with respect to *r* yields the following expression for the force of attraction *FA*:

$$F_A = \frac{dE_A}{dr} = \frac{d\left(-\frac{A}{r}\right)}{dr} = -\left(\frac{-A}{r^2}\right) = \frac{A}{r^2} \text{thus} \quad F_A = \frac{1}{4\pi\varepsilon_0 r^2}(|Z_1|e)(|Z_2|e)$$

Incorporation into this equation values for *e* and ε0 leads to:

$$F_A = \frac{1}{4\pi(8.85 \times 10^{-12}\ \text{F/m})(r^2)}[|Z_1|(1.602 \times 10^{-19}\ \text{C})][|Z_2|(1.602 \times 10^{-19}\ \text{C})]$$

$$= \frac{(2.31 \times 10^{-28}\ \text{N}\cdot\text{m}^2)(|Z_1|)(|Z_2|)}{r^2}$$

PRIMARY INTERATOMIC BONDS - Ionic Bonding

Example-Computation of Attractive and Repulsive Forces between Two Ions

The atomic radii of K^+ and Br^- ions are 0.138 and 0.196 nm, respectively.

For this problem, r is taken as the interionic separation r_0 for KBr, which is equal to the sum of the K^+ and Br^- ionic radii inasmuch as the ions touch one another—that is,

$$\begin{aligned} r_0 &= r_{K^+} + r_{Br^-} \\ &= 0.138 \text{ nm} + 0.196 \text{ nm} \\ &= 0.334 \text{ nm} \\ &= 0.334 \times 10^{-9} \text{ m} \end{aligned}$$

taking ion 1 to be K^+ and ion 2 as Br^- (i.e., $Z1 = +1$ and $Z2 = -1$), then the force of attraction is equal to:

$$F_A = \frac{(2.31 \times 10^{-28} \text{ N}\cdot\text{m}^2)(|+1|)(|-1|)}{(0.334 \times 10^{-9} \text{ m})^2} = 2.07 \times 10^{-9} \text{ N}$$

At the equilibrium separation distance the sum of attractive and repulsive forces is zero. This means that:

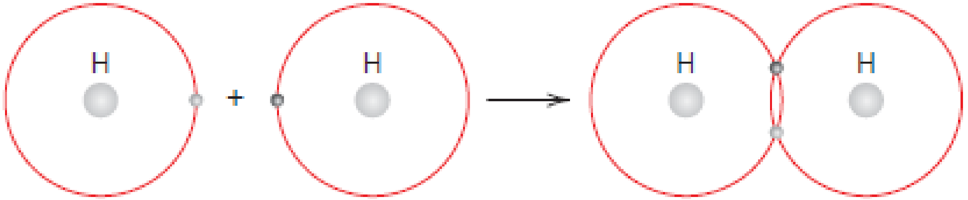
$$F_R = -F_A = -(2.07 \times 10^{-9} \text{ N}) = -2.07 \times 10^{-9} \text{ N}$$

PRIMARY INTERATOMIC BONDS - Covalent Bonding

A second bonding type, **covalent bonding,** is found in materials whose atoms have small differences in electronegativity—that is, that lie near one another in the periodic table.

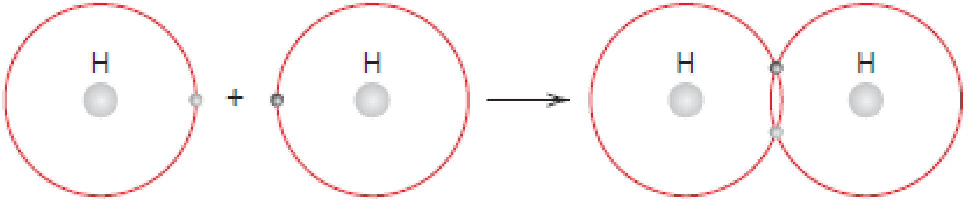
For these materials, stable electron configurations are assumed by the sharing of electrons between adjacent atoms. Two covalently bonded atoms will each contribute at least one electron to the bond, and the shared electrons may be considered to belong to both atoms.

Covalent bonding is schematically illustrated for a molecule of hydrogen (H_2). The hydrogen atom has a single 1*s* electron. Each of the atoms can acquire a helium electron configuration (two 1*s* valence electrons) when they share their single electron (right side of the figure). Furthermore, there is an overlapping of electron orbitals in the region between the two bonding atoms.



PRIMARY INTERATOMIC BONDS - Covalent Bonding

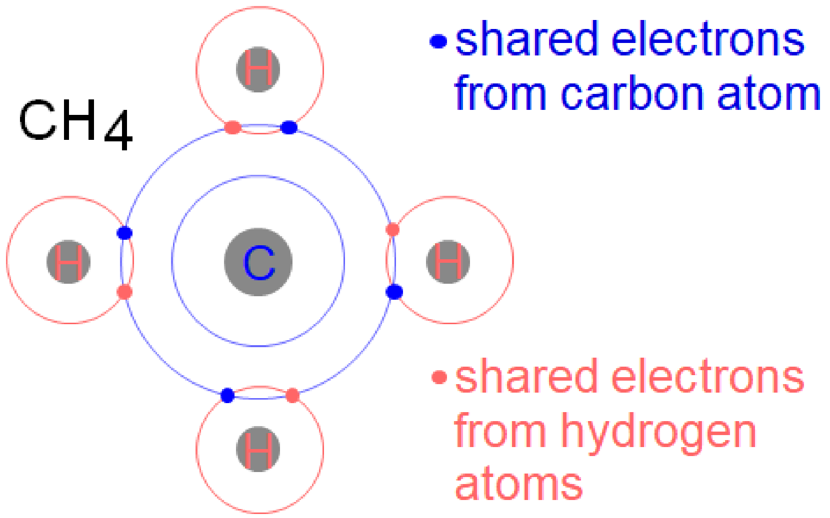
In addition, the covalent bond is *directional*—that is, it is between specific atoms and may exist only in the direction between one atom and another that participates in the electron sharing.



PRIMARY INTERATOMIC BONDS - Covalent Bonding

Many non-metallic elemental molecules (e.g., Cl_2, F_2), as well as molecules containing dissimilar atoms, such as CH_4, H_2O, HNO_3, and HF, are covalently bonded.

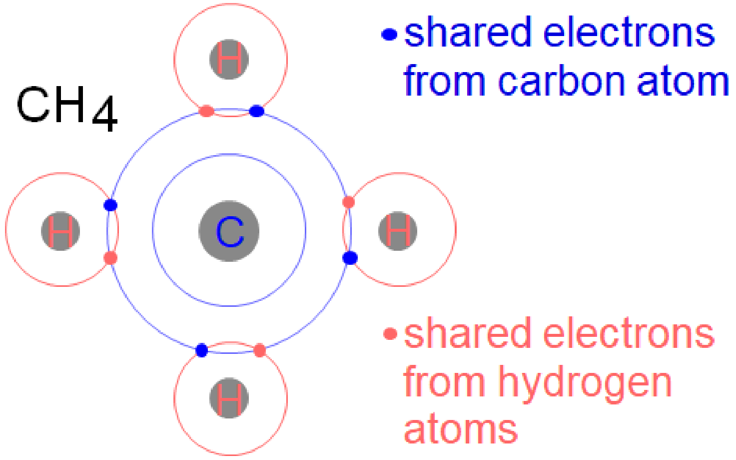
Furthermore, this type of bonding is found in elemental solids such as diamond (carbon), silicon, and germanium and other solid compounds composed of elements that are located on the right side of the periodic table, such as gallium arsenide (GaAs), indium antimonide (InSb), and silicon carbide (SiC).



● shared electrons from carbon atom

● shared electrons from hydrogen atoms

PRIMARY INTERATOMIC BONDS - Covalent Bonding



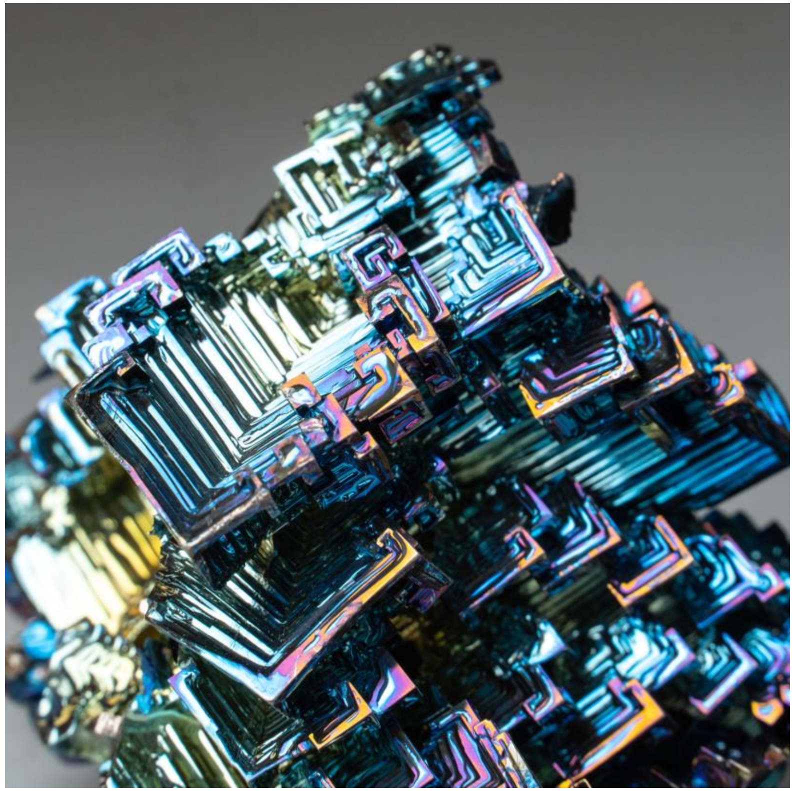
Covalent bonds may be very strong, as in diamond, which is very hard and has a very high melting temperature, >3550°C (6400°F), or they may be very weak, as with bismuth, which melts at about 270°C (518°F).

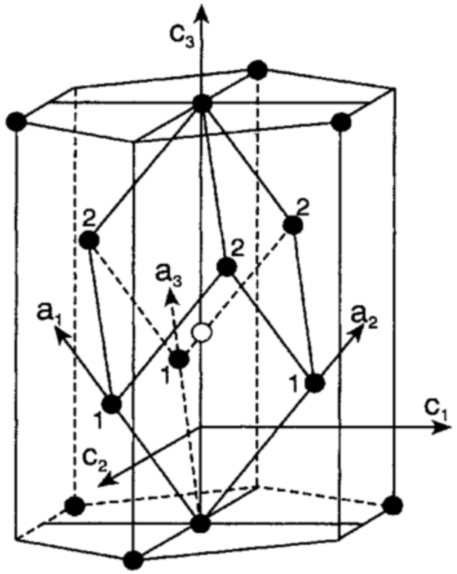
Electrons participating in covalent bonds are tightly bound to the bonding atoms, most covalently bonded materials are electrical insulators, or, in some cases, semiconductors.

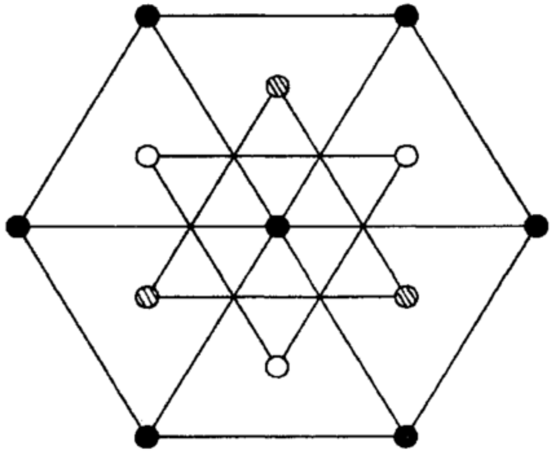
Mechanical behaviours of these materials vary widely: some are relatively strong, others are weak; some fail in a brittle manner, whereas others experience significant amounts of deformation before failure.

It is difficult to predict the mechanical properties of covalently bonded materials on the basis of their bonding characteristics.

PRIMARY INTERATOMIC BONDS - Covalent Bonding

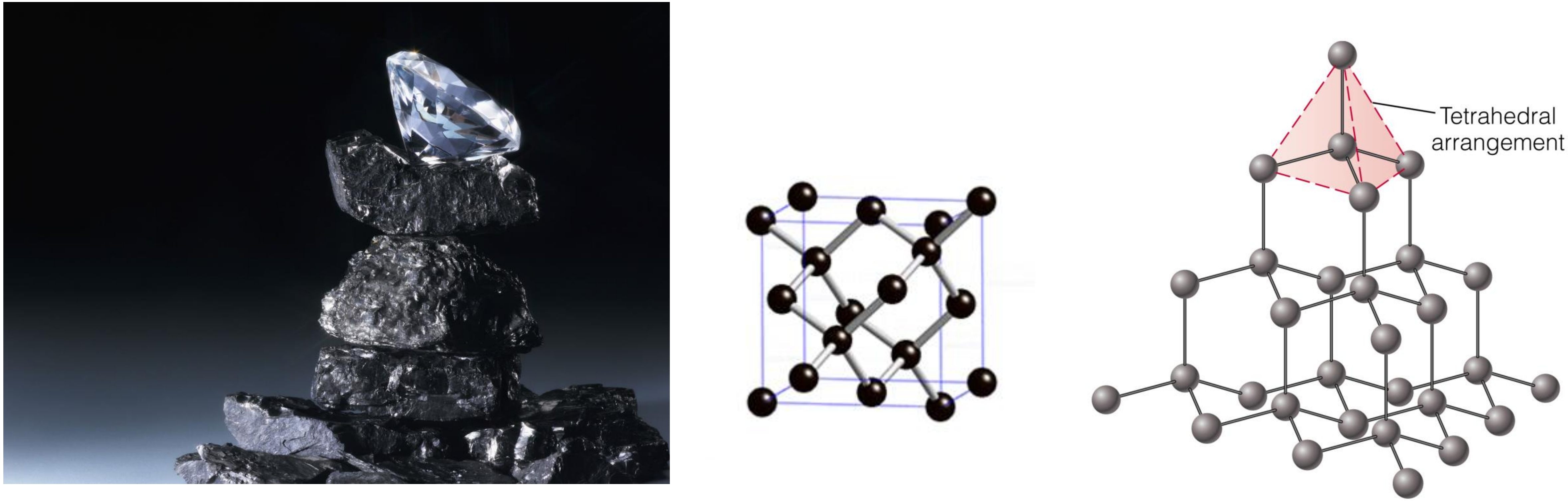






Bismuth Crystal

PRIMARY INTERATOMIC BONDS - Covalent Bonding



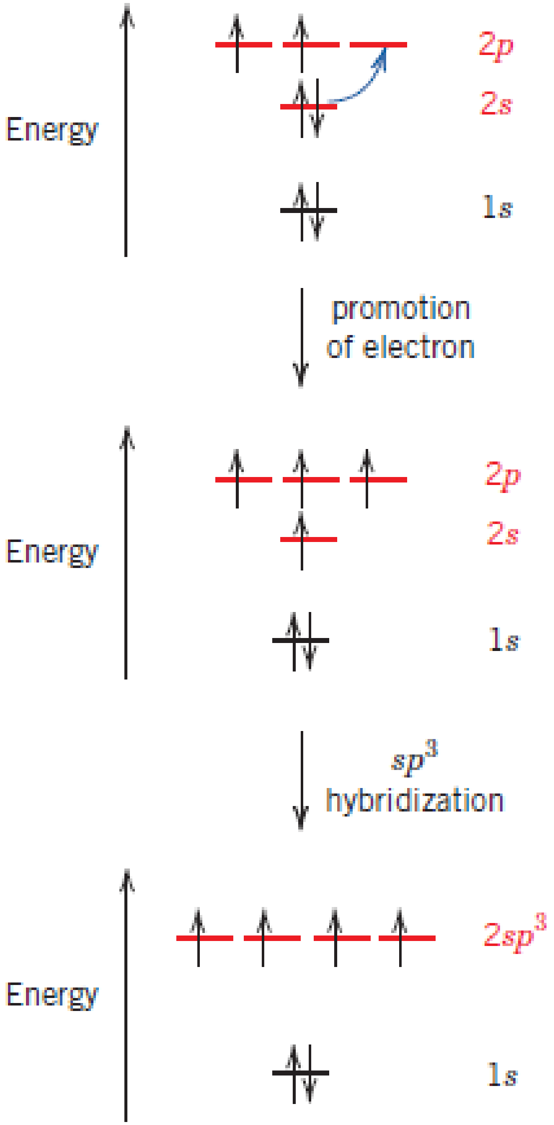
Diamond Crystal

PRIMARY INTERATOMIC BONDS - Covalent Bonding

Bond Hybridization in Carbon

Often associated with the covalent bonding of carbon (as well other non-metallic substances) is the phenomenon of ***hybridization***—the mixing (or combining) of two or more atomic orbitals with the result that more orbital overlap during bonding results.

For example, consider the electron configuration of carbon: $1s^2 2s^2 2p^2$. Under some circumstances, one of the 2*s* orbitals is promoted to the empty 2*p* orbital, which gives rise to a $1s^2 2s^1 2p^3$ configuration.

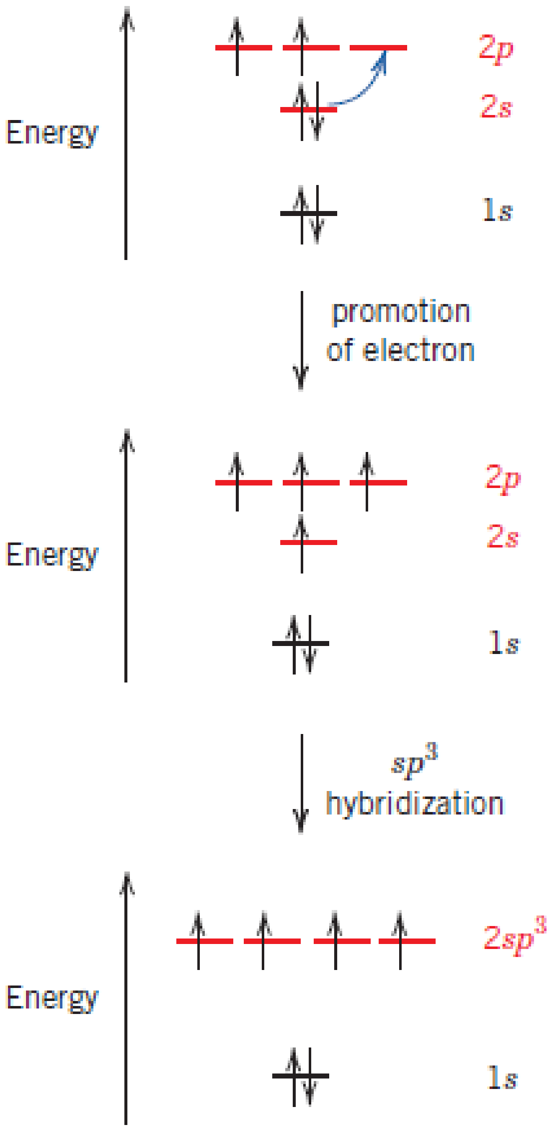


Schematic diagram that shows the formation of sp^3 hybrid orbitals in carbon. (*a*) Promotion of a 2*s* electron to a 2*p* state; (*b*) this promoted electron in a 2*p* state; (*c*) four $2sp^3$ orbitals that form by mixing the single 2*s* orbital with the three 2*p* orbitals.

PRIMARY INTERATOMIC BONDS - Covalent Bonding

Furthermore, the 2*s* and 2*p* orbitals can mix to produce four sp^3 orbitals that are equivalent to one another, have parallel spins, and are capable of covalently bonding with other atoms.

This orbital mixing is termed ***hybridization***, which leads to the electron configuration shown; here, each sp^3 orbital contains one electron, and, therefore, is half-filled.



Schematic diagram that shows the formation of sp^3 hybrid orbitals in carbon. (*a*) Promotion of a 2*s* electron to a 2*p* state; (*b*) this promoted electron in a 2*p* state; (*c*) four $2sp^3$ orbitals that form by mixing the single 2*s* orbital with the three 2*p* orbitals.

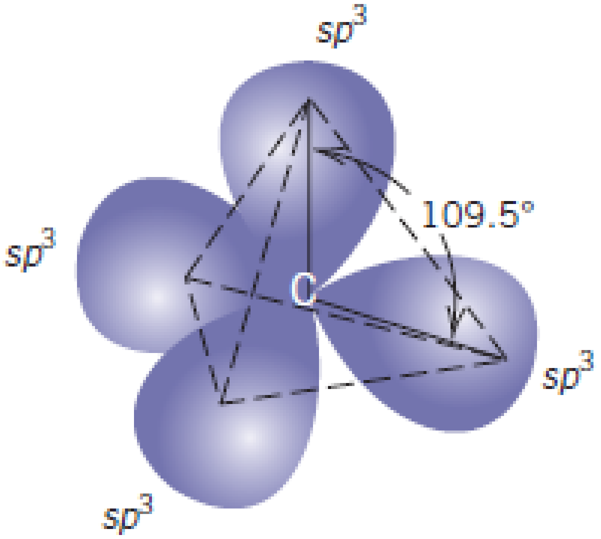
PRIMARY INTERATOMIC BONDS - Covalent Bonding

Bond Hybridization in Carbon

Bonding hybrid orbitals are directional in nature—that is, each extends to and overlaps the orbital of an adjacent bonding atom.

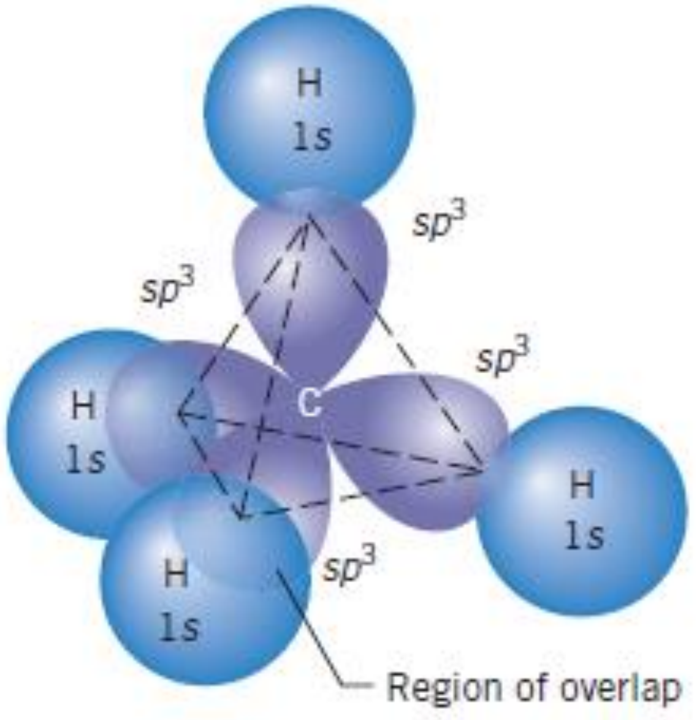
Furthermore, for carbon, each of its four sp^3 hybrid orbitals is directed symmetrically from a carbon atom to the vertex of a tetrahedron—a configuration represented schematically; the angle between each set of adjacent bonds is 109.5°.

Schematic diagram showing four sp^3 hybrid orbitals that point to the corners of a tetrahedron; the angle between orbitals is 109.5°.



PRIMARY INTERATOMIC BONDS - Covalent Bonding

The bonding of *sp*^{3} hybrid orbitals to the 1*s* orbitals of four hydrogen atoms, as in a molecule of methane (CH4), is presented.



Schematic diagram that shows bonding of carbon sp^3 hybrid orbitals to the 1*s* orbitals of four hydrogen atoms in a molecule of methane (CH_4).

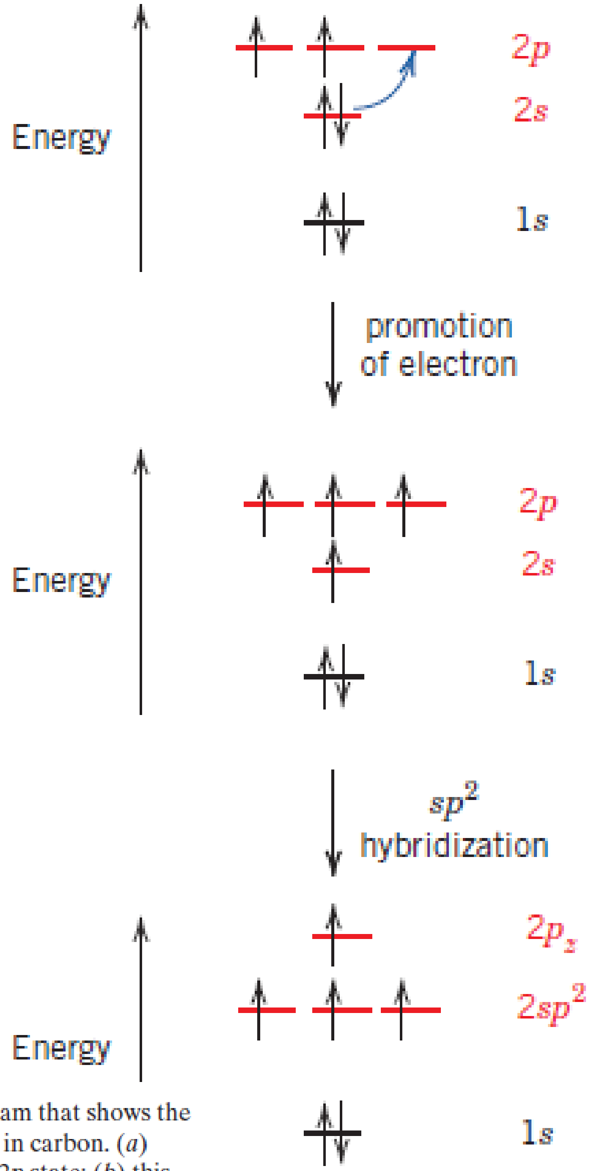
PRIMARY INTERATOMIC BONDS - Covalent Bonding

Bond Hybridization in Carbon

For diamond, its carbon atoms are bonded to one another with sp^3 covalent hybrids—each atom is bonded to four other carbon atoms.

The crystal structure for diamond is shown. Diamond's carbon–carbon bonds are extremely strong, which accounts for its high melting temperature and ultrahigh hardness (it is the hardest of all materials).

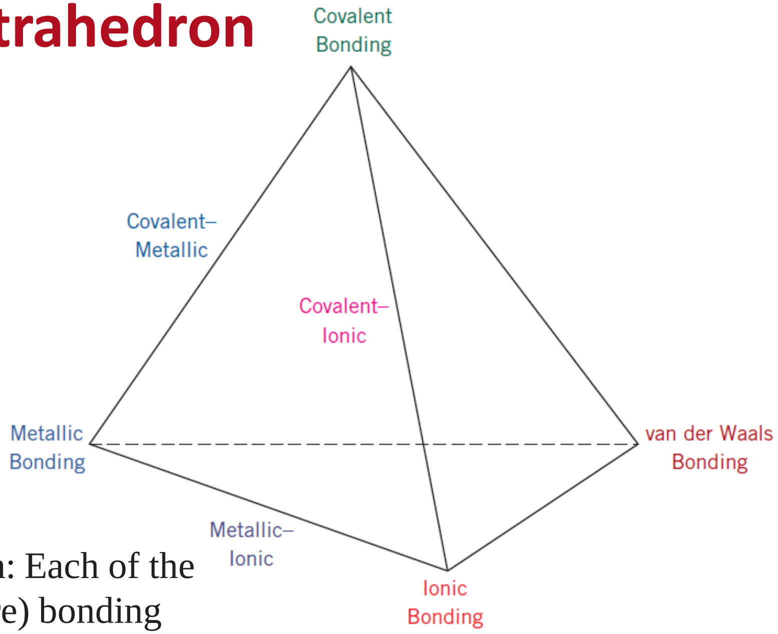
Many polymeric materials are composed of long chains of carbon atoms that are also bonded together using sp^3 tetrahedral bonds; these chains form a zigzag structure because of this 109.5° inter-bonding angle.



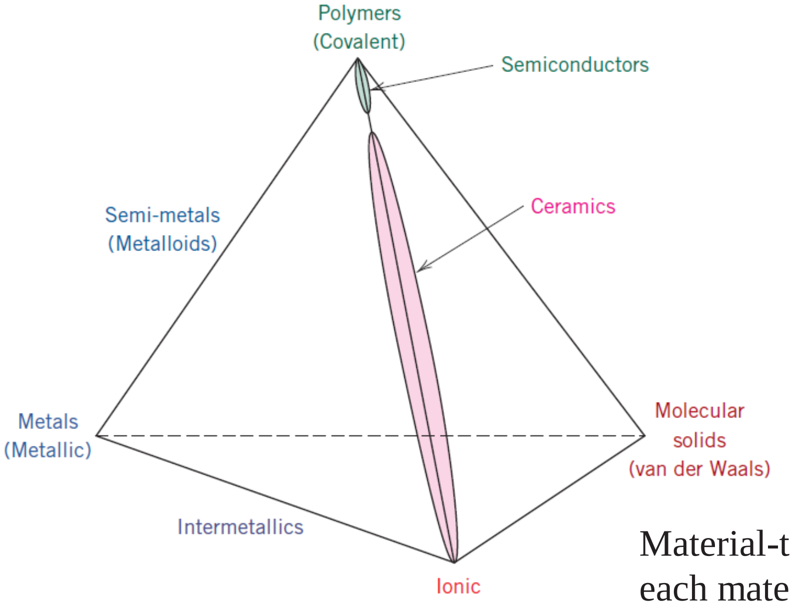
Schematic diagram that shows the formation of sp^2 hybrid orbitals in carbon. (*a*) Promotion of a 2*s* electron to a 2*p* state; (*b*) this promoted electron in a 2*p* state; (*c*) three $2sp^2$ orbitals that form by mixing the single 2*s* orbital with two 2*p* orbitals—the $2p_z$ orbital remains unhybridized.

PRIMARY INTERATOMIC BONDS - Mixed Bonding

Bonding Tetrahedron



Bonding tetrahedron: Each of the four extreme (or pure) bonding types is located at one corner of the tetrahedron; three mixed bonding types are included along tetrahedron edges.



Material-type tetrahedron: correlation of each material classification (metals, ceramics, polymers, etc.) with its type(s) of bonding.

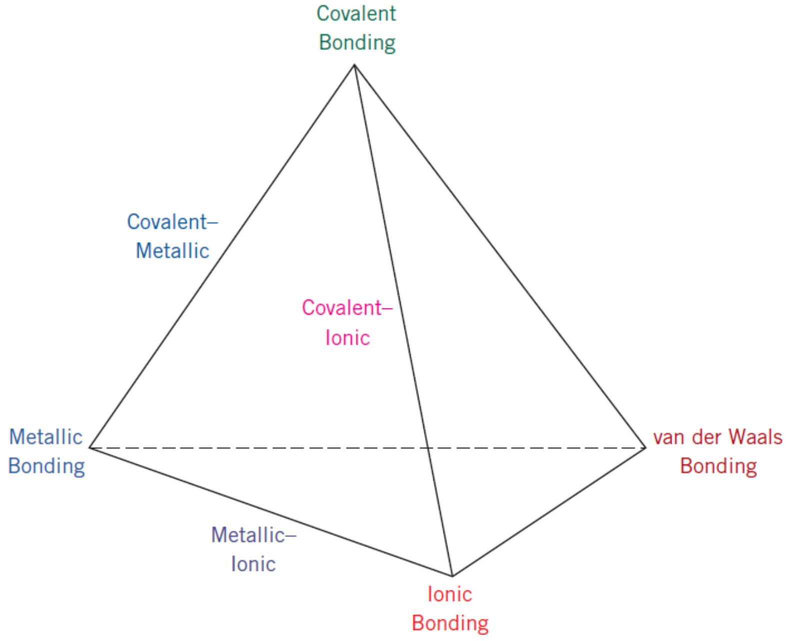
For many real materials, the atomic bonds are mixtures of two or more of these extremes (i.e., *mixed bonds*). Three mixed-bond types—covalent–ionic, covalent–metallic, and metallic–ionic—are also included on edges of this tetrahedron.

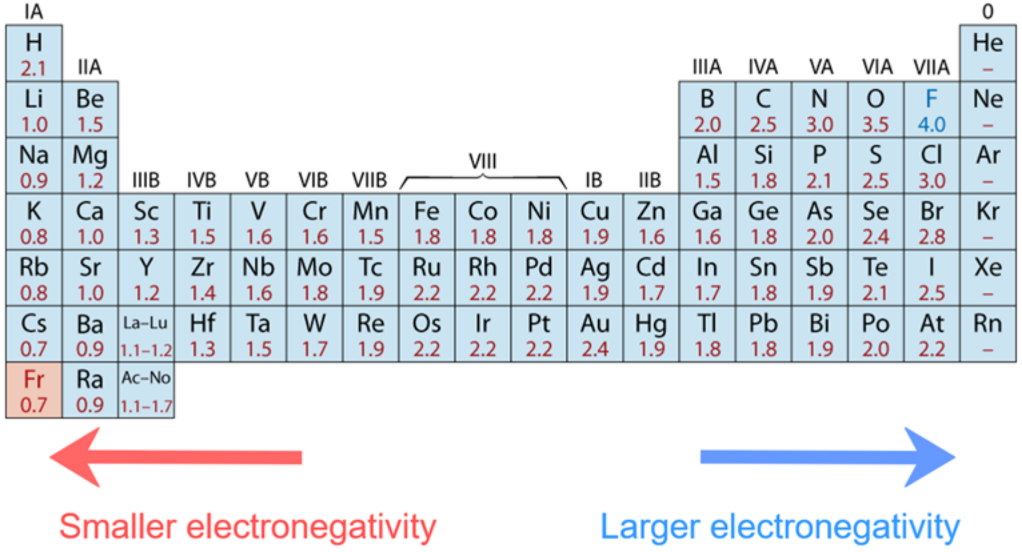
PRIMARY INTERATOMIC BONDS - Mixed Bonding

Bonding Tetrahedron

For mixed covalent–ionic bonds, there is some ionic character to most covalent bonds and some covalent character to ionic ones.

As such, there is a continuum between these two extreme bond types. This type of bond is represented between the ionic and covalent bonding vertices. The degree of either bond type depends on the relative positions of the constituent atoms in the periodic table or the difference in their electronegativities.





IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
H 2.1																	He –
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne –
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar –
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr –
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe –
Cs 0.7	Ba 0.9	La–Lu 1.1–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn –
Fr 0.7	Ra 0.9	Ac–No 1.1–1.7															

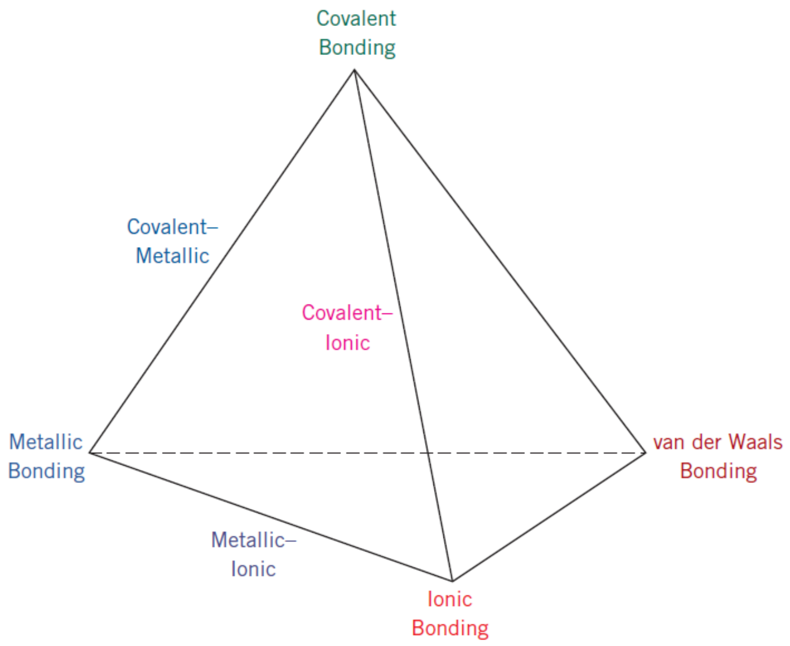
Smaller electronegativity

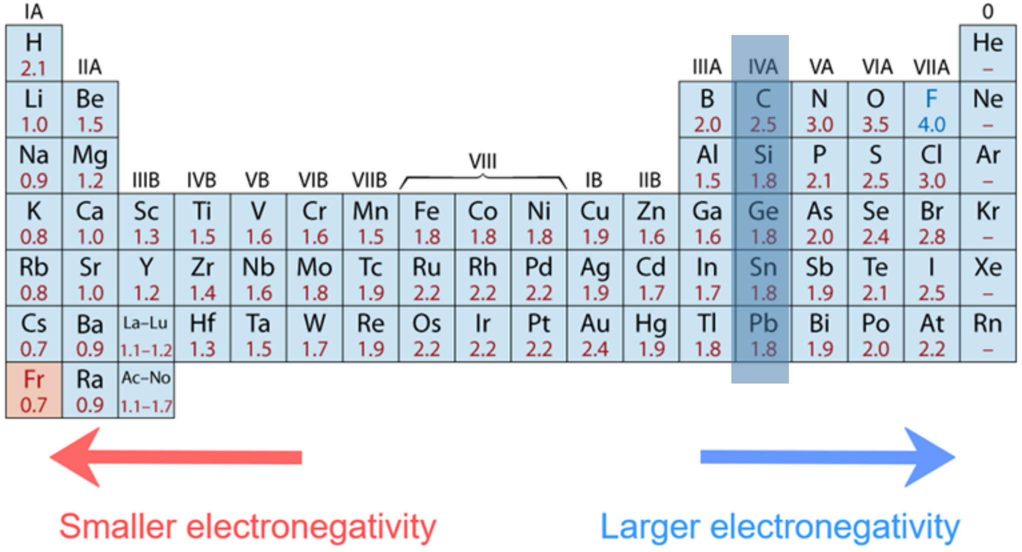
Larger electronegativity

PRIMARY INTERATOMIC BONDS - Mixed Bonding

The wider the separation (both horizontally— relative to Group IVA— and vertically) from the lower left to the upper right corner (i.e., the greater the difference in electronegativity), the more ionic is the bond.

Conversely, the closer the atoms are together (i.e., the smaller the difference in electronegativity), the greater is the degree of covalency.





IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
H 2.1																	He –
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne –
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar –
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr –
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe –
Cs 0.7	Ba 0.9	La–Lu 1.1–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn –
Fr 0.7	Ra 0.9	Ac–No 1.1–1.7															

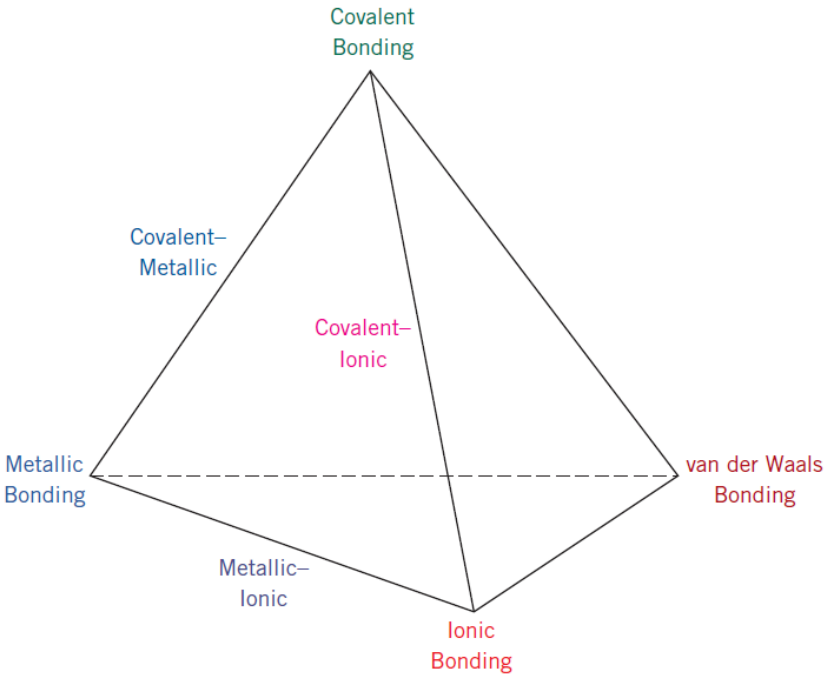
PRIMARY INTERATOMIC BONDS - Mixed Bonding

Bonding Tetrahedron

Percent ionic character (%IC) of a bond between elements A and B (A being the most electronegative) may be approximated by the expression:

$$\%IC = \{1 - \exp[-(0.25)(X_A - X_B)^2]\} \times 100$$

where X_A and X_B are the electronegativities for the respective elements.



PRIMARY INTERATOMIC BONDS - Mixed Bonding

Example- Calculation of the Percent Ionic Character for the C-H Bond

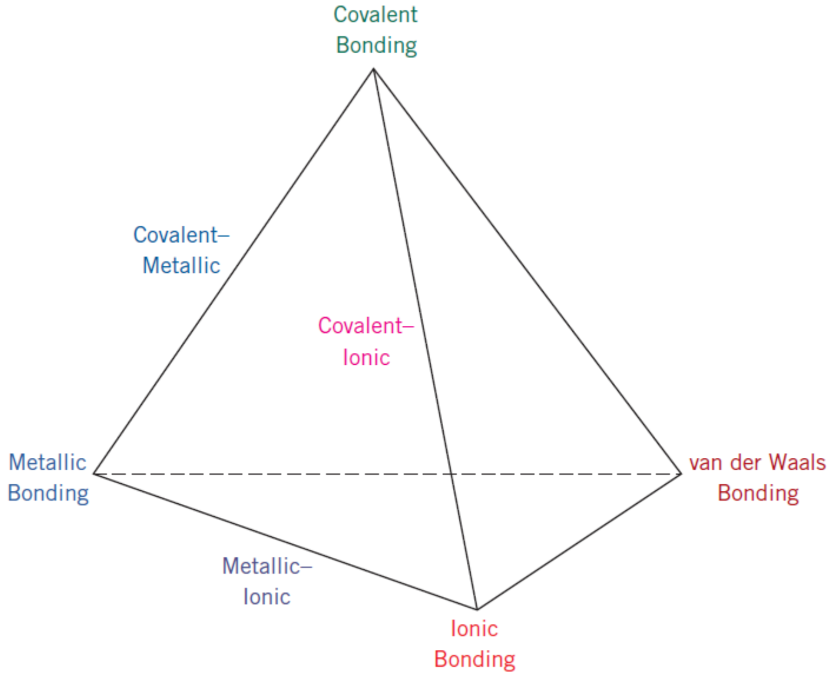
The %IC of a bond between two atoms/ions, A and B (A being the more electronegative), is a function of their electronegativities X_A and X_B.

The electronegativities for C and H are $X_C = 2.5$ and $X_H = 2.1$. Therefore, the %IC is:

$$\%IC = \{1 - \exp[-(0.25)(X_C - X_H)^2]\} \times 100 = \{1 - \exp[-(0.25)(2.5 - 2.1)^2]\} \times 100$$

$$= 3.9\%$$

Thus the C-H atomic bond is primarily covalent (96.1%).



PRIMARY INTERATOMIC BONDS - Mixed Bonding

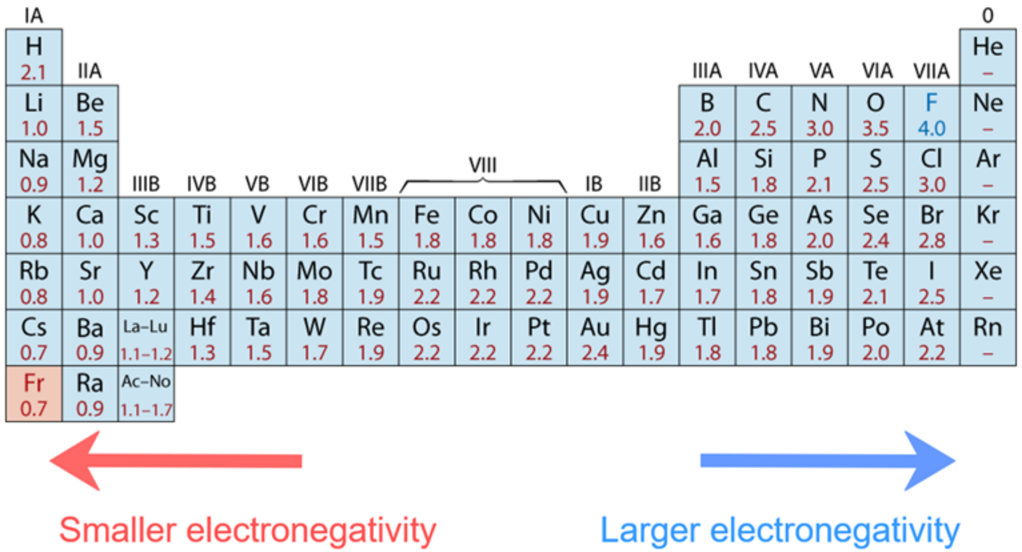
Example- Calculation of the Percent Ionic Character for the Mg-O Bond

The electronegativities for Mg and O are $X_{Mg} = 1.2$ and $X_O = 3.5$. Therefore, the %IC is:

$\%IC = \{1 - \exp[-(0.25)(X_{Mg} - X_O)^2]\} \times 100 = \{1 - \exp[-(0.25)(1.2 - 3.5)^2]\} \times 100$

$= 73.4\%$

Thus the Mg-O atomic bond is primarily ionic.



IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
H 2.1																	He –
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne –
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar –
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr –
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe –
Cs 0.7	Ba 0.9	La–Lu 1.1–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn –
Fr 0.7	Ra 0.9	Ac–No 1.1–1.7															

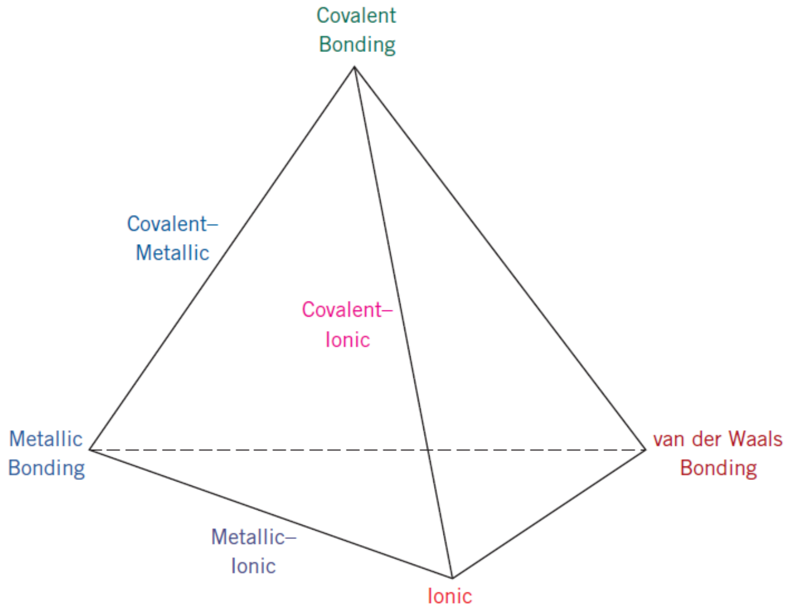
Smaller electronegativity

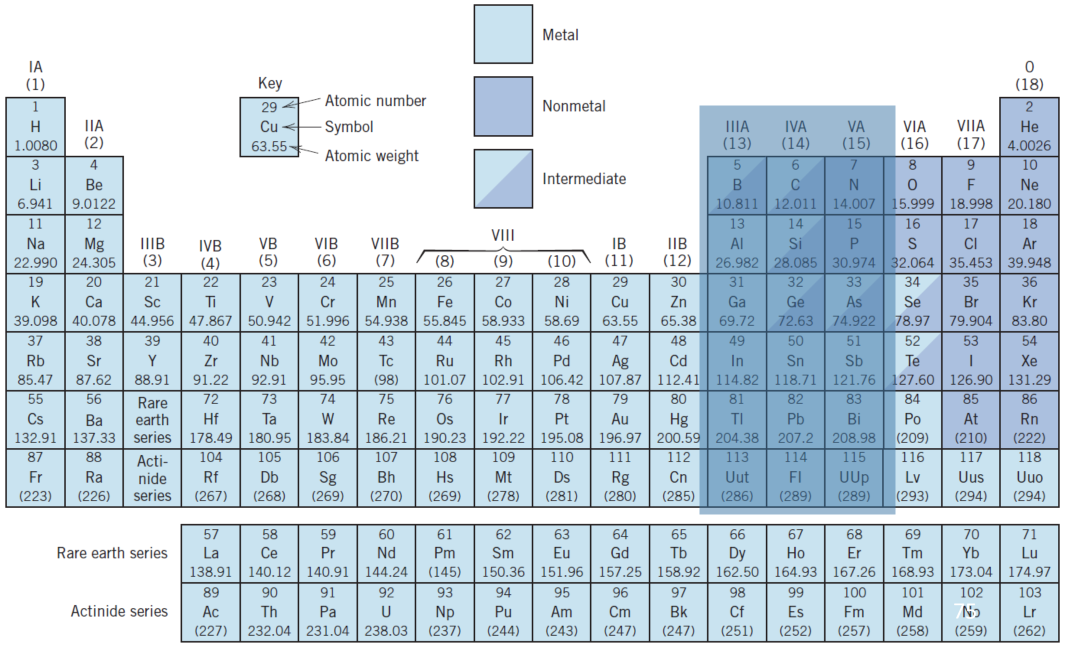
Larger electronegativity

PRIMARY INTERATOMIC BONDS - Mixed Bonding

Another type of mixed bond is found for some elements in Groups IIIA, IVA, and VA of the periodic table (viz., B, Si, Ge, As, Sb, Te, Po, and At).

Interatomic bonds for these elements are mixtures of metallic and covalent. These materials are called the *metalloids* or *semi-metals*, and their properties are intermediate between the metals and nonmetals.





Key: 29 ← Atomic number; Cu ← Symbol; 63.55 ← Atomic weight

- ▢ Metal
- ▨ Nonmetal
- ◩ Intermediate

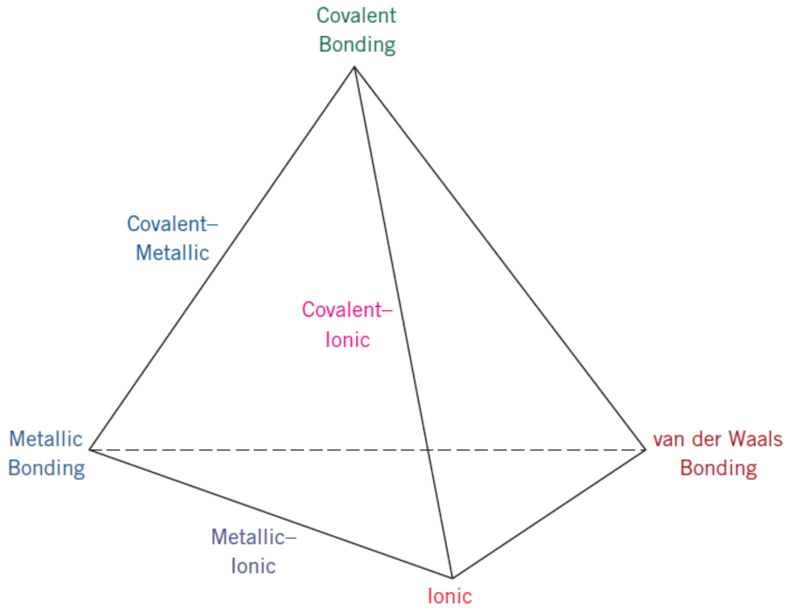
IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	(9)	(10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
1 H 1.0080																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.085	15 P 30.974	16 S 32.064	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.922	34 Se 78.97	35 Br 79.904	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	Rare earth series	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	Acti-nide series	104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (269)	109 Mt (278)	110 Ds (281)	111 Rg (280)	112 Cn (285)	113 Uut (286)	114 Fl (289)	115 UUp (289)	116 Lv (293)	117 Uus (294)	118 Uuo (294)

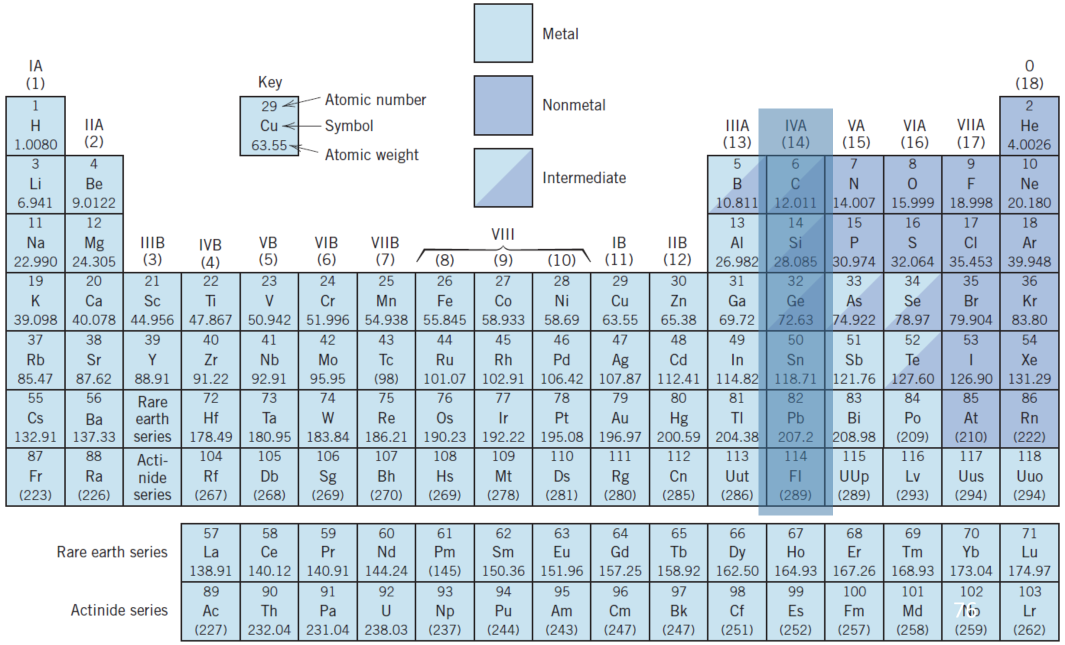
Rare earth series	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.92	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
Actinide series	89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

PRIMARY INTERATOMIC BONDS - Mixed Bonding

In addition, for Group IV elements, there is a gradual transition from covalent to metallic bonding as one moves vertically down this column—for example, bonding in carbon (diamond) is purely covalent, whereas for tin and lead, bonding is predominantly metallic.

Mixed metallic–ionic bonds are observed for compounds composed of two metals when there is a significant difference between their electronegativities. This means that some electron transfer is associated with the bond inasmuch as it has an ionic component.





Key: 29
Cu
63.55
- Atomic number
- Symbol
- Atomic weight

- Metal
- Nonmetal
- Intermediate

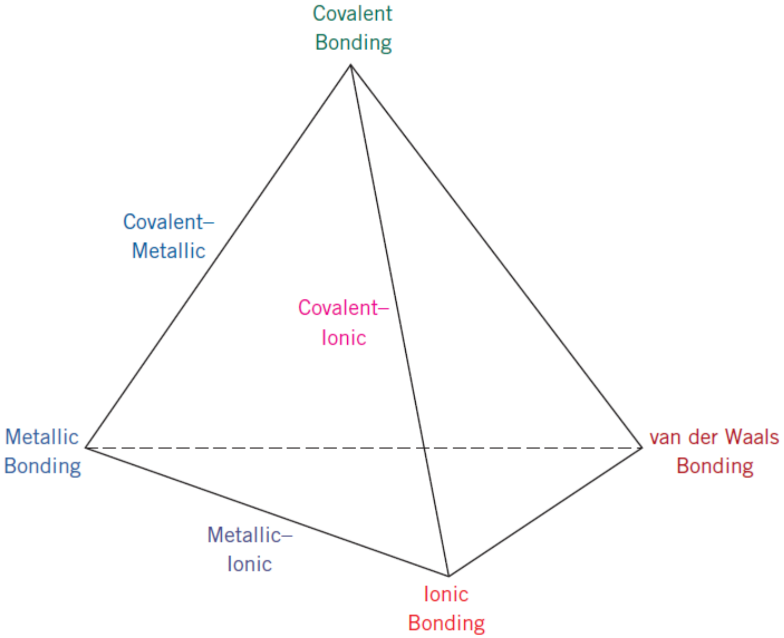
IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	(9)	(10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
1 H 1.0080																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.085	15 P 30.974	16 S 32.064	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.922	34 Se 78.97	35 Br 79.904	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	Rare earth series	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	Actinide series	104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (269)	109 Mt (278)	110 Ds (281)	111 Rg (280)	112 Cn (285)	113 Uut (286)	114 Fl (289)	115 UUp (289)	116 Lv (293)	117 Uus (294)	118 Uuo (294)

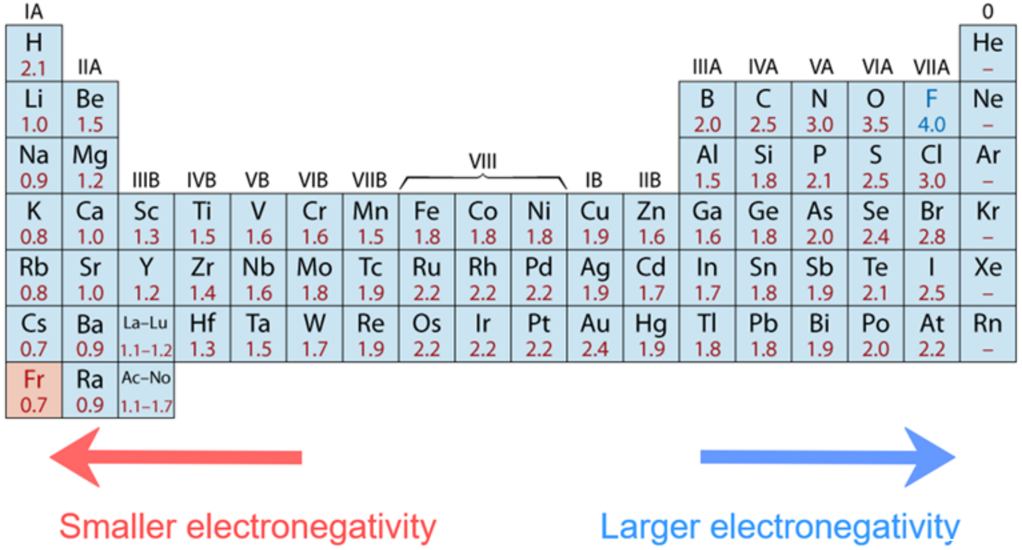
Rare earth series	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.92	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
Actinide series	89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

PRIMARY INTERATOMIC BONDS - Mixed Bonding

Furthermore, the larger this electronegativity difference, the greater the degree of ionicity. For example, there is little ionic character to the titanium–aluminium bond for the intermetallic compound $TiAl_3$ because electronegativities of both Al and Ti are the same (1.5).

However, a much greater degree of ionic character is present for $AuCu_3$; the electronegativity difference for copper and gold is 0.5.





IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
H 2.1																	He –
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne –
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar –
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr –
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe –
Cs 0.7	Ba 0.9	La–Lu 1.1–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn –
Fr 0.7	Ra 0.9	Ac–No 1.1–1.7															

PRIMARY INTERATOMIC BONDS - Metallic Bonding

Metallic bonding, the final primary bonding type, is found in metals and their alloys. A relatively simple model has been proposed that very nearly approximates the bonding scheme.

With this model, these valence electrons are not bound to any particular atom in the solid and are more or less free to drift throughout the entire metal.

They may be thought of as belonging to the metal as a whole, or forming a “sea of electrons” or an “electron cloud.”

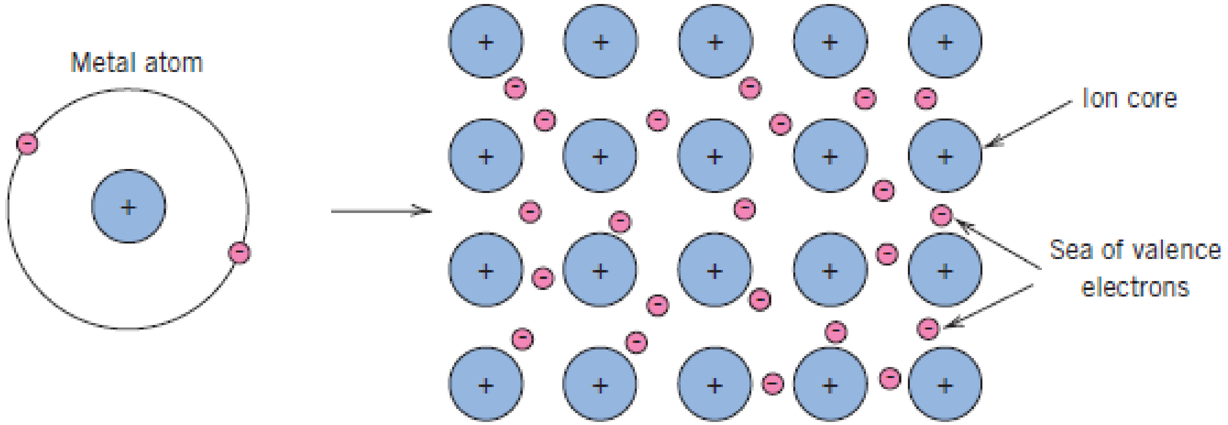
The remaining non-valence electrons and atomic nuclei form what are called *ion cores*, which possess a net positive charge equal in magnitude to the total valence electron charge per atom.

PRIMARY INTERATOMIC BONDS - Metallic Bonding

The free electrons shield the positively charged ion cores from the mutually repulsive electrostatic forces that they would otherwise exert upon one another; consequently, the metallic bond is nondirectional in character.

In addition, these free electrons act as a “glue” to hold the ion cores together.

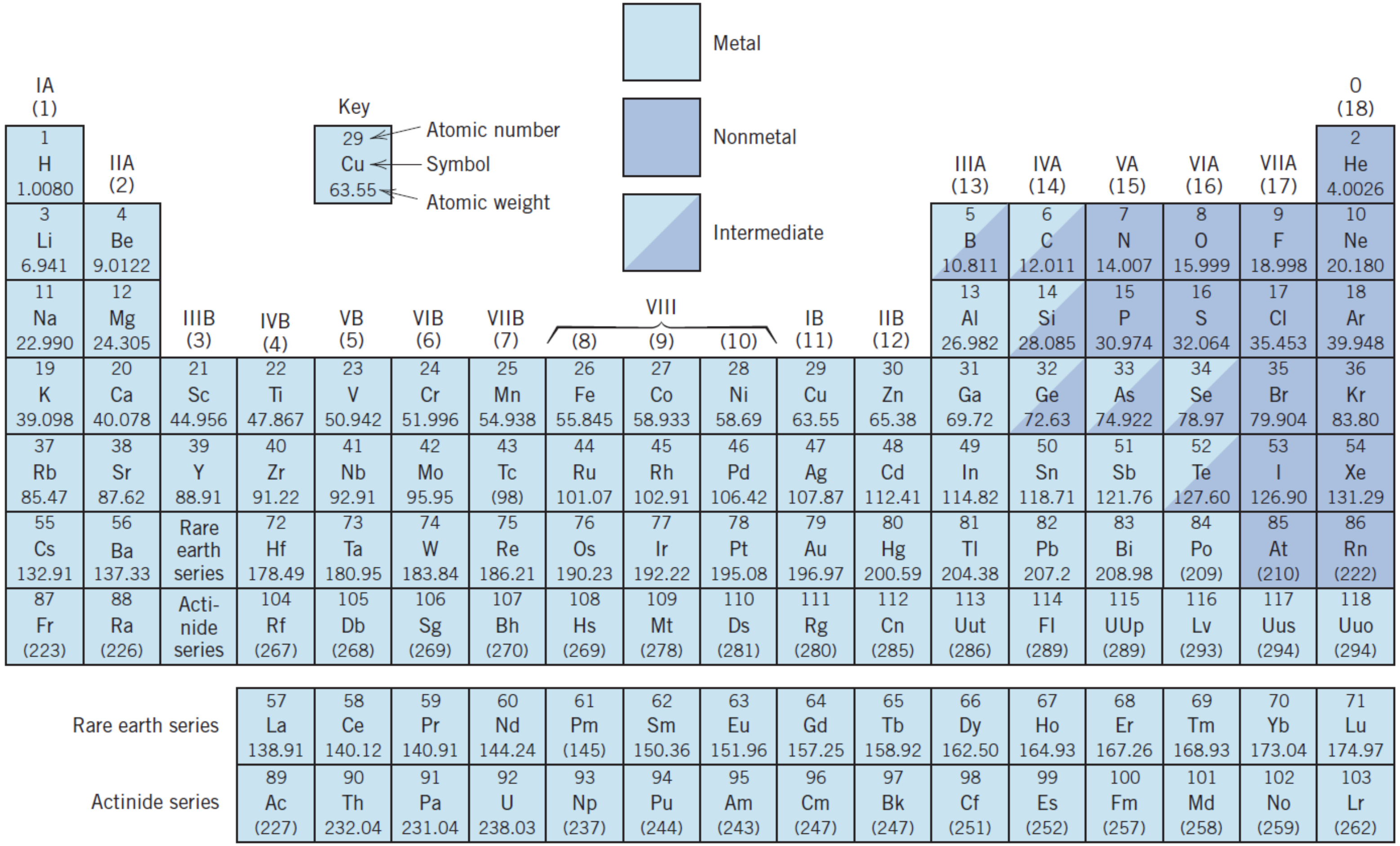
Bonding may be weak or strong; energies range from 62 kJ/mol for mercury to 850 kJ/mol for tungsten. Their respective melting temperatures are −39°C and 3414°C (−39°F and 6177°F).



PRIMARY INTERATOMIC BONDS - Metallic Bonding

Metallic bonding is found in the periodic table for Group IA and IIA elements and, in fact, for all elemental metals.

Metals are good conductors of both electricity and heat as a consequence of their free electrons. Furthermore, we note that at room temperature, most metals and their alloys fail in a ductile manner—that is, fracture occurs after the materials have experienced significant degrees of permanent deformation.

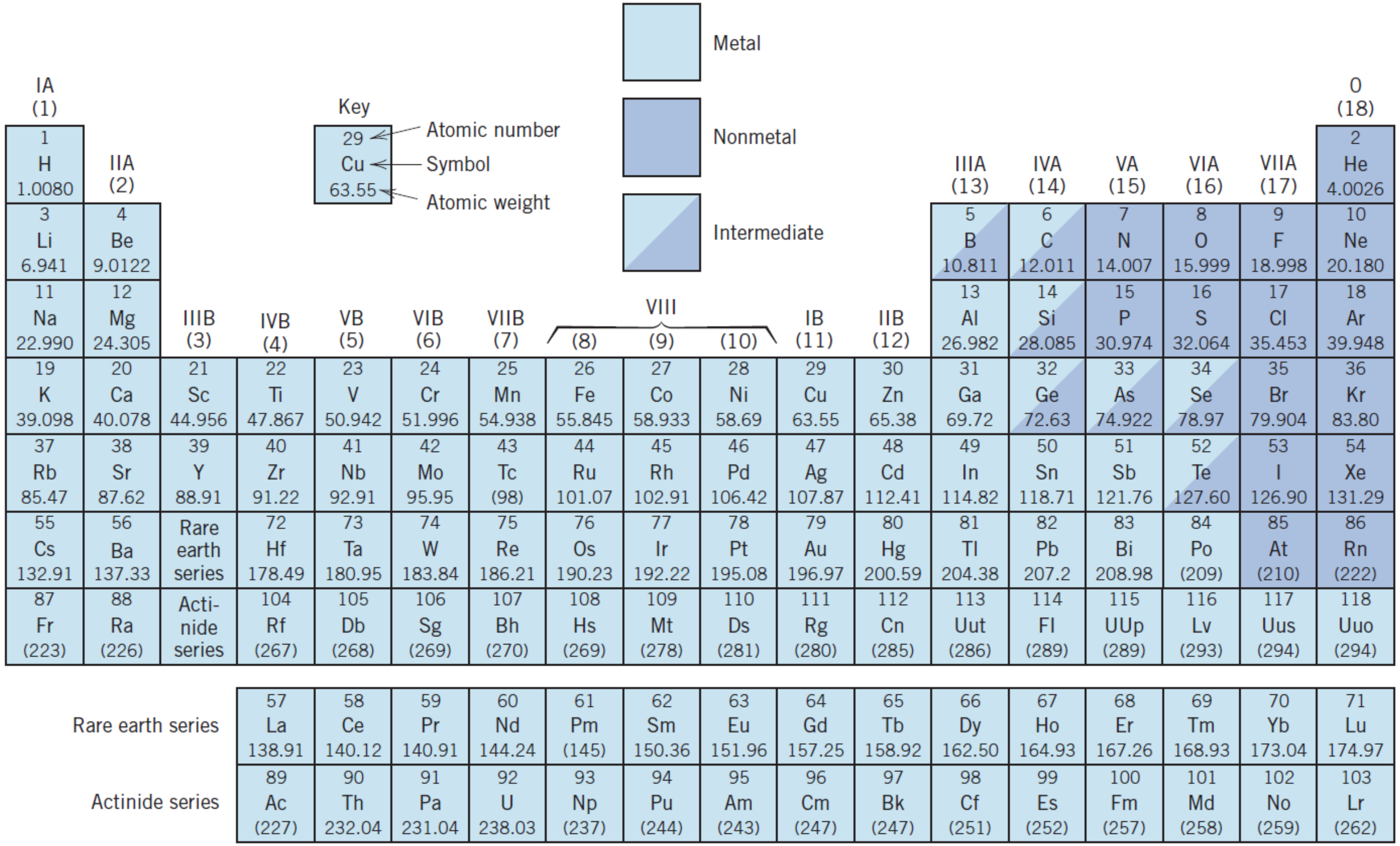


IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	VIII (9)	VIII (10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
1 H 1.0080																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.085	15 P 30.974	16 S 32.064	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.922	34 Se 78.97	35 Br 79.904	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	Rare earth series	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	Acti- nide series	104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (269)	109 Mt (278)	110 Ds (281)	111 Rg (280)	112 Cn (285)	113 Uut (286)	114 Fl (289)	115 UUp (289)	116 Lv (293)	117 Uus (294)	118 Uuo (294)

Rare earth series	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.92	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
Actinide series	89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

PRIMARY INTERATOMIC BONDS - Metallic Bonding

This behaviour is explained in terms of a deformation mechanism, which is implicitly related to the characteristics of the metallic bond.



IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	VIII (9)	VIII (10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
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37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
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87 Fr (223)	88 Ra (226)	Actinide series	104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (269)	109 Mt (278)	110 Ds (281)	111 Rg (280)	112 Cn (285)	113 Uut (286)	114 Fl (289)	115 UUp (289)	116 Lv (293)	117 Uus (294)	118 Uuo (294)

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IONIC VS COVALENT BONDS

Ionic Bonds

In an ionic bond, one atom essentially donates an electron to stabilize the other atom.
In other words, the electron spends most of its time close to the bonded atom. Atoms that participate in an ionic bond have different electronegativity values from each other.

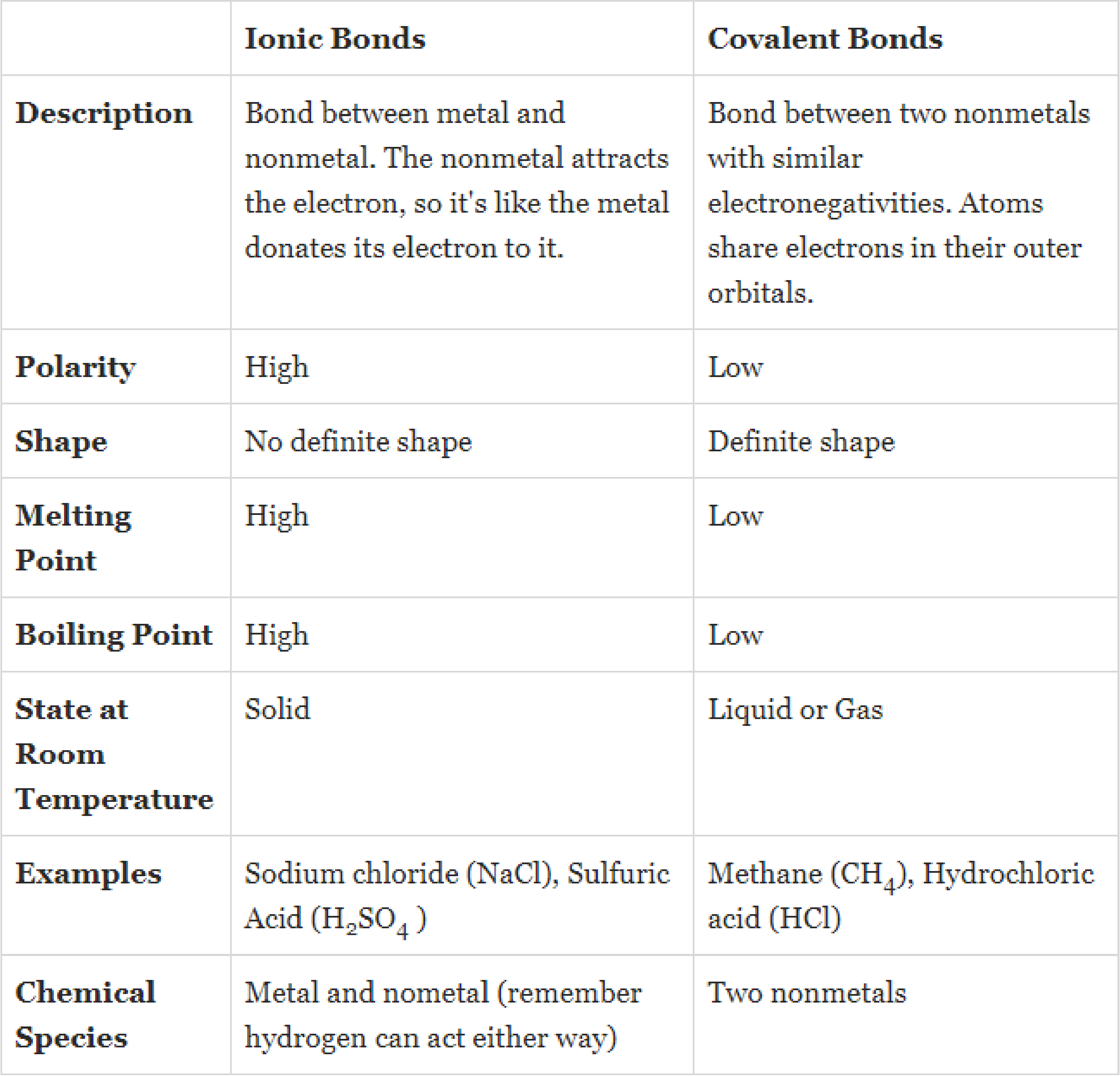
A polar bond is formed by the attraction between oppositely-charged ions. For example, sodium and chloride form an ionic bond, to make NaCl, or table salt.
You can predict an ionic bond will form when two atoms have different electronegativity values and detect an ionic compound by its properties, including a tendency to dissociate into ions in water.

Covalent Bonds

In a covalent bond, the atoms are bound by shared electrons.
In a true covalent bond, the electronegativity values are the same (e.g., H_2, O_3), although in practice the electronegativity values just need to be close.
If the electron is shared equally between the atoms forming a covalent bond, then the bond is said to be nonpolar. Usually, an electron is more attracted to one atom than to another, forming a polar covalent bond. For example, the atoms in water, H_2O, are held together by polar covalent bonds.

You can predict a covalent bond will form between two non-metallic atoms. Also, covalent compounds may dissolve in water, but don't dissociate into ions.

IONIC VS COVALENT BONDS



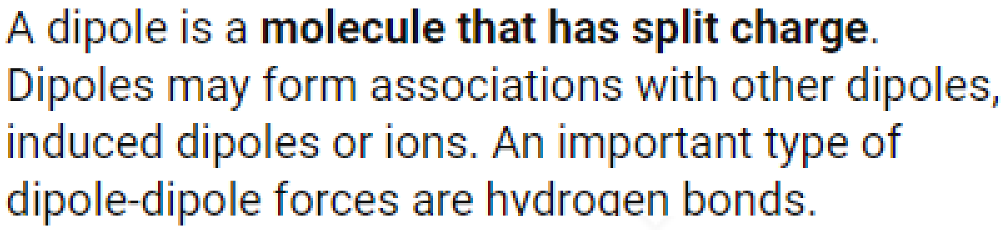
	Ionic Bonds	**Covalent Bonds**
Description	Bond between metal and nonmetal. The nonmetal attracts the electron, so it's like the metal donates its electron to it.	Bond between two nonmetals with similar electronegativities. Atoms share electrons in their outer orbitals.
Polarity	High	Low
Shape	No definite shape	Definite shape
Melting Point	High	Low
Boiling Point	High	Low
State at Room Temperature	Solid	Liquid or Gas
Examples	Sodium chloride (NaCl), Sulfuric Acid (H_2SO_4)	Methane (CH_4), Hydrochloric acid (HCl)
Chemical Species	Metal and nometal (remember hydrogen can act either way)	Two nonmetals

POLAR COVALENT BOND

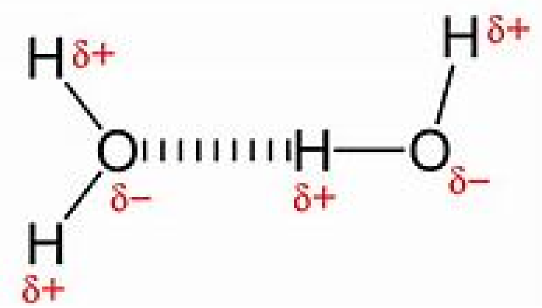
Chemical bonds may be classified as being either polar or nonpolar. The difference is how the electrons in the bond are arranged.

Polar Bond Definition

A polar bond is a covalent bond between two atoms wh the electrons forming the bond are unequally distributed. This causes t molecule to have a slight electrical **dipole** moment where one end is slightly positive and the other is slightly negative.
The charge of the electric dipoles is less than a full unit charge, so they are considered partial charges and denoted by delta plus (δ+) and delta minus (δ-). Because positive and negative charges are separated in the bond, molecules with polar covalent bonds interact with dipoles in other molecules. This produces dipole-dipole intermolecular forces between the molecules.



A dipole is a **molecule that has split charge**. Dipoles may form associations with other dipoles, induced dipoles or ions. An important type of dipole-dipole forces are hydrogen bonds.



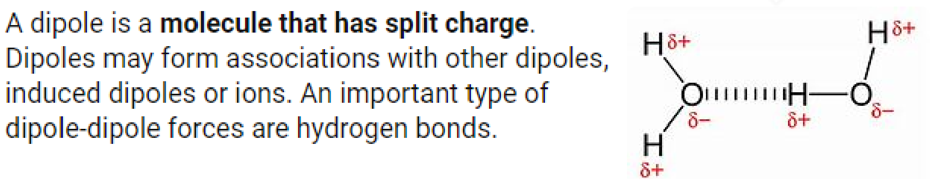
POLAR COVALENT BOND

Polar bonds are the dividing line between pure covalent bonding and pure ionic bonding. Pure covalent bonds (nonpolar covalent bonds) share electron pairs equally between atoms. Technically, nonpolar bonding only occurs when the atoms are identical to each other (e.g., H2 gas), but chemists consider *any bond between atoms with a difference in electronegativity less than 0.4 to be a nonpolar covalent bond.*

Carbon dioxide (CO2) and methane (CH4) are nonpolar molecules.

In ionic bonds, the electrons in the bond are essentially donated to one atom by the other (e.g., NaCl). Ionic bonds form between atoms when the electronegativity difference between them is greater than 1.7.

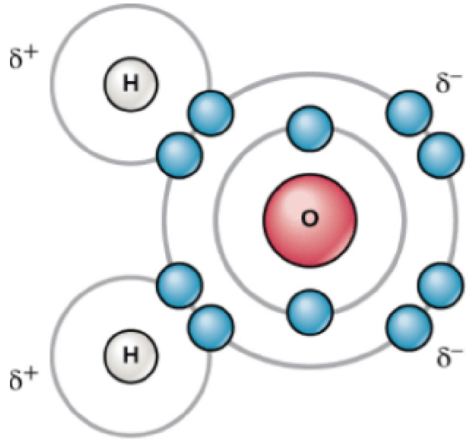
Technically ionic bonds are completely polar bonds, so the terminology can be confusing.



A dipole is a **molecule that has split charge**. Dipoles may form associations with other dipoles, induced dipoles or ions. An important type of dipole-dipole forces are hydrogen bonds.

POLAR COVALENT BOND

A polar bond refers to a type of covalent bond where electrons aren't equally shared and electronegativity values are slightly different. Polar covalent bonds form between atoms with an electronegativity difference between 0.4 and 1.7.

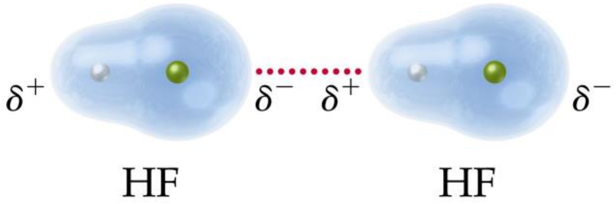


Polar covalent bond

Examples of Molecules with Polar Covalent Bonds

Water (H_2O) is a polar bonded molecule.
The electronegativity value of oxygen is 3.44, while the electronegativity of hydrogen is 2.20.

The inequality in electron distribution accounts for the bent shape of the molecule. The oxygen "side" of the molecule has a net negative charge, while the two hydrogen atoms (on the other "side") have a net positive charge.

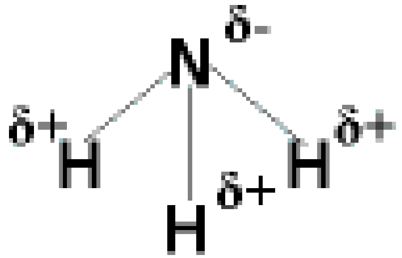


Hydrogen fluoride (HF) is another example of a molecule that has a polar covalent bond.

POLAR COVALENT BOND

Fluorine is the more electronegative atom, so the electrons in the bond are more closely associated with the fluorine atom than with the hydrogen atom.

A dipole forms with the fluorine side having a net negative charge and the hydrogen side having a net positive charge. Hydrogen fluoride is a linear molecule because there are only two atoms, so no other geometry is possible.



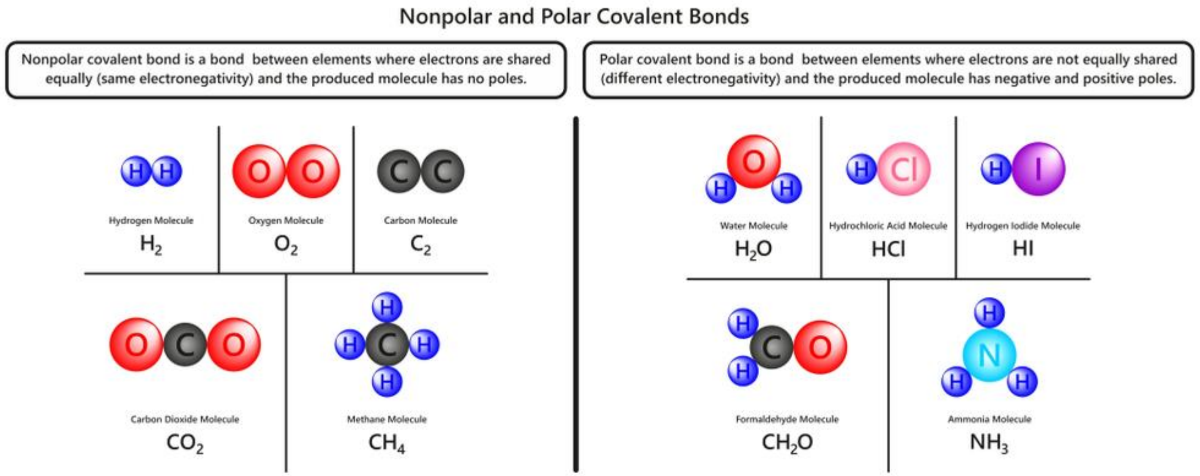
The ammonia molecule (NH_3) has polar covalent bonds between the nitrogen and hydrogen atoms. The dipole is such that the nitrogen atom is more negatively charged, with the three hydrogen atoms all on one side of the nitrogen atom with a positive charge.

POLAR COVALENT BOND

<u>Which Elements Form Polar Bonds?</u>

Polar covalent bonds form between two nonmetal atoms that have sufficiently different electronegativities from each other. Because the electronegativity values are slightly different, the bonding electron pair isn't equally shared between the atoms. For example, polar covalent bonds typically form between hydrogen and any other nonmetal.

The electronegativity value between metals and nonmetals is large, so they form ionic bonds with each other.

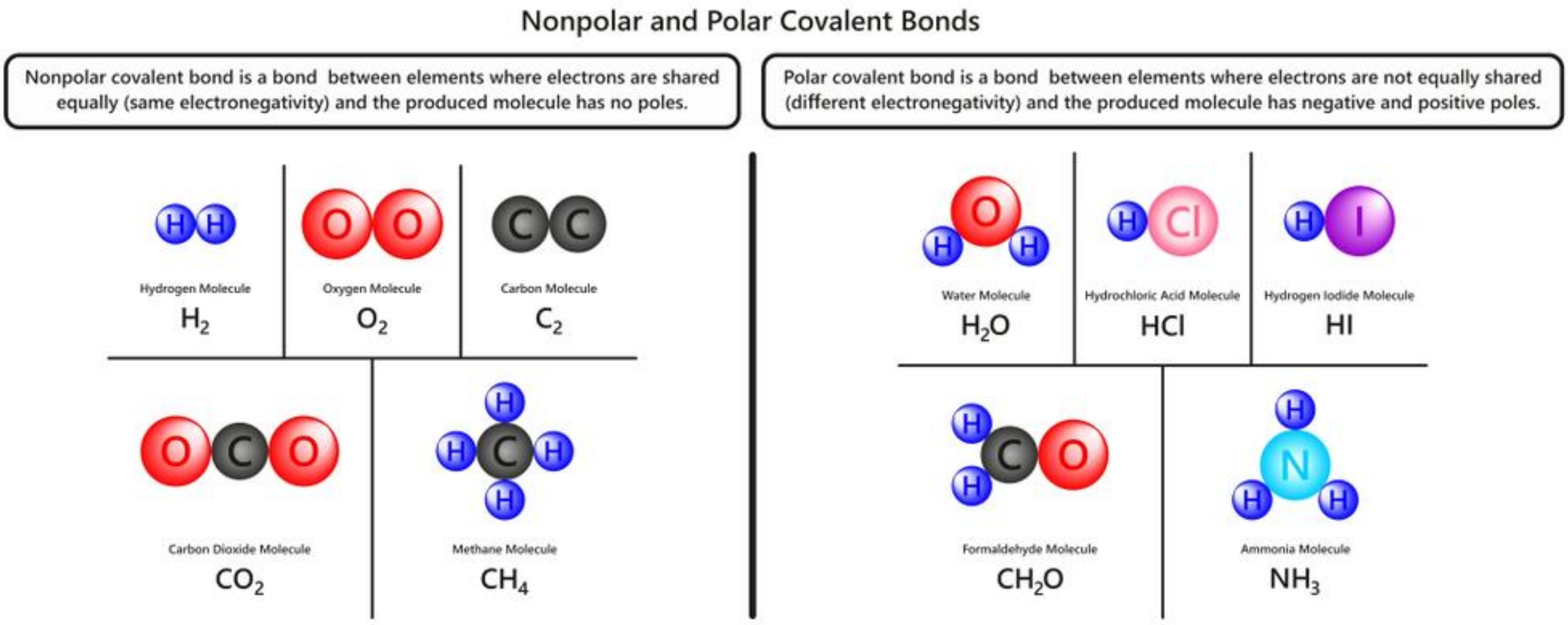


Nonpolar and Polar Covalent Bonds

Nonpolar covalent bond is a bond between elements where electrons are shared equally (same electronegativity) and the produced molecule has no poles.

Polar covalent bond is a bond between elements where electrons are not equally shared (different electronegativity) and the produced molecule has negative and positive poles.

POLAR COVALENT BOND



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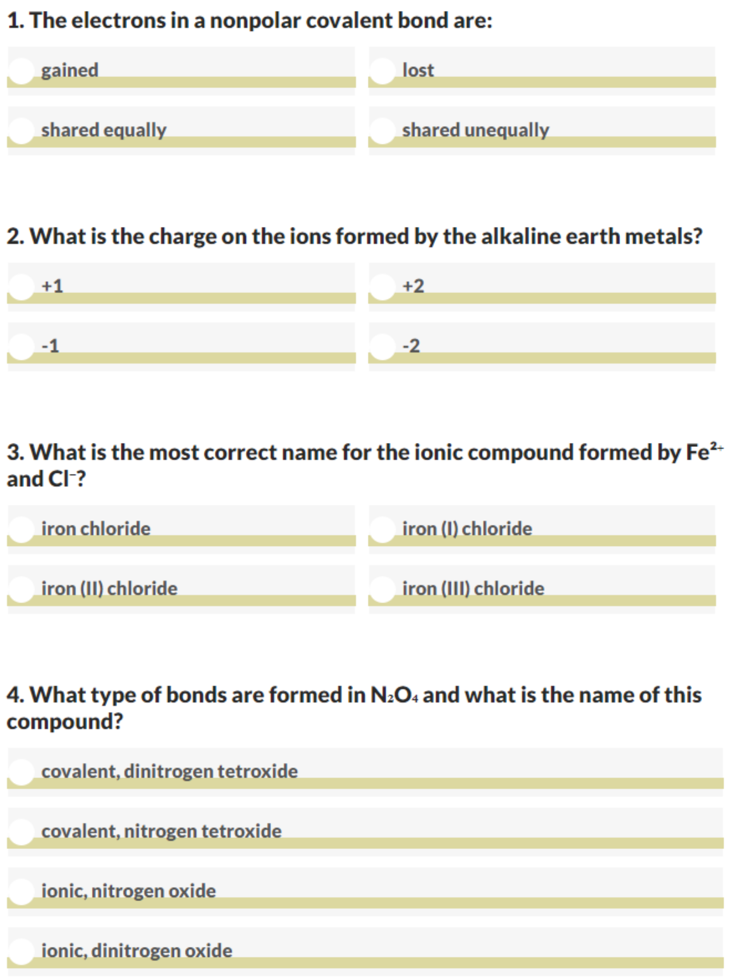
IONIC VS COVALENT BONDS

QUIZ!



1. The electrons in a nonpolar covalent bond are:

○ gained
○ lost
○ shared equally
○ shared unequally

2. What is the charge on the ions formed by the alkaline earth metals?

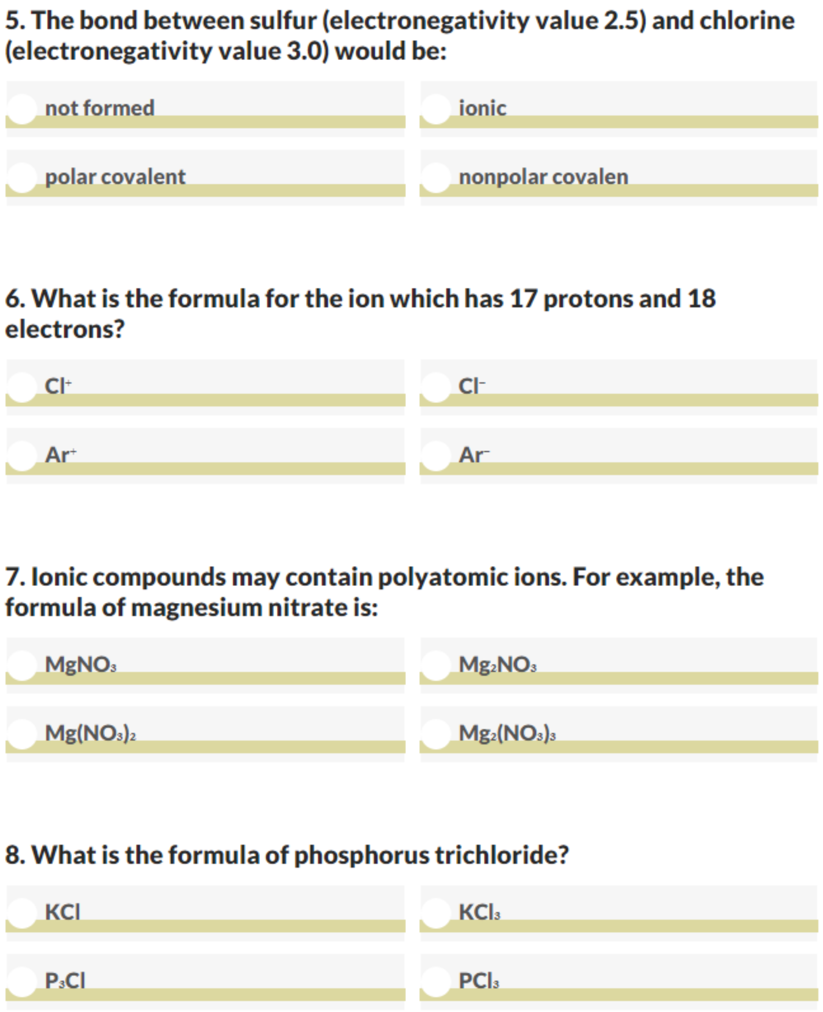
○ +1
○ +2
○ -1
○ -2

3. What is the most correct name for the ionic compound formed by Fe^{2+} and Cl^{-}?

○ iron chloride
○ iron (I) chloride
○ iron (II) chloride
○ iron (III) chloride

4. What type of bonds are formed in N_2O_4 and what is the name of this compound?

○ covalent, dinitrogen tetroxide
○ covalent, nitrogen tetroxide
○ ionic, nitrogen oxide
○ ionic, dinitrogen oxide



5. The bond between sulfur (electronegativity value 2.5) and chlorine (electronegativity value 3.0) would be:

○ not formed
○ ionic
○ polar covalent
○ nonpolar covalen

6. What is the formula for the ion which has 17 protons and 18 electrons?

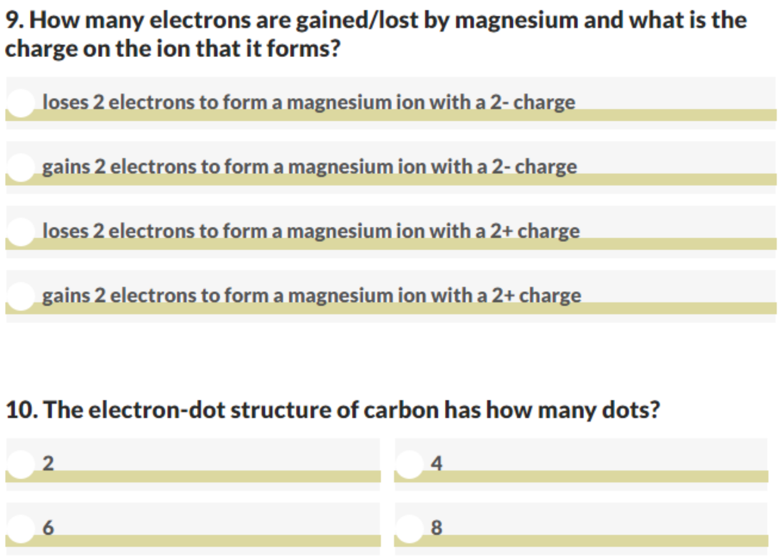
○ Cl^{+}
○ Cl^{-}
○ Ar^{+}
○ Ar^{-}

7. Ionic compounds may contain polyatomic ions. For example, the formula of magnesium nitrate is:

○ $MgNO_3$
○ Mg_2NO_3
○ $Mg(NO_3)_2$
○ $Mg_2(NO_3)_3$

8. What is the formula of phosphorus trichloride?

○ KCl
○ KCl_3
○ P_3Cl
○ PCl_3



9. How many electrons are gained/lost by magnesium and what is the charge on the ion that it forms?

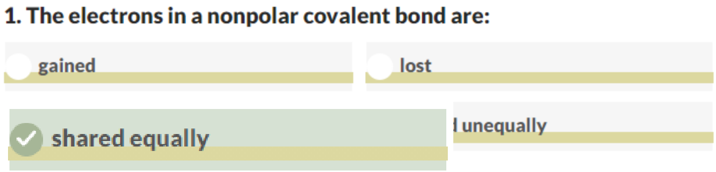
○ loses 2 electrons to form a magnesium ion with a 2- charge
○ gains 2 electrons to form a magnesium ion with a 2- charge
○ loses 2 electrons to form a magnesium ion with a 2+ charge
○ gains 2 electrons to form a magnesium ion with a 2+ charge

10. The electron-dot structure of carbon has how many dots?

○ 2
○ 4
○ 6
○ 8

IONIC VS COVALENT BONDS

QUIZ!



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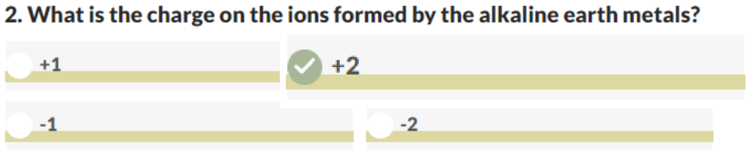
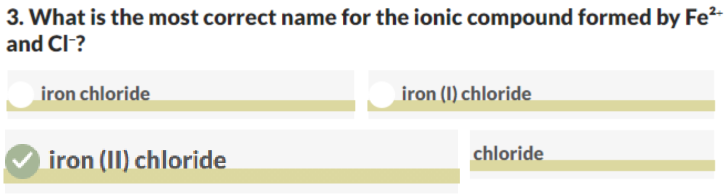


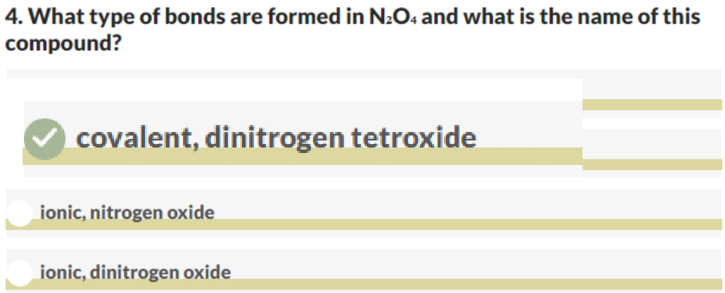
2. What is the charge on the ions formed by the alkaline earth metals?

○ +1
✓ +2
○ -1
○ -2



3. What is the most correct name for the ionic compound formed by Fe^{2+} and Cl^-?

○ iron chloride
○ iron (I) chloride
✓ iron (II) chloride
chloride



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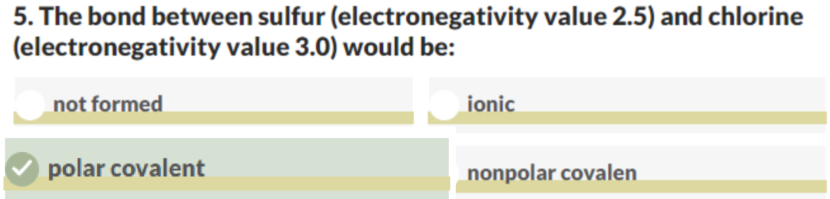
○ ionic, nitrogen oxide
○ ionic, dinitrogen oxide



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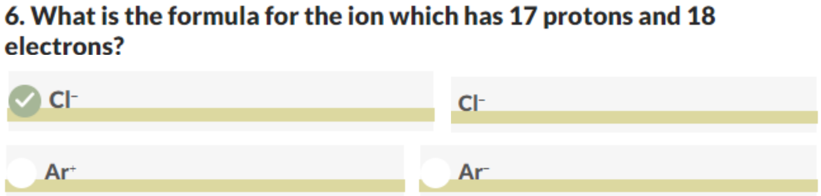
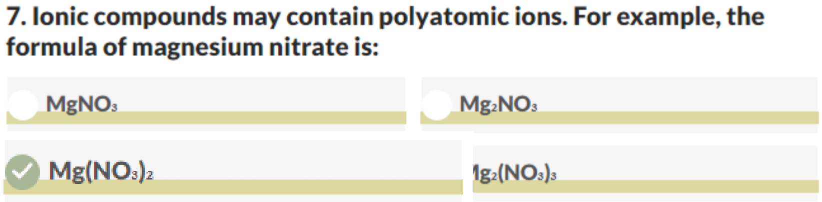


6. What is the formula for the ion which has 17 protons and 18 electrons?

✓ Cl^-
Cl^-
○ Ar^+
○ Ar^-



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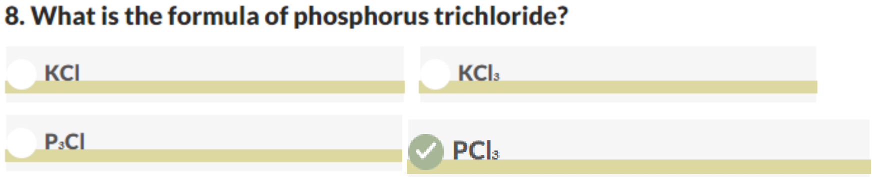

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○ KCl
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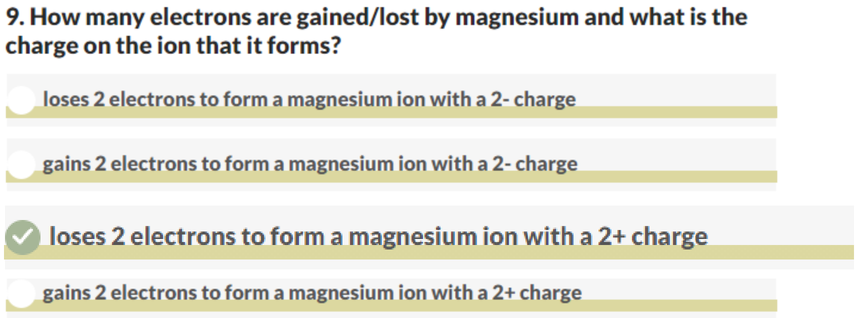
○ P_3Cl
✓ PCl_3



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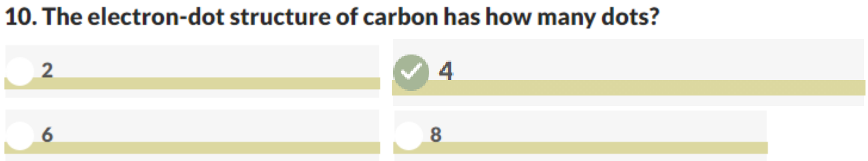


10. The electron-dot structure of carbon has how many dots?

○ 2
✓ 4
○ 6
○ 8

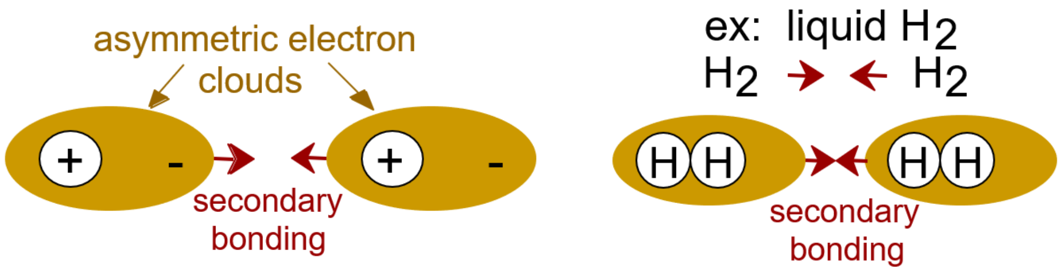
[Ionic vs Covalent Bonds - Understand the Difference (thoughtco.com)]

SECONDARY BONDING - VANDER WAALS BONDING

Secondary bonds, or **van der Waals** (physical) **bonds,** are weak in comparison to the primary or chemical bonds; bonding energies range between about 4 and 30 kJ/mol.

Secondary bonding exists between virtually all atoms or molecules, but its presence may be obscured if any of the three primary bonding types is present.

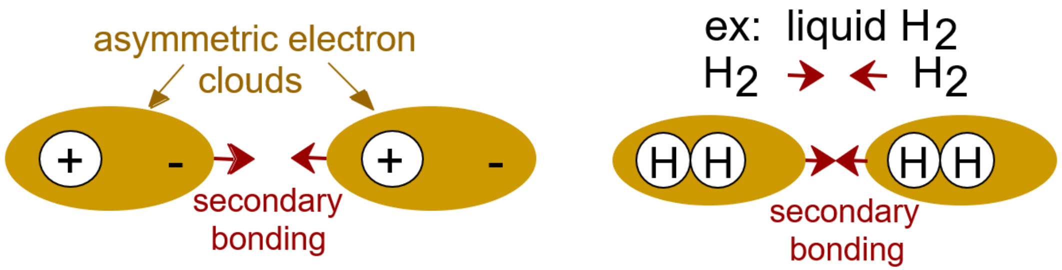
Secondary bonding is evidenced for the inert gases, which have stable electron structures. In addition, secondary (or *inter*molecular) bonds are possible between atoms or groups of atoms, which themselves are joined together by primary (or *intra*molecular) ionic or covalent bonds.



SECONDARY BONDING - VANDER WAALS BONDING

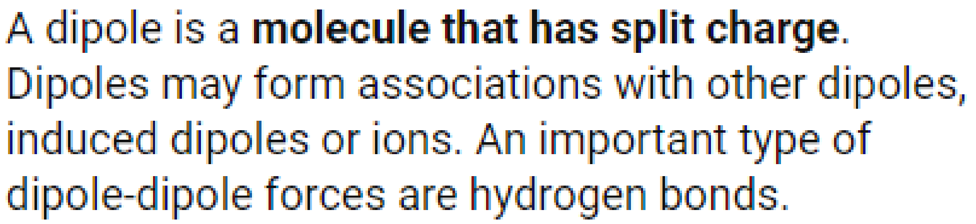
Secondary bonding forces arise from atomic or molecular **dipoles.** In essence, an electric dipole exists whenever there is some separation of positive and negative portions of an atom or molecule.

The bonding results from the coulombic attraction between the positive end of one dipole and the negative region of an adjacent one.

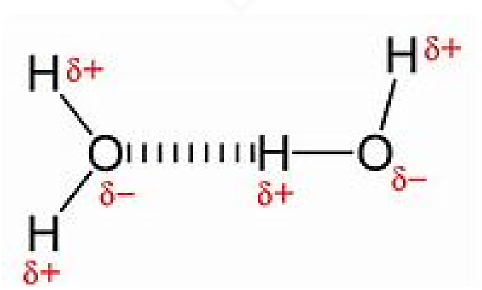


SECONDARY BONDING - VANDER WAALS BONDING

Dipole interactions occur between induced dipoles, between induced dipoles and polar molecules (which have permanent dipoles), and between polar molecules. **Hydrogen bonding,** a special type of secondary bonding, is found to exist between some molecules that have hydrogen as one of the constituents.



A dipole is a **molecule that has split charge**. Dipoles may form associations with other dipoles, induced dipoles or ions. An important type of dipole-dipole forces are hydrogen bonds.

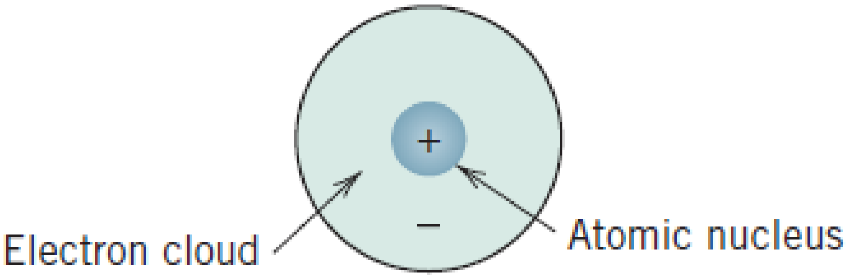


SECONDARY BONDING - VANDER WAALS BONDING

Fluctuating Induced Dipole Bonds

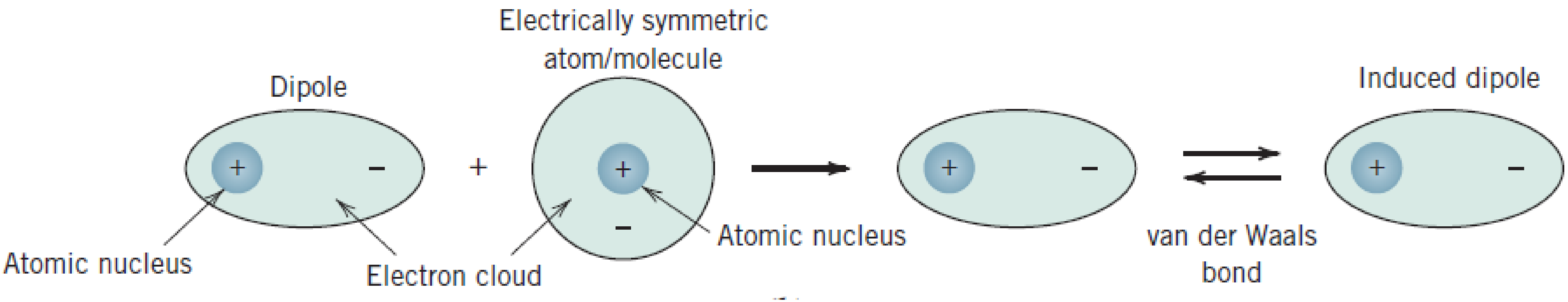
A dipole may be created or induced in an atom or molecule that is normally electrically symmetric— that is, the overall spatial distribution of the electrons is symmetric with respect to the positively charged nucleus.

All atoms experience constant vibrational motion that can cause instantaneous and short lived distortions of this electrical symmetry for some of the atoms or molecules and the creation of small electric dipoles.



SECONDARY BONDING - VANDER WAALS BONDING

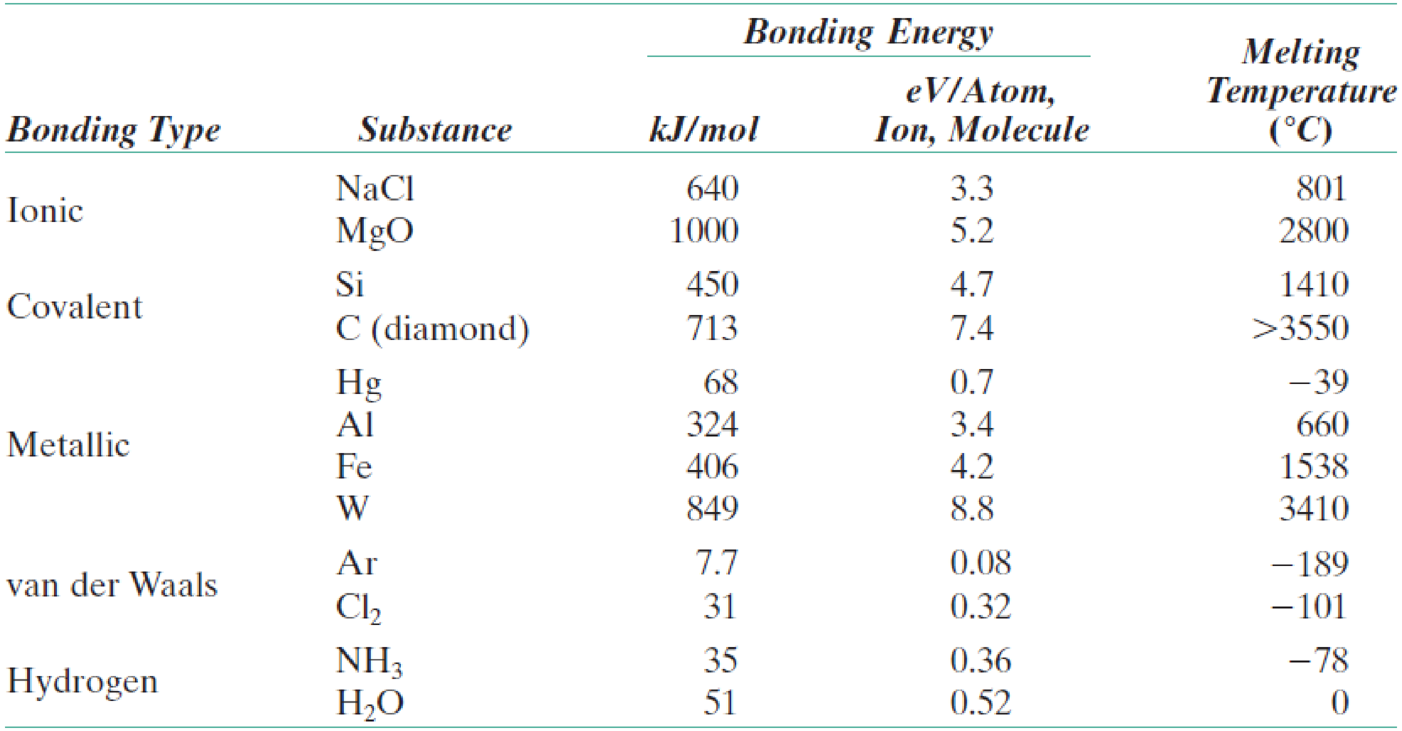
One of these dipoles can in turn produce a displacement of the electron distribution of an adjacent molecule or atom, which induces the second one also to become a dipole that is then weakly attracted or bonded to the first; this is one type of van der Waals bonding.



SECONDARY BONDING - VANDER WAALS BONDING

These attractive forces, which are temporary and fluctuate with time, may exist between large numbers of atoms or molecules.

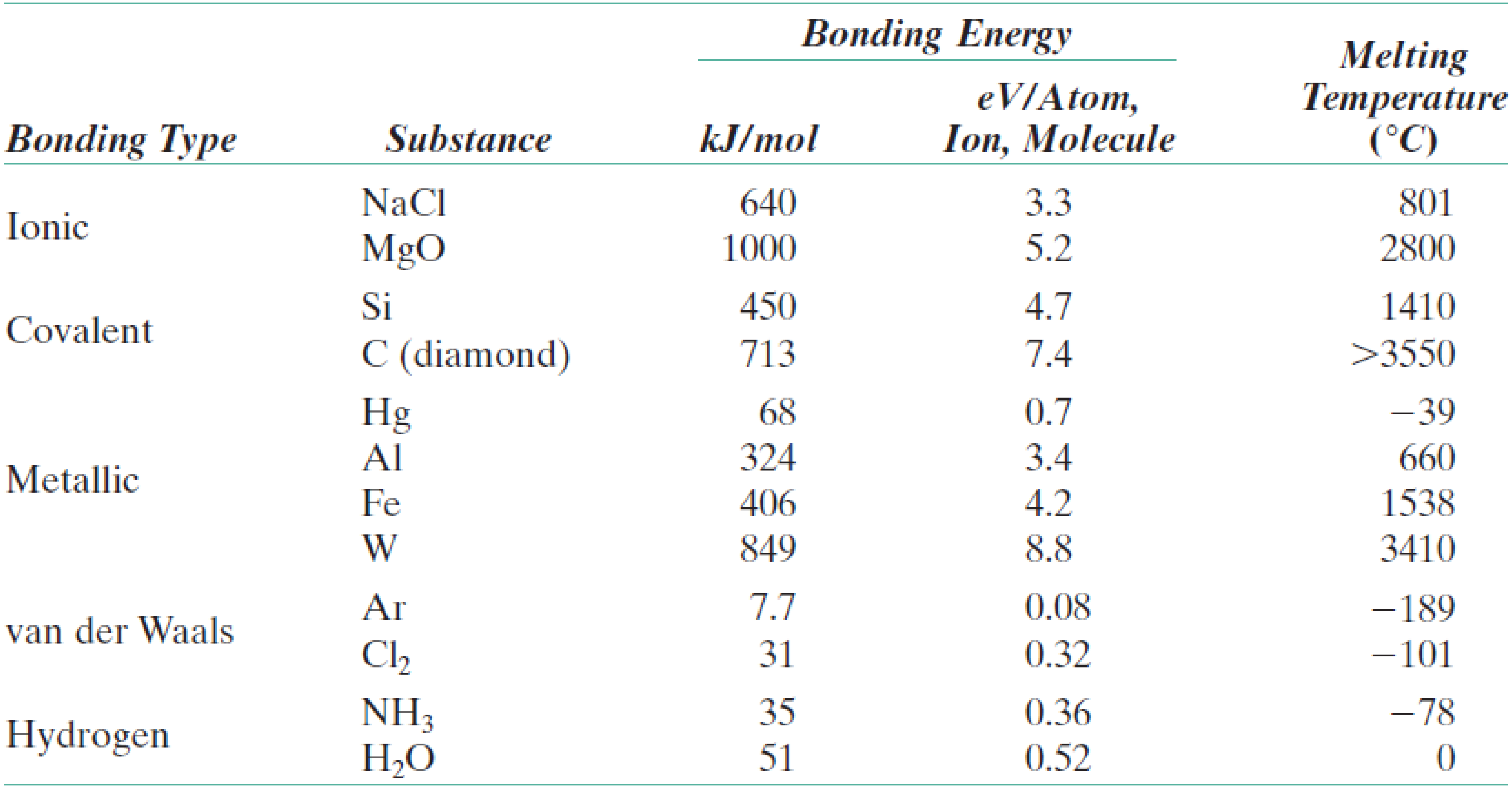
The liquefaction and, in some cases, the solidification of the inert gases and other electrically neutral and symmetric molecules such as H_2 and Cl_2 are realized because of this type of bonding.



		Bonding Energy		*Melting Temperature (°C)*
Bonding Type	***Substance***	***kJ/mol***	***eV/Atom, Ion, Molecule***	
Ionic	NaCl	640	3.3	801
	MgO	1000	5.2	2800
Covalent	Si	450	4.7	1410
	C (diamond)	713	7.4	>3550
Metallic	Hg	68	0.7	−39
	Al	324	3.4	660
	Fe	406	4.2	1538
	W	849	8.8	3410
van der Waals	Ar	7.7	0.08	−189
	Cl_2	31	0.32	−101
Hydrogen	NH_3	35	0.36	−78
	H_2O	51	0.52	0

SECONDARY BONDING - VANDER WAALS BONDING

Melting and boiling temperatures are extremely low in materials for which induced dipole bonding predominates; of all possible intermolecular bonds, these are the weakest.



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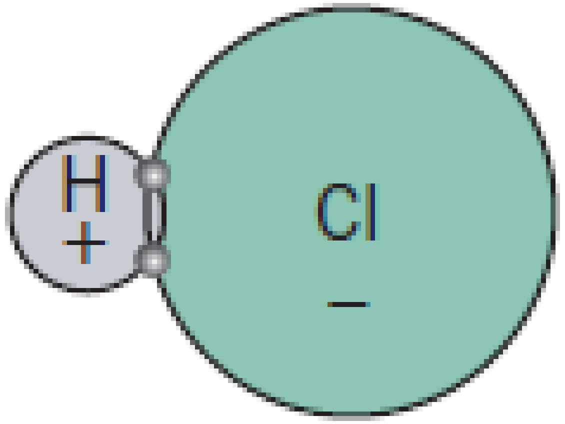
SECONDARY BONDING - VANDER WAALS BONDING

Polar Molecule–Induced Dipole Bonds

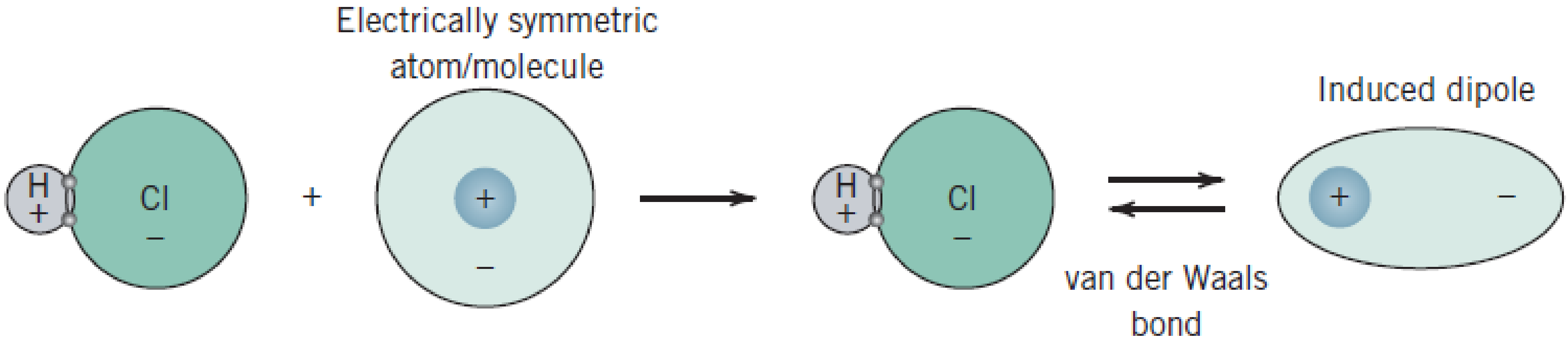
Permanent dipole moments exist in some molecules by virtue of an asymmetrical arrangement of positively and negatively charged regions; such molecules are termed **polar molecules.**

The figure shows a schematic representation of a hydrogen chloride molecule; a permanent dipole moment arises from net positive and negative charges that are respectively associated with the hydrogen and chlorine ends of the HCl molecule.



SECONDARY BONDING - VANDER WAALS BONDING

Polar molecules can also induce dipoles in adjacent nonpolar molecules, and a bond forms as a result of attractive forces between the two molecules; this bonding scheme is represented schematically.

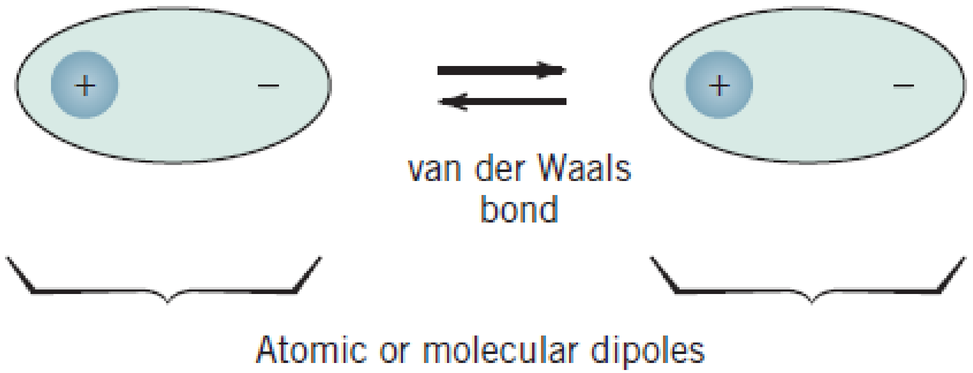


Furthermore, the magnitude of this bond is greater than for fluctuating induced dipoles.

SECONDARY BONDING - VANDER WAALS BONDING

Permanent Dipole Bonds

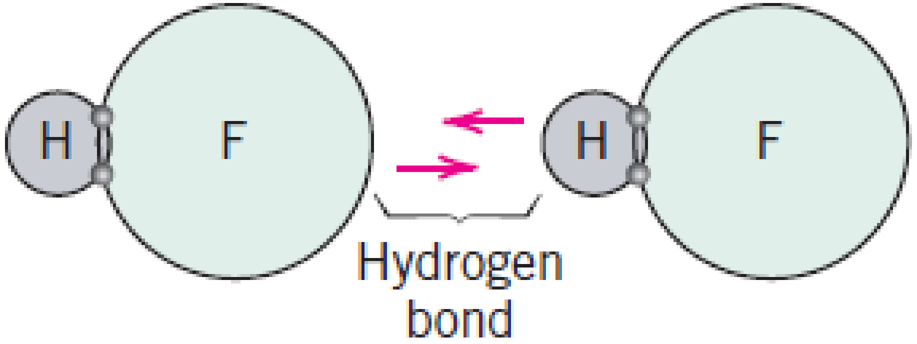
Coulombic forces also exist between adjacent polar molecules. The associated bonding energies are significantly greater than for bonds involving induced dipoles.



The strongest secondary bonding type, the hydrogen bond, is a special case of polar molecule bonding. It occurs between molecules in which hydrogen is covalently bonded to fluorine (as in HF), oxygen (as in H_2O), or nitrogen (as in NH_3). For each H-F, H-O, or H-N bond, the single hydrogen electron is shared with the other atom.

SECONDARY BONDING - VANDER WAALS BONDING

Thus, the hydrogen end of the bond is essentially a positively charged bare proton unscreened by any electrons. This highly positively charged end of the molecule is capable of a strong attractive force with the negative end of an adjacent molecule, as demonstrated for HF.



In essence, this single proton forms a bridge between two negatively charged atoms. The magnitude of the hydrogen bond is generally greater than that of the other types of secondary bonds and may be as high as 51 kJ/mol. Melting and boiling temperatures for hydrogen fluoride, ammonia, and water are abnormally high in light of their low molecular weights, as a consequence of hydrogen bonding.

SECONDARY BONDING - VANDER WAALS BONDING

In spite of the small energies associated with secondary bonds, they nevertheless are involved in a number of natural phenomena and many products that we use on a daily basis.

Examples of physical phenomena include the *solubility* of one substance in another, surface tension and capillary action, *vapor pressure*, *volatility*, and *viscosity*.

Common applications that make use of these phenomena include *adhesives*—van der Waals bonds form between two surfaces so that they adhere to one another; *surfactants*—compounds that lower the surface tension of a liquid and are found in soaps, detergents, and foaming agents.

SECONDARY BONDING - VANDER WAALS BONDING

Emulsifiers—substances that, when added to two immiscible materials (usually liquids), allow particles of one material to be suspended in another (common emulsions include sunscreens, salad dressings, milk, and mayonnaise); and *desiccants*—materials that form hydrogen bonds with water molecules (and remove moisture from closed containers—e.g., small packets that are often found in cartons of packaged goods); and finally, the strengths, stiffnesses, and softening temperatures of polymers, to some degree, depend on secondary bonds that form between chain molecules.

Food for thought:

- What happens when you put a plastic container filled with water in a freezer? Explain!
- Why a bottle of water bought from your local shop isn’t suitable to take onboard an airplane?