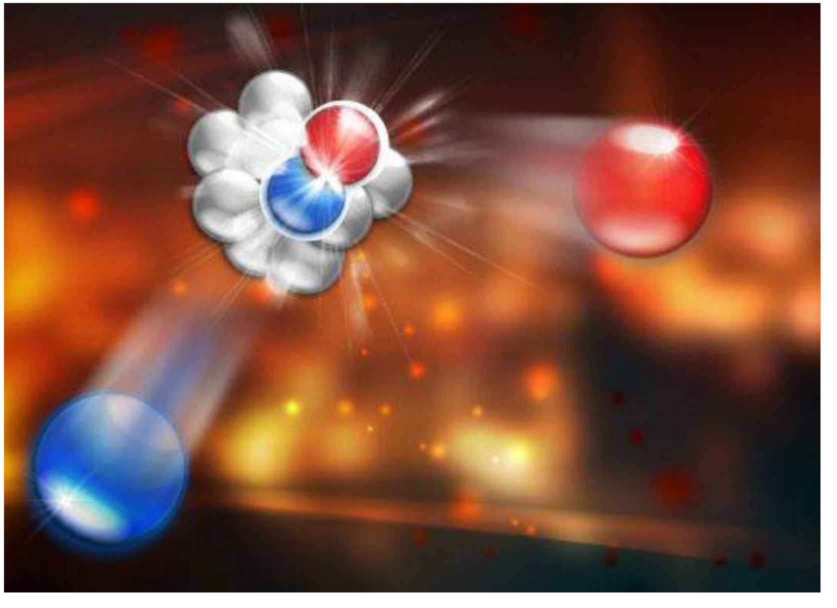
CHAPTER 2

ATOMIC STRUCTURE AND INTERATOMIC BONDING



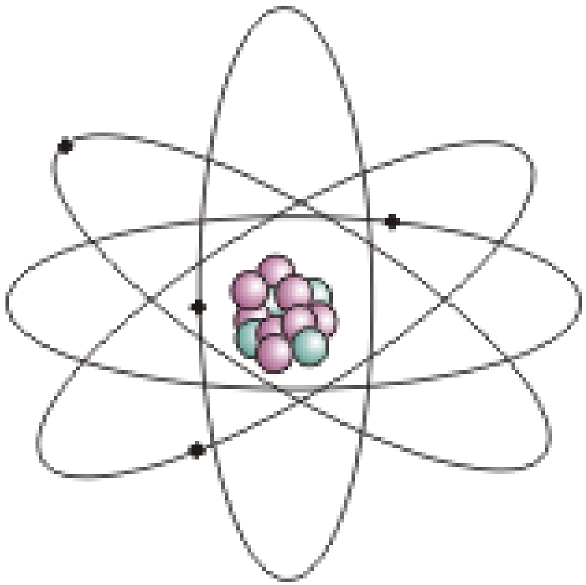
What are ATOMS?

- All matter is made up of tiny particles called atoms.
- Since the atom is too small to be seen even with the most powerful microscopes, scientists rely upon on models to help us to understand the atom.

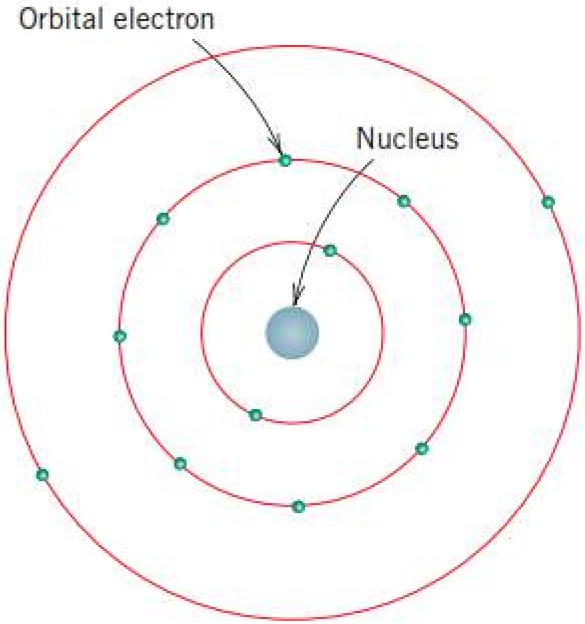
Even with the world's best microscopes we cannot clearly see the structure or behavior of the atom.

Is this really an ATOM?

Even though we do not know what an atom looks like, scientific models must be based on evidence. Many of the atom models that you have seen may look like the one below which shows the parts and structure of the atom.



Bohr Atomic Model



- Bohr model present early attempt to describe electron in atom
- Electrons are particles moving in discrete orbitals
- Electron energy is quantized into shells

ELECTRON ENERGY STATES

Bohr's model vs Wave mechanical's models

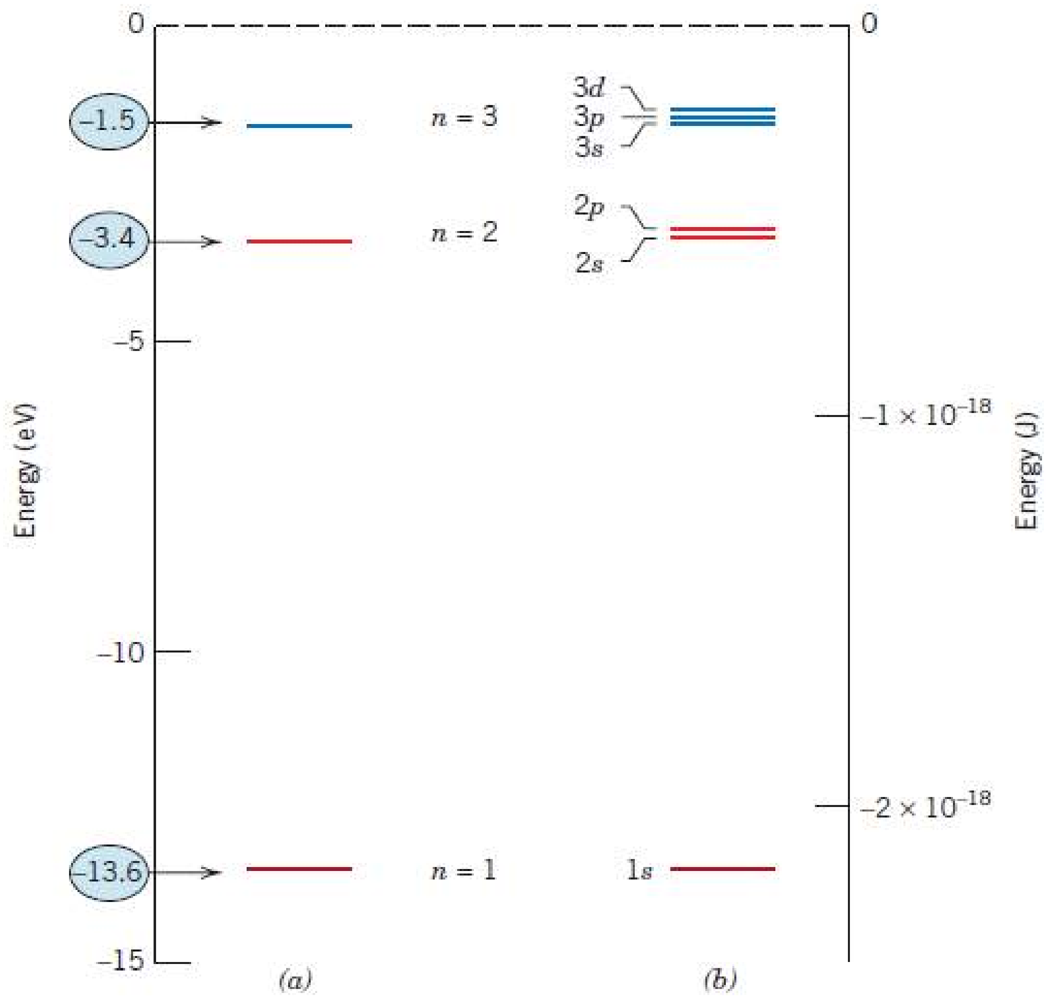


Figure 2.2 (*a*) The first three electron energy states for the Bohr hydrogen atom. (*b*) Electron energy states for the first three shells of the wave-mechanical hydrogen atom. (Adapted from W. G. Moffatt, G. W. Pearsall, and J. Wulff, *The Structure and Properties of Materials,* Vol. I, *Structure,* p. 10. Copyright © 1964 by John Wiley & Sons, New York. Reprinted by permission of John Wiley & Sons, Inc.)

Disadvantages of Bohr's Model

1. Electron couldn't circle around the nucleus like a planet – because they would lose energy (by emitting electromagnetic radiation & spiral to the nucleus)
2. Bohr was not able to explain electron orbits of large atoms with many electrons

What does an ATOM look like?

Atoms are made of a nucleus that contains *protons*, *neutrons* and *electrons* that orbit around the nucleus at different levels, known as shells.

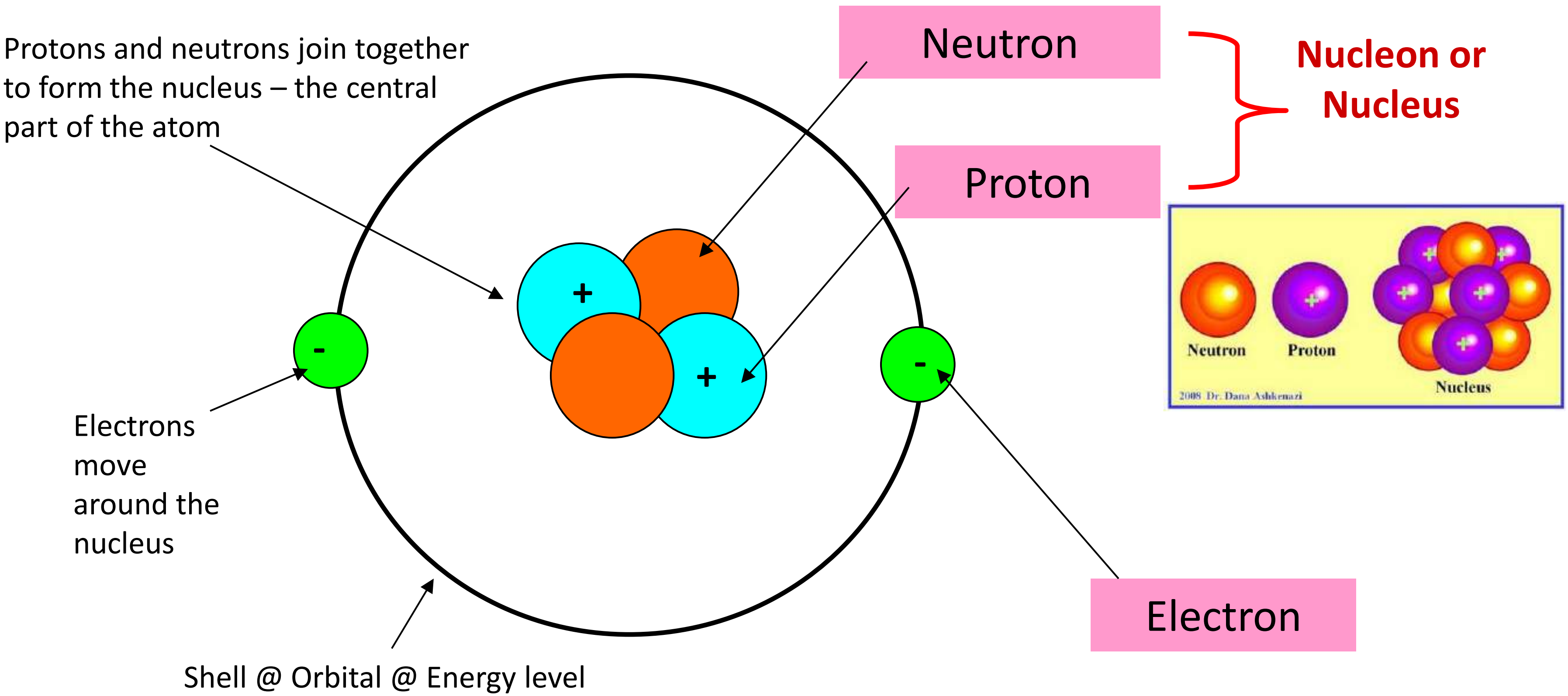


Figure : A simplified diagram of atom

Atomic Structure

- atom – electrons – 9.11×10^{-31} kg
  protons
  neutrons } $1.672623 / 1.674929 \times 10^{-27}$ kg
- atomic number = # of protons in nucleus of atom
  = # of electrons in neutral species

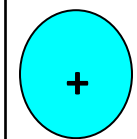
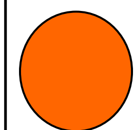
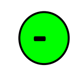
- A [=] atomic mass unit = amu = 1/12 mass of ^{12}C
- 1 amu $=1.660540 \times 10^{-27}$ kg

Atomic wt = wt of 6.022×10^{23} molecules or atoms

1 amu/atom = 1 g/mol

C 12.011
H 1.008 etc.

- These particles have the following properties:

Particle	Charge	Location	Mass (amu)	Symbol
Proton	Positive (+ve)	Nucleus	1.0073	
Neutron	Neutral	Nucleus	1.0087	
Electron	Negative (-ve)	Orbital	0.000549	

- *To describe the mass of atom, a unit of mass called the atomic mass unit (amu) is used.*

- The number of protons, neutrons and electrons in an atom completely determine its properties and identity. This is what makes one atom different from another.

Why are all ATOMS are ELECTRICALLY NEUTRAL?

Most atoms are electrically neutral, meaning that they have an **equal number of protons and electrons**. The positive and negative charges cancel each other out. Therefore, the atom is said to be electrically neutral.

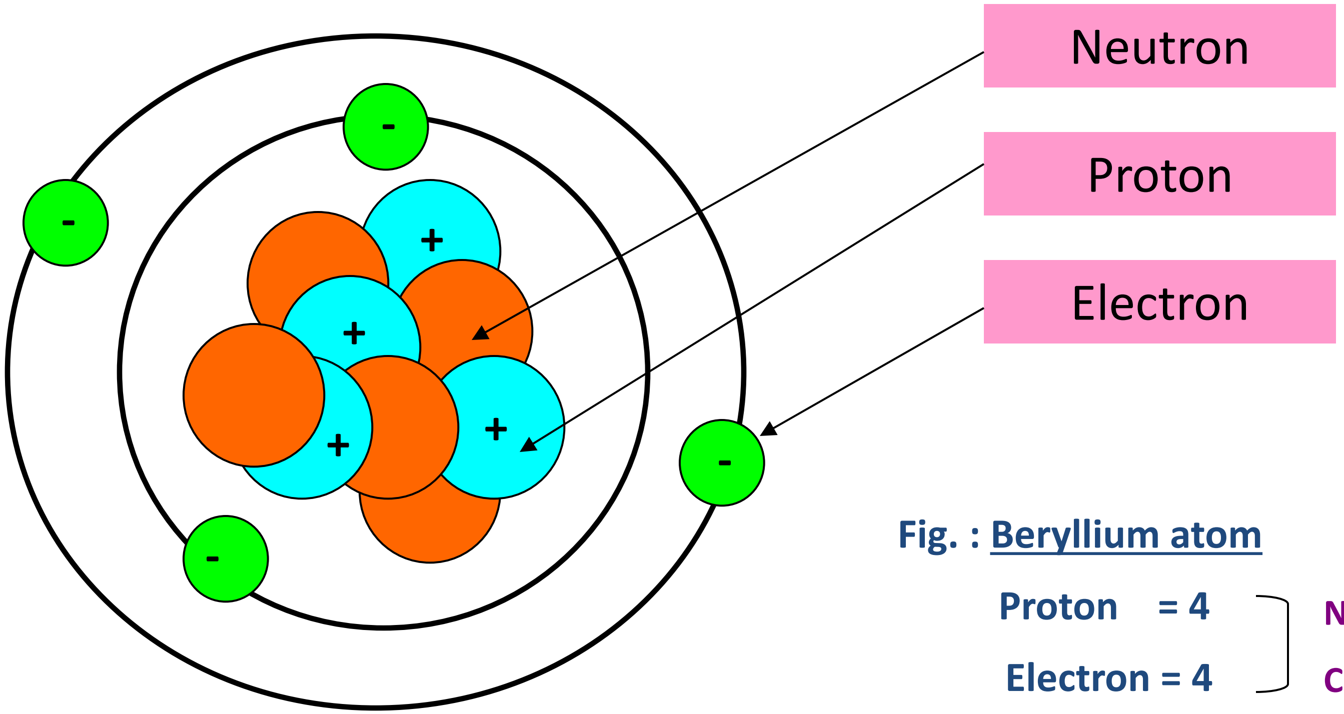


Fig. : Beryllium atom

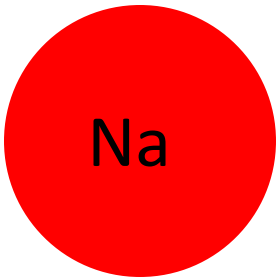
Proton = 4
Electron = 4

NEUTRAL CHARGE

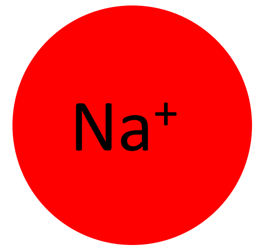
If an atom gains or loses electrons, the atom is no longer neutral and it become **electrically charged** . The atom is then called an ***ION*.**

cation **- ion with a positive charge**

- If a neutral atom loses one or more electrons, it becomes a cation.



11 protons
11 electrons

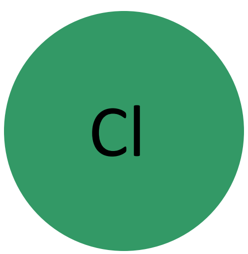


11 protons
10 electrons

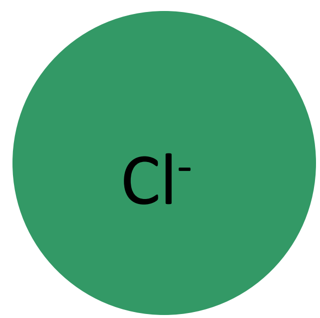
Cations are smaller than their "parent atom" because there is less e-e repulsion

anion **- ion with a negative charge**

- If a neutral atom gains one or more electrons, it becomes an anion.



17 protons
17 electrons



17 protons
18 electrons

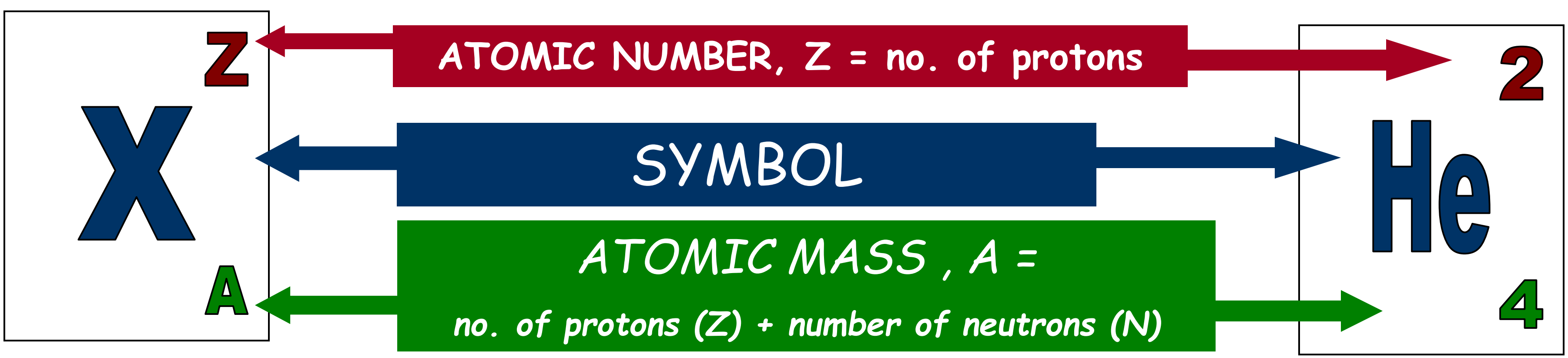
Anions are larger than their "parent atom" because there is more e-- e repulsion

ATOMIC NUMBER and ATOMIC MASS

Atom can be described using :

1) ATOMIC NUMBER

2) ATOMIC MASS



The element helium has the atomic number 2, is represented by the symbol He, its atomic mass is 4 and its name is helium.

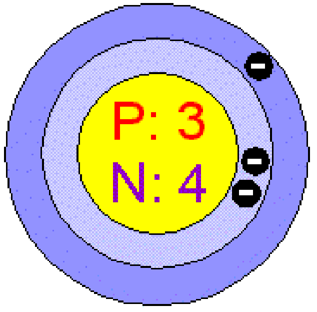
	II A	III B	IV B	V B	VI B	VII B	VIII			I B	II B	III A	IV A	V A	VI A	VII A	
1 H 1.0079																	2 He 4.00260
3 Li 6.94	4 Be 9.01218											5 B 10.811	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.998403	10 Ne 20.179
11 Na 22.98977	12 Mg 24.305											13 Al 26.98154	14 Si 28.0855	15 P 30.97376	16 S 32.066	17 Cl 35.453	18 Ar 39.948
19 K 39.0983	20 Ca 40.08	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.996	25 Mn 55.9381	26 Fe 55.847	27 Co 58.9332	28 Ni 58.69	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc 98.9072	44 Ru 101.07	45 Rh 102.9055	46 Pd 106.42	47 Ag 107.868	48 Cd 112.41	49 In 114.82	50 Sn 118.710	51 Sb 121.75	52 Te 127.60	53 I 126.9047	54 Xe 131.30
55 Cs 132.9054	56 Ba 137.33	57 La 138.33	72 Hf 178.49	73 Ta 180.9479	74 W 183.85	75 Re 186.207	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.9665	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.9804	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226.0254)	89 Ac (227)															

58 Ce 140.12	59 Pr 140.9077	60 Nd 144.24	61 Pm (145)	62 Sm 150.4	63 Eu 151.965	64 Gd 157.25	65 Tb 158.9254	66 Dy 162.50	67 Ho 164.9303	68 Er 167.26	69 Tm 168.9342	70 Yb 173.04	71 Lu 174.967
90 Th 232.0381	91 Pa (231.036)	92 U 238.029	93 Np 237.0482	94 Pu (244.069)	95 Am (243.06)	96 Cm (247.070)	97 Bk (247.070)	98 Cf (251.08)	99 Es (252.083)	100 Fm (257.095)	101 Md (258.10)	102 No (259.101)	103 Lw (260.11)

ATOMIC NUMBER tells how many ***PROTONS (Z)*** are in its atoms which determine the ***atom's identity.***

The list of elements (ranked according to an increasing no. of protons) can be looked up on the Periodic Table. So, if an atom has 2 protons (atomic no. = 2), it must be helium(He).

ATOMIC MASS tells the sum of the masses of *PROTONS (Z)* and *NEUTRONS (N)*. within the nucleus E.g :



lithium:

Atomic number = 3
3 protons, Z
4 neutrons, N
Atomic mass, A = 3 + 4 = 7

BUT... although each element has a defined number of protons, the number of neutrons is not fixed → ***isotopes***

ISOTOPES

- Atoms which have the *same number of protons* but *different numbers of neutrons.*

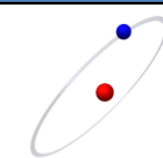
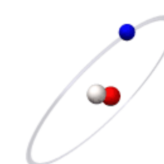
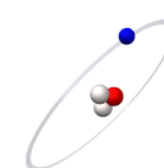
Different mass number

$^{1}_{1}H$ $\quad$ $^{2}_{1}H$ (D) $\quad$ $^{3}_{1}H$ (T)

Same atomic no. @ no. of protons

- Atoms which have the *same atomic number* but *different atomic mass .*

- Eg : Hydrogen has 3 isotopes.

Natural Isotope	Proton	Neutron	Atomic Mass	
Hydrogen 1 (hydrogen)	1	0	1	
Hydrogen 2 (deuterium)	1	1	2	
Hydrogen 3 (tritium)	1	2	3	

Element	Name	Proton Number	Nucleon Number	Number of proton	Number of neutron
Hydrogen	Hydrogen	1	1	1	0
	Deuterium	1	12	1	1
	Tritium	1	23	1	2
Oxygen	Oxygen-16	8	16	8	8
	Oxygen-17	8	17	8	9
	Oxygen-18	8	18	8	10
Carbon	Carbon-12	6	12	6	6
	Carbon-13	6	13	6	7
	Carbon-14	6	14	6	8
Chlorine	Chlorine-35	17	35	17	18
	Chlorine-37	17	37	17	20
Sodium	Sodium-23	11	23	11	12
	Sodium-24	11	24	11	13

Atomic Weight

- Corresponds to the weighted average of the atomic masses of the atom's naturally occurring isotopes.

1 **H** 1.0079	II A											III A	IV A	V A	VI A	VII A	2 **He** 4.00260
3 **Li** 6.94	4 **Be** 9.01218											5 **B** 10.811	6 **C** 12.011	7 **N** 14.0067	8 **O** 15.9994	9 **F** 18.998403	10 **Ne** 20.179
11 **Na** 22.98977	12 **Mg** 24.305	III B	IV B	V B	VI B	VII B	VIII			I B	II B	13 **Al** 26.98154	14 **Si** 28.0855	15 **P** 30.97376	16 **S** 32.066	17 **Cl** 35.453	18 **Ar** 39.948
19 **K** 39.0983	20 **Ca** 40.08	21 **Sc** 44.9559	22 **Ti** 47.88	23 **V** 50.9415	24 **Cr** 51.996	25 **Mn** 54.9380	26 **Fe** 55.847	27 **Co** 58.9332	28 **Ni** 58.69	29 **Cu** 63.546	30 **Zn** 65.39	31 **Ga** 69.723	32 **Ge** 72.61	33 **As** 74.9216	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.80
37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.9059	40 **Zr** 91.224	41 **Nb** 92.9064	42 **Mo** 95.94	43 **Tc** 98.9072	44 **Ru** 101.07	45 **Rh** 102.9055	46 **Pd** 106.42	47 **Ag** 107.868	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.710	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.9045	54 **Xe** 131.30
55 **Cs** 132.9054	56 **Ba** 137.33	57 **La** 138.33	72 **Hf** 178.49	73 **Ta** 180.9479	74 **W** 183.85	75 **Re** 186.207	76 **Os** 190.2	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.9665	80 **Hg** 200.59	81 **Tl** 204.383	82 **Pb** 207.2	83 **Bi** 208.9804	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226.0254)	89 **Ac** (227)															
				58 **Ce** 140.12	59 **Pr** 140.9077	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.965	64 **Gd** 157.25	65 **Tb** 158.9254	66 **Dy** 162.50	67 **Ho** 164.9303	68 **Er** 167.26	69 **Tm** 168.9342	70 **Yb** 173.04	71 **Lu** 174.967
				90 **Th** 232.0381	91 **Pa** (231.036)	92 **U** 238.029	93 **Np** 237.0482	94 **Pu** (244.069)	95 **Am** (243.06)	96 **Cm** (247.070)	97 **Bk** (247.070)	98 **Cf** (251.08)	99 **Es** (252.083)	100 **Fm** (257.095)	101 **Md** (258.10)	102 **No** (259.101)	103 **Lw** (260.11)

ATOMIC STRUCTURE

- Some of the following properties
  1) Chemical
  2) Electrical
  3) Thermal
  4) Optical

are determined by electronic structure

QUANTUM NUMBERS

Principle quantum number, n

- Refer to electron shell

Subsidiary quantum number, l

- Refer to subshell / orbital

The magnetic quantum number, m_l

- Refer to spatial orientation of a single atomic orbital

Electron spin quantum number, m_s

- Refer to spin directions for an electron (clockwise and counterclockwise)

ELECTRONIC STRUCTURE

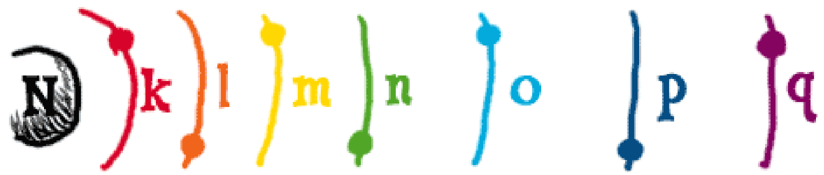
- Electrons have wavelike and particulate properties.
- Two of the wavelike characteristics are
  - electrons are in orbitals defined by a probability.
  - each orbital at discrete energy level is determined by quantum numbers.

  - <u>Quantum #</u> <u>Designation</u>

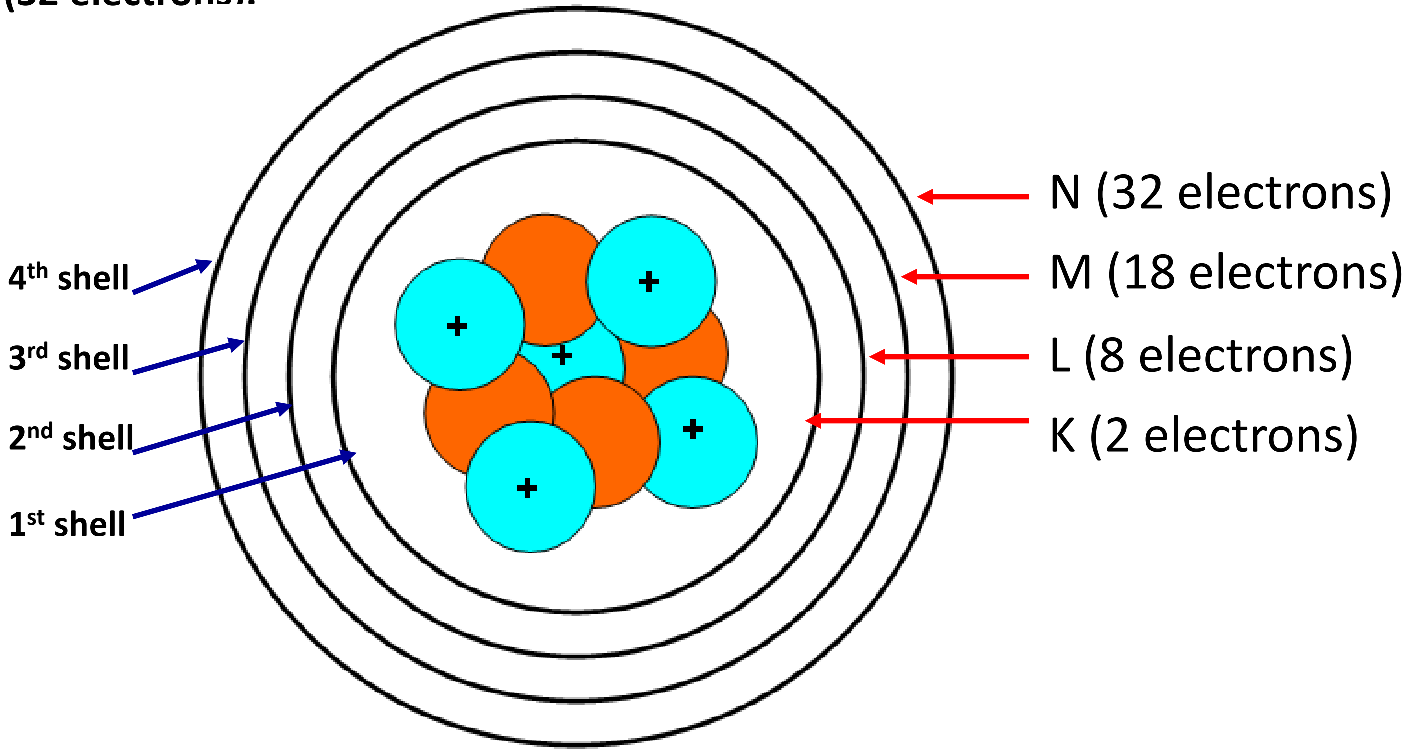
Quantum #	Designation
n = principal (energy level-shell)	*K, L, M, N, O* (1, 2, 3, etc.)
l = subsidiary (orbitals)	*s, p, d, f* (0, 1, 2, 3,..., *n*-1)
m_l = no electron state in each electron subshell	1, 3, 5, 7 (-l to +l)
m_s = spin moment on each electron	½, -½

ELECTRON SHELLS

The electron cloud that surrounded the nucleus is divided into 7 shells (a.k.a energy level) – K (1st shell, closest to nucleus) followed by L, M, N, O, P, Q.



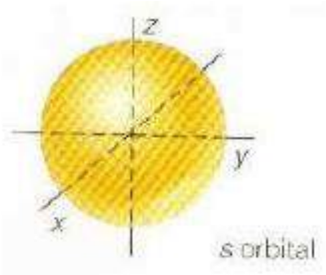
Each of the shell, hold a limited no. of electrons. E.g : K (2 electrons), L (8 electrons), M (18 electrons), N (32 electrons).



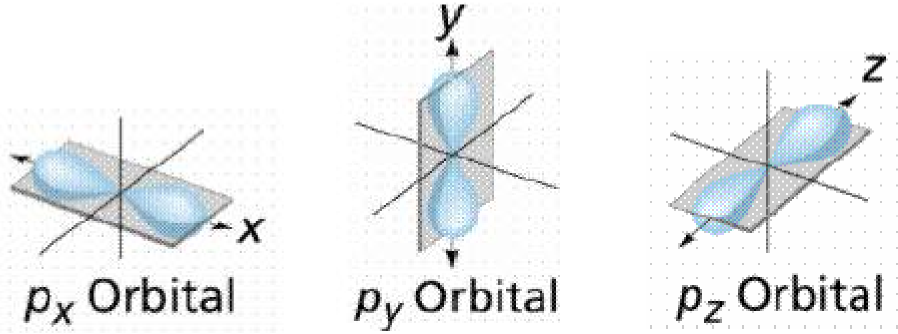
ORBITALS

- Within each shell, the electrons occupy sub shell (energy sublevels) – s, p, d, f, g, h, i. Each sub shell holds a different types of orbital.
- Each orbital holds a max. of 2 electrons.
- Each orbital has a characteristic energy state and characteristic shape.

- **s - orbital**
  –spherical shape
  –Located closest to nucleus (first energy level)
  –Max 2 electrons



- **p - orbital**
  - There is 3 distinct p - orbitals (px, py, pz)
  - Dumbbell shape
  - Second energy level
  - 6 electrons



d- orbital

- There is 5 distinct d – orbitals
- Max 10 electrons
- Third energy level

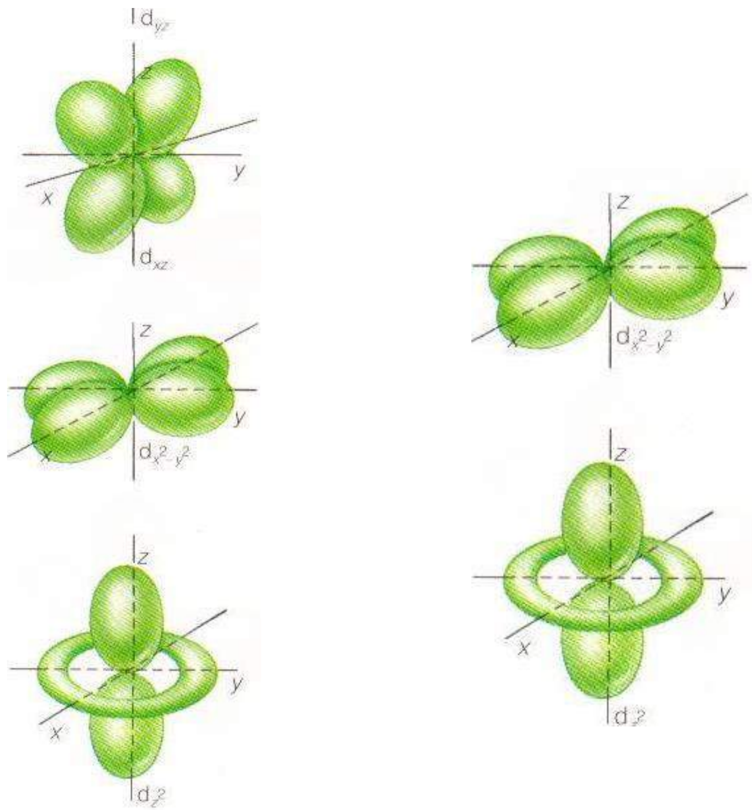


Table : The number of available electron states in some of the electrons shells and subshells.

Principal Quantum Number n	Shell Designation	Subshells	Number of States	Number of Electrons Per Subshell	Number of Electrons Per Shell
1	*K*	*s*	1	2	2
2	*L*	*s*	1	2	8
		p	3	6	
3	*M*	*s*	1	2	18
		p	3	6	
		d	5	10	
4	*N*	*s*	1	2	32
		p	3	6	
		d	5	10	
		f	7	14	

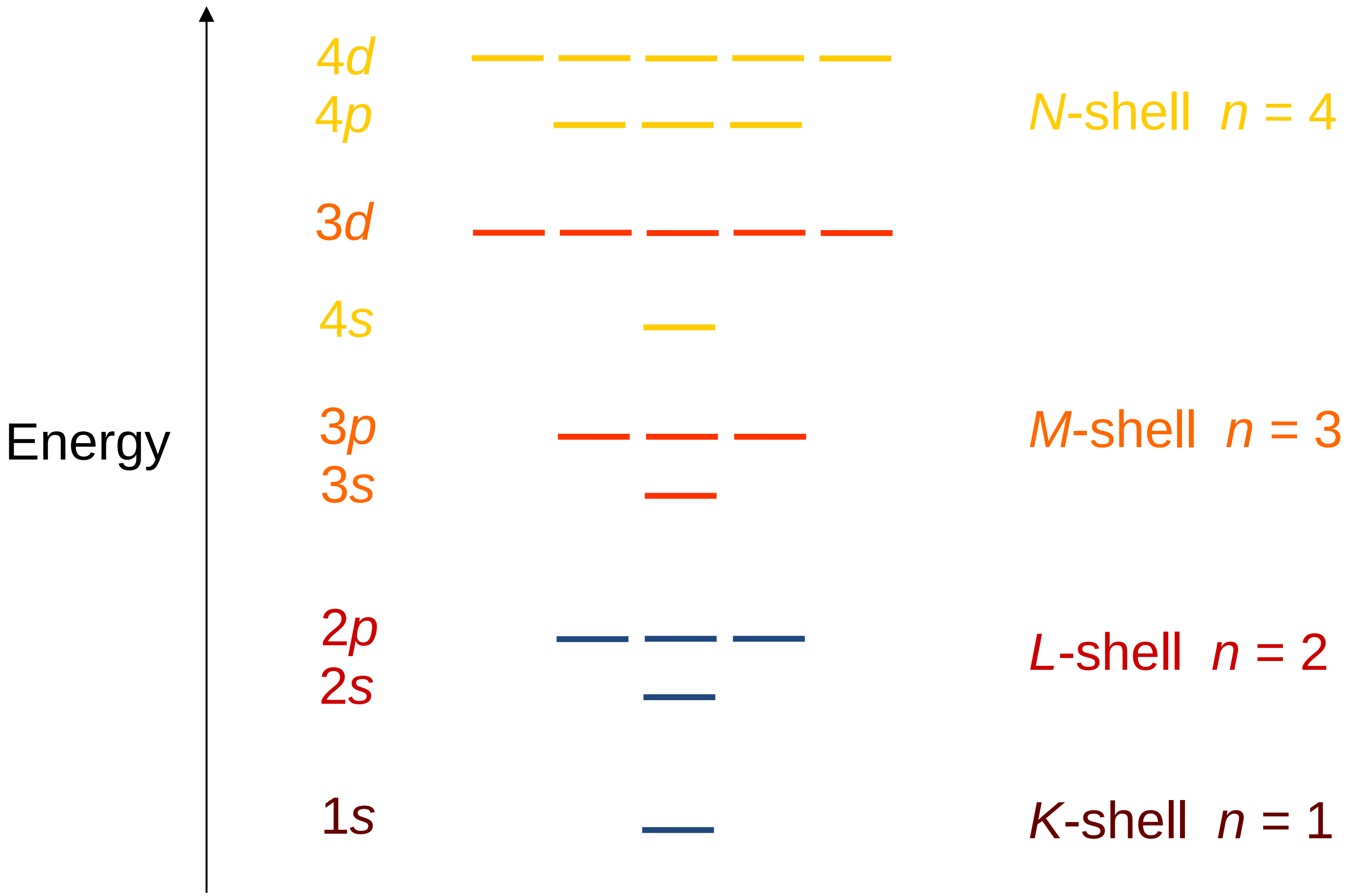
The max. no. of electrons that can occupy a specific shell can be found using the following formula:

Electron Capacity = $2n^2$

ELECTRON ENERGY STATES

Electrons...

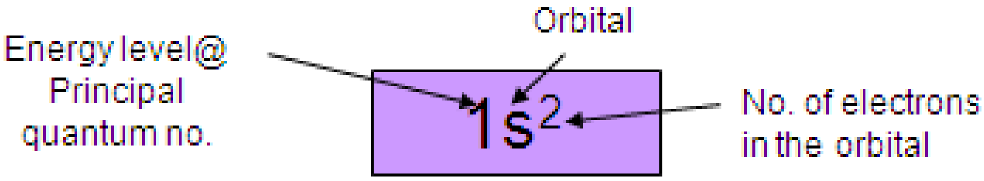
- have discrete energy states
- tend to occupy lowest available energy state.



ELECTRON CONFIGURATION

Electron configuration – the ways in which electrons are arranged around the nucleus of atoms.

- The following representation is used :



- Example: it means that there are two electrons in the ‘s’ orbital of the first energy level. The element is helium.

ELECTRON CONFIGURATION

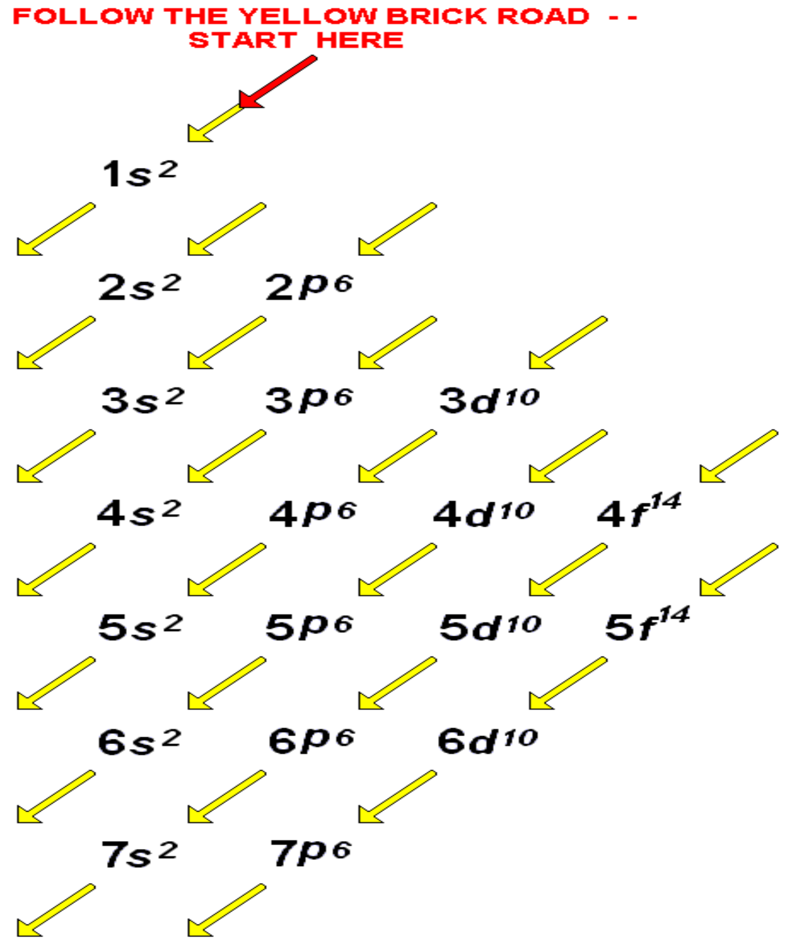
1. Aufbau principle: Electrons enter orbital of the lowest energy first.

2. Pauli exclusion principle: Each electron state can hold no more than two electrons that must have opposite spins.

Based on the Aufbau principle, which assumes that electrons enter orbital of lowest energy first.

The electrons in their orbital are represented as follows :

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



ELECTRON CONFIGURATION

- Most elements: Electron configuration not stable.

Element	Atomic #	Electron configuration	
Hydrogen	1	$1s^{1}$	
Helium	2	$1s^{2}$	(stable)
Lithium	3	$1s^{2}2s^{1}$	
Beryllium	4	$1s^{2}2s^{2}$	
Boron	5	$1s^{2}2s^{2}2p^{1}$	
Carbon	6	$1s^{2}2s^{2}2p^{2}$	
...		...	
Neon	10	$1s^{2}2s^{2}2p^{6}$	(stable)
Sodium	11	$1s^{2}2s^{2}2p^{6}3s^{1}$	
Magnesium	12	$1s^{2}2s^{2}2p^{6}3s^{2}$	
Aluminum	13	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{1}$	
...		...	
Argon	18	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}$	(stable)
...	...	...	
Krypton	36	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}3d^{10}4s^{2}4p^{6}$	(stable)

- Why? Valence (outer) shell usually not filled completely.

ELECTRON CONFIGURATION

- Valence electrons – those in unfilled shells
- Filled shells more stable
- Valence electrons are most available for bonding and tend to control the chemical properties

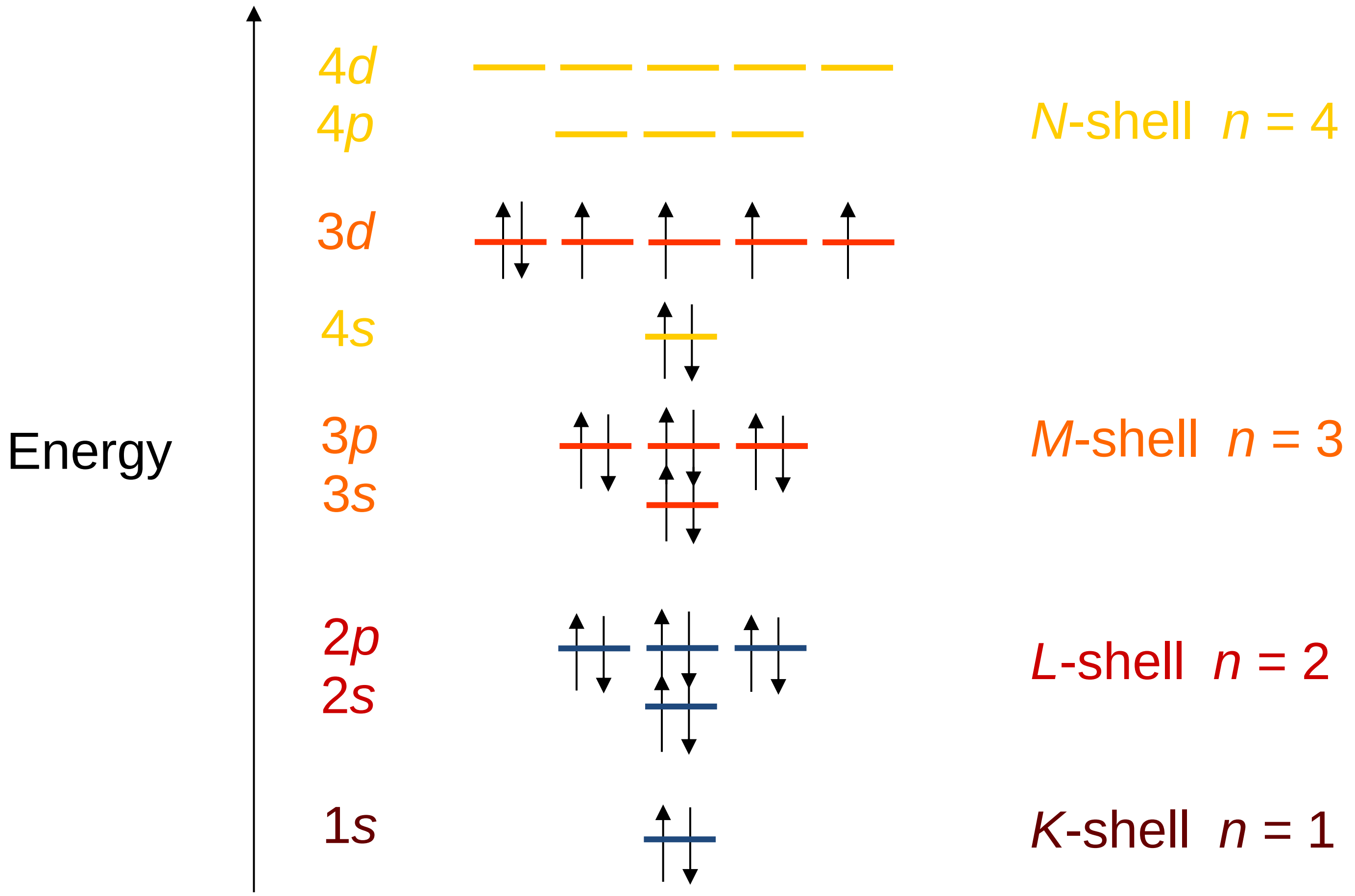
  – example: C (atomic number = 6)

$1s^2$ $2s^2\ 2p^2$

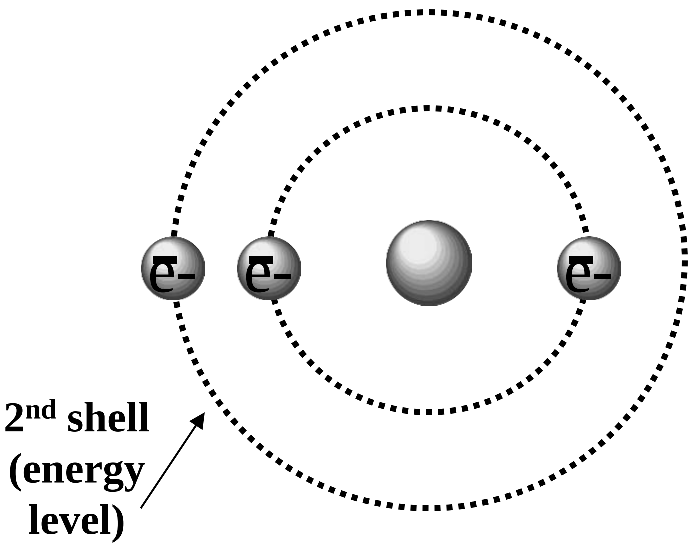
valence electrons

How to Write the Electron Configuration of the Element?

ex: Fe - atomic # = 26 $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^6$ $4s^2$

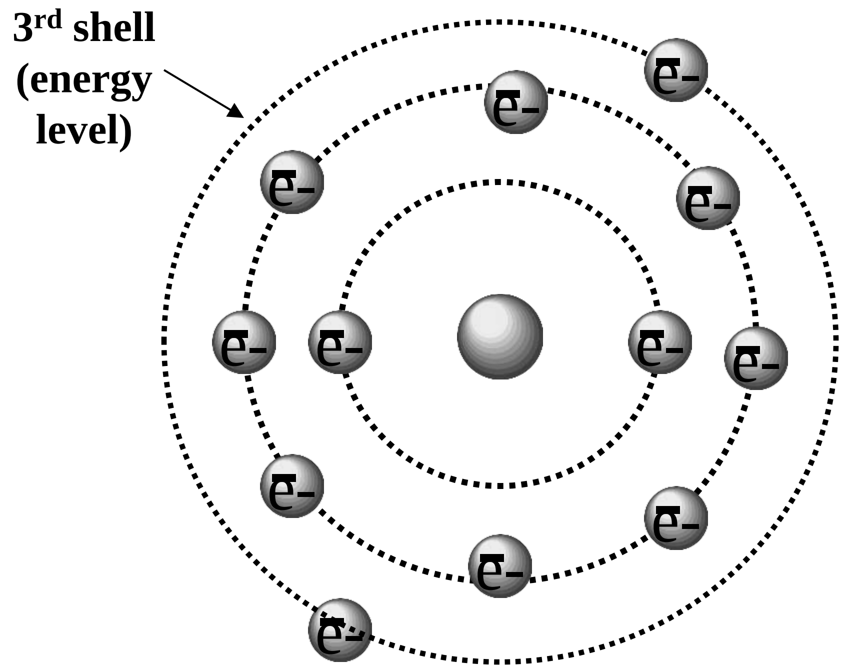


How to Write the Electron Configuration of the Element?



Lithium (3 electrons)

Answer : $ls^2\ 2s^1$



Magnesium (12 electrons)

Answer : $1s^2\ 2s^2\ 2p^6\ 3s^2$

EXERCISE

Write the electron configuration of the following species:

1. Ca (20 e)
2. O (8 e)
3. Cu (29 e)
4. O^{2-} (O = 8 e)
5. Fe^{2+} (Fe = 26 e)

Element	*Symbol*	*Atomic Number*	*Electron Configuration*
Hydrogen	H	1	$1s^1$
Helium	He	2	$1s^2$
Lithium	Li	3	$1s^22s^1$
Beryllium	Be	4	$1s^22s^2$
Boron	B	5	$1s^22s^22p^1$
Carbon	C	6	$1s^22s^22p^2$
Nitrogen	N	7	$1s^22s^22p^3$
Oxygen	O	8	$1s^22s^22p^4$
Fluorine	F	9	$1s^22s^22p^5$
Neon	Ne	10	$1s^22s^22p^6$
Sodium	Na	11	$1s^22s^22p^63s^1$
Magnesium	Mg	12	$1s^22s^22p^63s^2$
Aluminum	Al	13	$1s^22s^22p^63s^23p^1$
Silicon	Si	14	$1s^22s^22p^63s^23p^2$
Phosphorus	P	15	$1s^22s^22p^63s^23p^3$
Sulfur	S	16	$1s^22s^22p^63s^23p^4$
Chlorine	Cl	17	$1s^22s^22p^63s^23p^5$
Argon	Ar	18	$1s^22s^22p^63s^23p^6$
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	$1s^22s^22p^63s^23p^63d^{10}4s^24p^6$

Basics of the PERIODIC TABLE

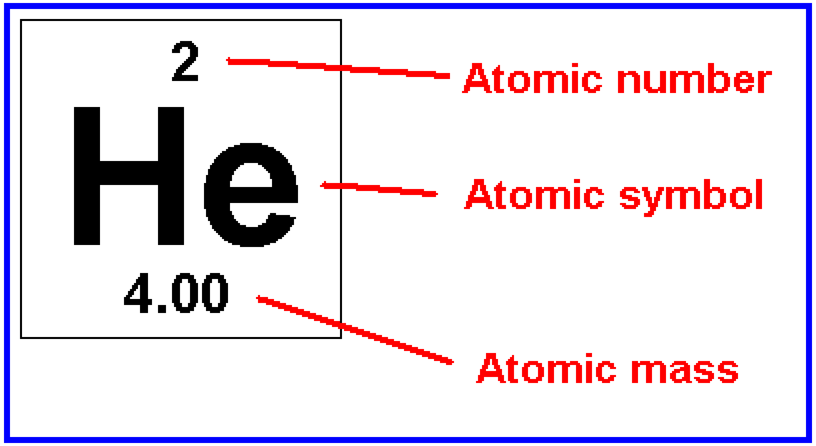
periodic: a repeating pattern

table: an organized collection of information

Periodic Table (P.T.)

An arrangement of elements in order of atomic number; elements with similar properties are in the same group.

The periodic table below is a simplified representation which usually gives the :



Two main classifications in P.T.

1) period: horizontal row on the P.T.

- Designate electron energy levels

2) group or **family:** vertical column on the P.T.

Groups to know

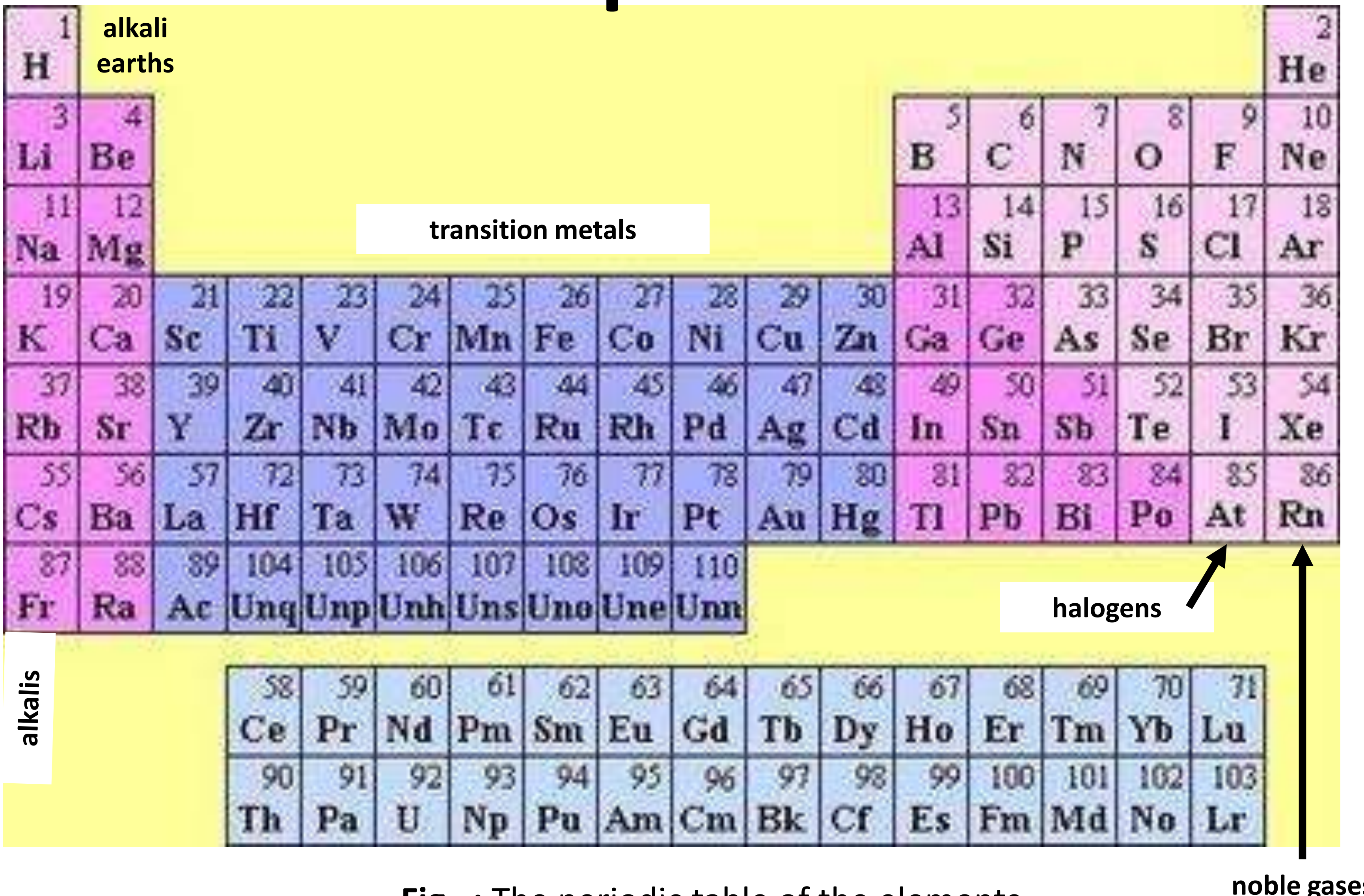


Fig. : The periodic table of the elements.

5 Periodic Trends to Know

Periodic trend: property of an element that can be predicted by position on Periodic Table.

Trend #1:

Elements in the same group have similar properties because they have same number of **valence e-** (e- in outermost energy level).

H 1							He 2
Li 2,1	Be 2,2	B 2,3	C 2,4	N 2,5	O 2,6	F 2,7	Ne 2,8
Na 2,8,1	Mg 2,8,2	Al 2,8,3	Si 2,8,4	P 2,8,5	S 2,8,6	Cl 2,8,7	Ar 2,8,8
K 2,8,8,1	Ca 2,8,8,2						

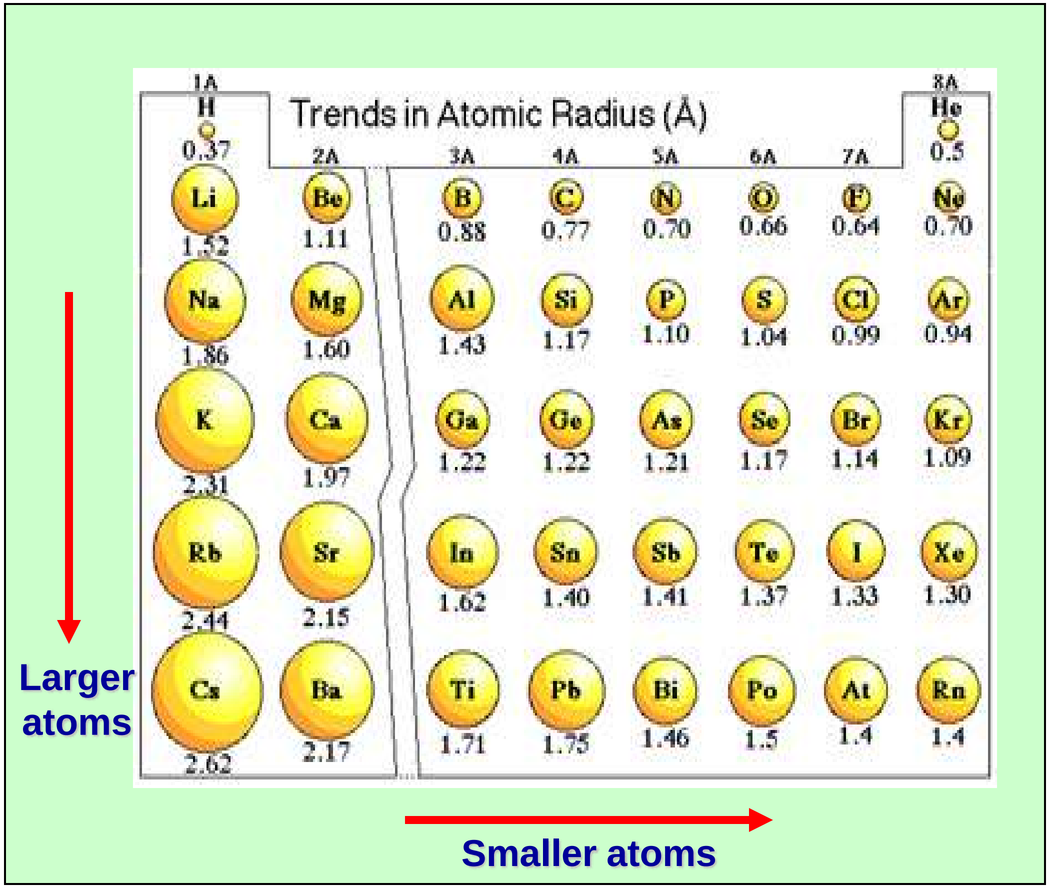
The number of valence e- increases as you go from left to right across a period; there is no change going down a group.

<u>Trend #2</u>: ATOMIC RADIUS

It is a distance from center of nucleus to valence e- energy level.

- **Atoms get smaller as you go across (left to right) a period.**
- Caused by increasing # of protons in nucleus.
- More protons pull e- closer = smaller radius

- **Atoms get larger as you go down a group.**
- Each new period represents a new energy level.
- More energy levels = larger radius



Trend #3: IONIZATION ENERGY (I.E.)

It is an energy required to remove one e- from a neutral atom.

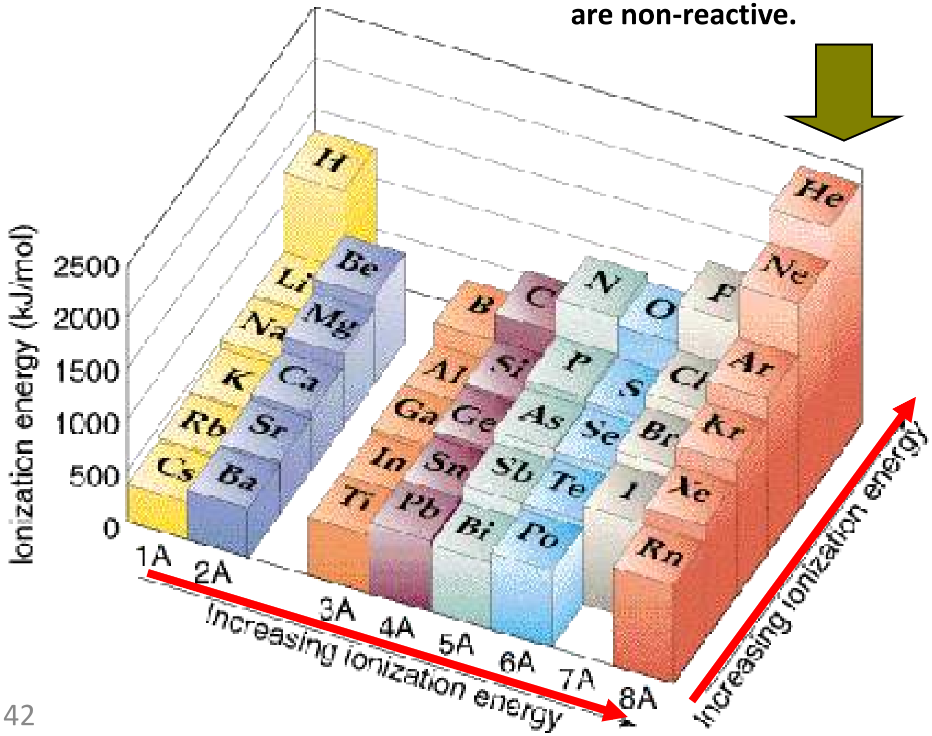
Eg : Na + energy → Na^{1+} + e^-

- **Down a group, I.E. decrease**
  - e- are removed more easily from energy levels farther from the nucleus.

- **Across a period, I.E. increase**
  - e- are removed less easily from atoms which are close to have *filled energy levels*.

Atoms are more stable if they have filled energy levels; therefore, atoms close to filling their energy levels will not easily give up their e-.

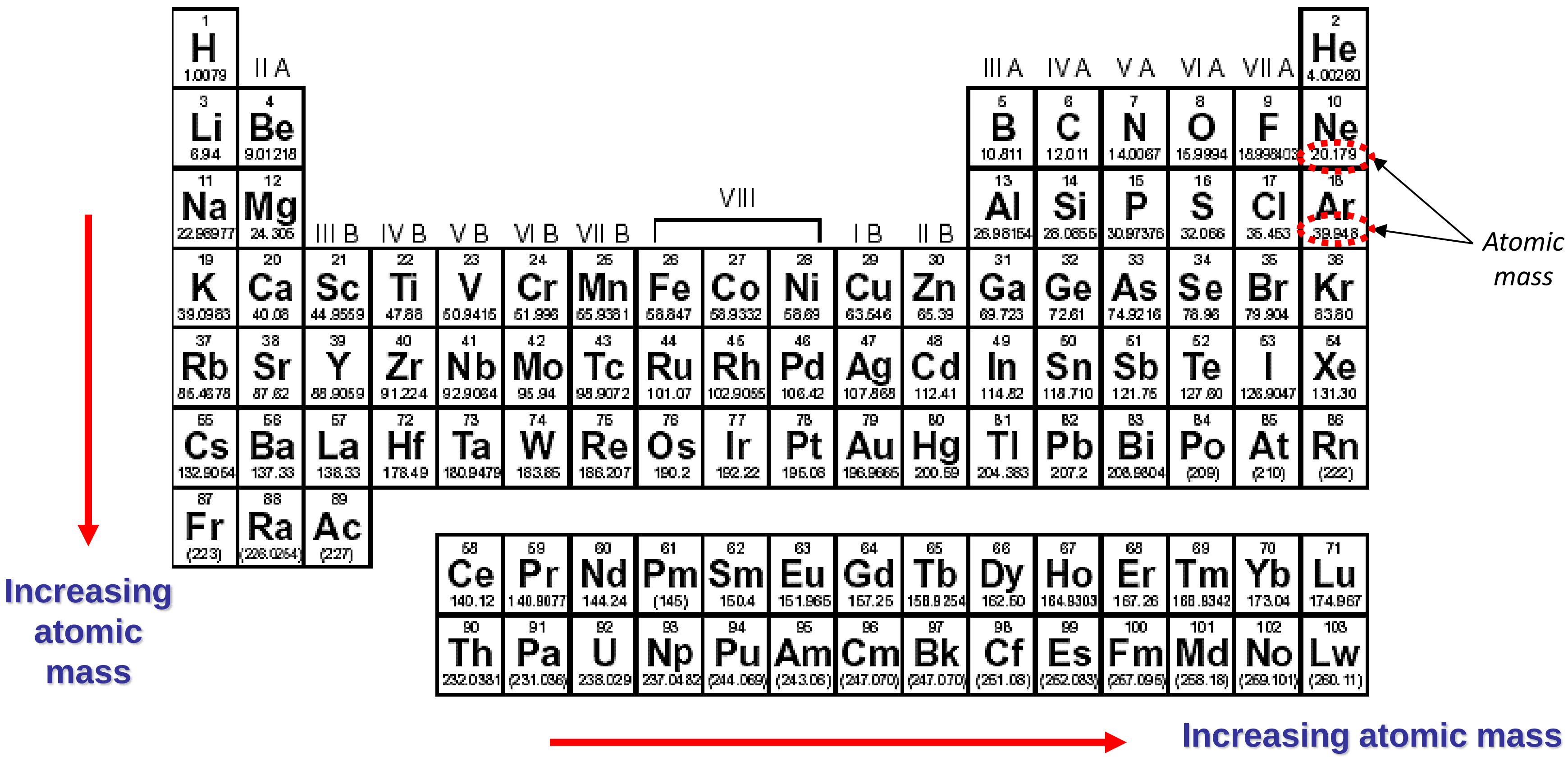
Noble gases have the highest I.E of all; they will only lose e- in extreme circumstances; thus, they are non-reactive.



Trend #4 : ATOMIC MASS

- **Down a group, atomic mass increase**
  - Protons (and e-) are added, increasing the mass.

- **Across (left to right) a period, atomic mass increase**
  - The same reason.



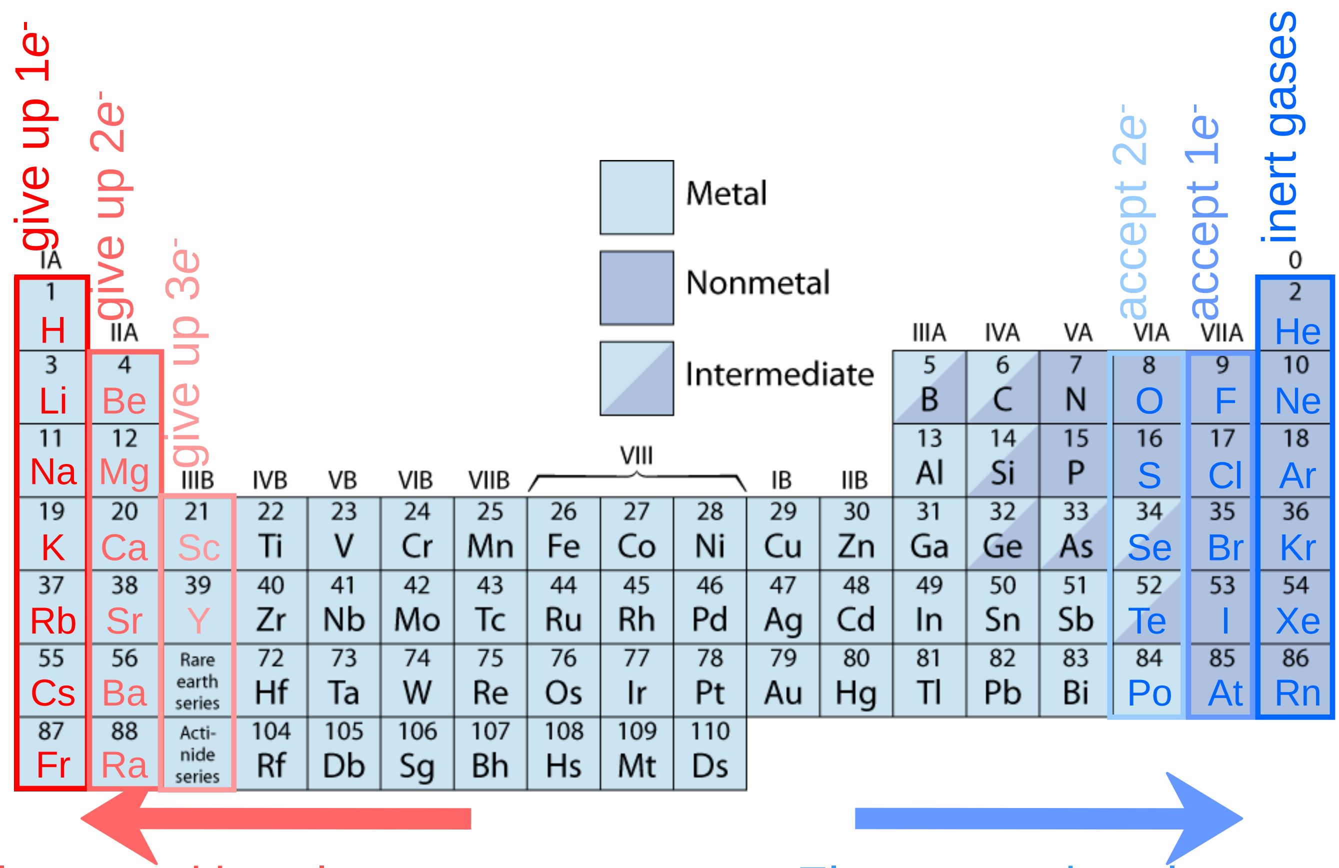
Trend #5 : Electronegativity (EN)

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 3.1	O 3.5	F 4.1	Ne –
Na 1.0	Mg 1.3											Al 1.5	Si 1.8	P 2.1	S 2.4	Cl 2.9	Ar –
K 0.9	Ca 1.1	Sc 1.2	Ti 1.3	V 1.5	Cr 1.6	Mn 1.6	Fe 1.7	Co 1.7	Ni 1.8	Cu 1.8	Zn 1.7	Ga 1.8	Ge 2.0	As 2.2	Se 2.5	Br 2.8	Kr –
Rb 0.9	Sr 1.0	Y 1.1	Zr 1.2	Nb 1.3	Mo 1.3	Tc 1.4	Ru 1.4	Rh 1.5	Pd 1.4	Ag 1.4	Cd 1.5	In 1.5	Sn 1.7	Sb 1.8	Te 2.0	I 2.2	Xe –
Cs 0.9	Ba 0.9	La 1.1	Hf 1.2	Ta 1.4	W 1.4	Re 1.5	Os 1.5	Ir 1.6	Pt 1.5	Au 1.4	Hg 1.5	Tl 1.5	Pb 1.6	Bi 1.7	Po 1.8	At 2.0	Rn –
Fr 0.9	Ra 0.9	Ac 1.0															

Lanthanides: 1.0-1.2
Actinides: 1.0-1.2

Smaller electronegativity

Larger electronegativity

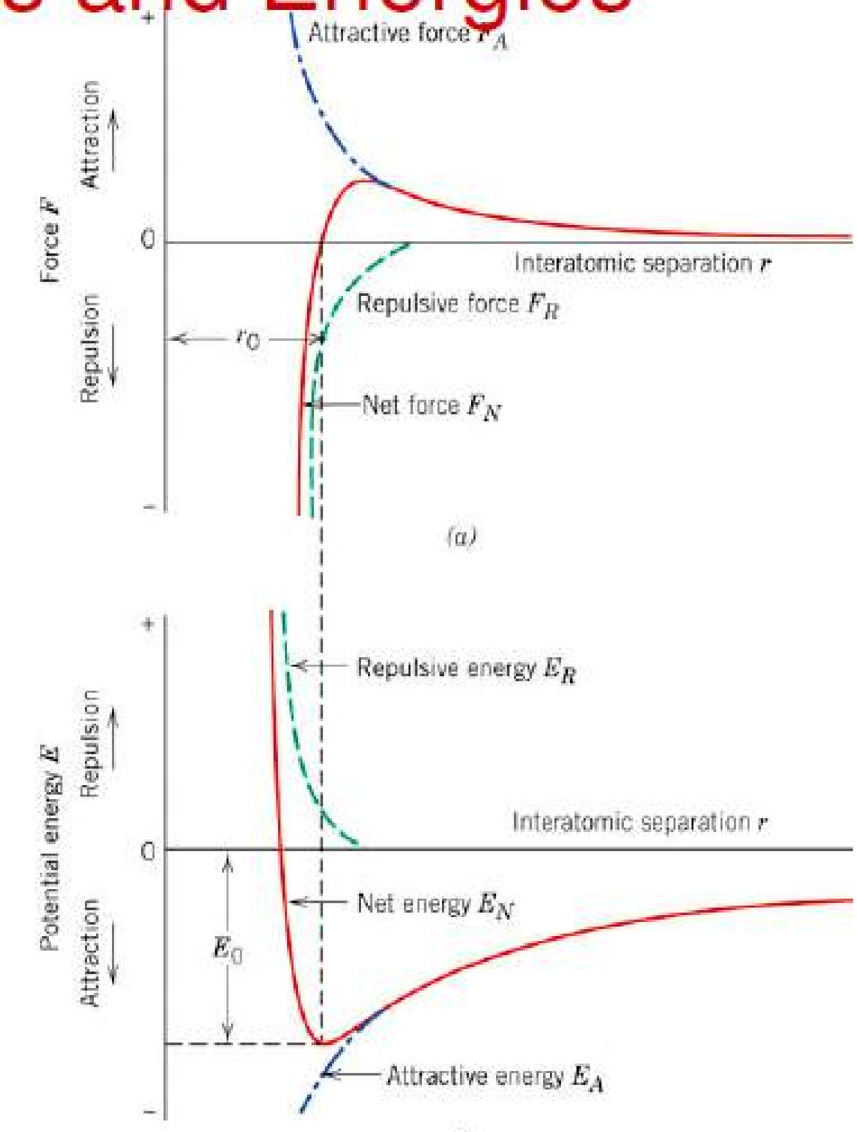


Electropositive elements:
Readily give up electrons
to become + ions.

Electronegative elements:
Readily acquire electrons
to become - ions.

Bonding Forces and Energies

- Considering the interaction between two isolated atoms as they are brought into close proximity from an infinite separation.
- At larger distances, the interactions are negligible.
- As the atoms approach, each exerts forces on the other.
  - Attractive
  - Repulsive
- Ultimately, the outer electron shells of the two atoms begin to overlap, and a strong repulsive force comes into play.



Bonding Models

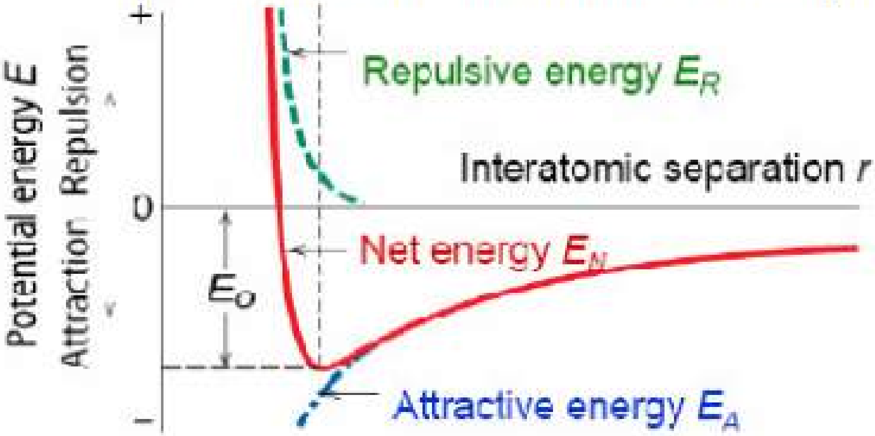
- Bonding holds atoms together to form solids materials
- In solids, atoms are held at preferred distances from each other (equilibrium distances)
- Distances larger or smaller than equilibrium distances are not preferred. Equilibrium spacing r_0 is approximately 0.3nm
- Consequently, as atomic bonds are stretched, atoms tend to attract each other, and as the bonds are compressed, atoms repel each other.
- Simple bonding models assume that the total bonding results from the sum of two forces: an attractive force (F_A) and a repulsive (F_R).

the net force $$F_N = F_A + F_R$$

- The repulsive force dominates at small distances, and the attractive force dominates at larger distances. At equilibrium they are just equal.

Bonding Forces and Energies

- It is convenient to work with energy than forces.
- Bonding energy (also called interaction energy or potential energy) between two isolated atoms at separation *r* is related to the force by



$$E(r) = \int_{\infty}^{r} F(r')dr'$$

$$= \int_{\infty}^{r} (F_A(r') + F_R(r'))dr'$$

$$= E_A(r) + E_R(r)$$

Bonding energy between two atoms

- The interaction energy at equilibrium is called the bonding energy between the two atoms.

- To break the bond, this energy must be supplied from outside.

- Breaking the bond means that the two atoms become infinitely separated.

- In real materials, containing many atoms, bonding is studied by expressing the bonding energy of the entire materials in terms of the separation distances between all atoms, see later discussion.

PRIMARY AND SECONDARY ATOMIC BONDING

- The forces of attraction that hold atoms together are called chemical bonds which can be divided into 2 categories :

1) <u>Primary Interatomic Bonding</u>
Metallic, ionic and covalent

2) <u>Secondary Atomic Bonding</u>
Van der Waals

Metallic atoms		Nonmetallic atoms
↓↓	↓↓	↓↓
Metallic bonds	Ionic bonds	Covalent bonds

- Chemical reactions between elements involve either the releasing/receiving or sharing of electrons .

Bonding

Primary bonding:

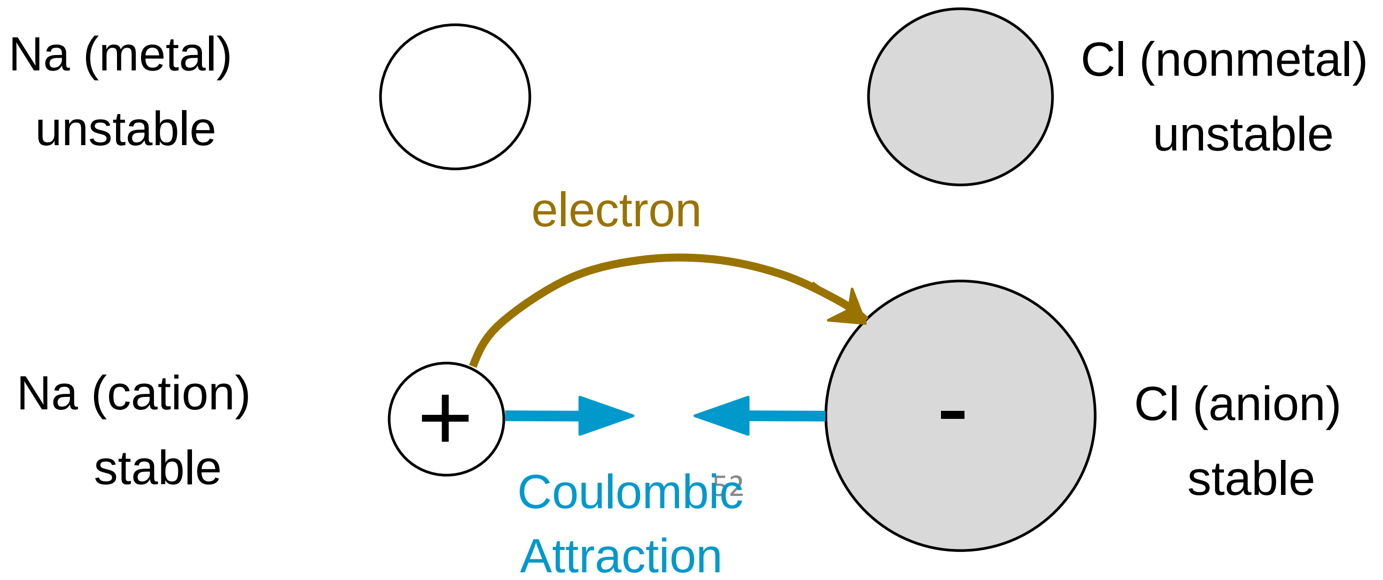
Ionic (transfer of valence electrons)
Covalent (sharing of valence electrons, directional)
Metallic (delocalization of valence electrons)

Secondary or van der Waals Bonding:
(Common, but weaker than primary bonding)
Dipole-dipole
H-bonds
Polar molecule-induced dipole
Fluctuating dipole (weakest)

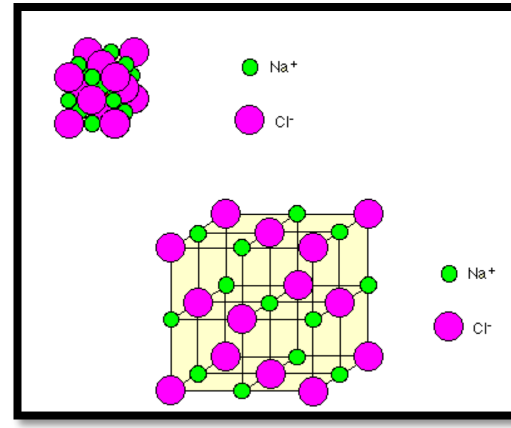
PRIMARY INTERATOMIC BONDING

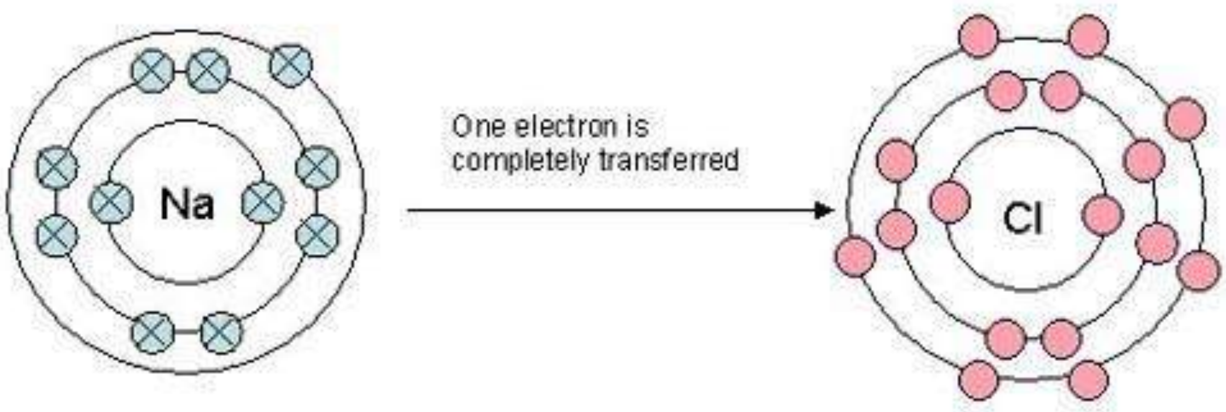
1) IONIC BONDING

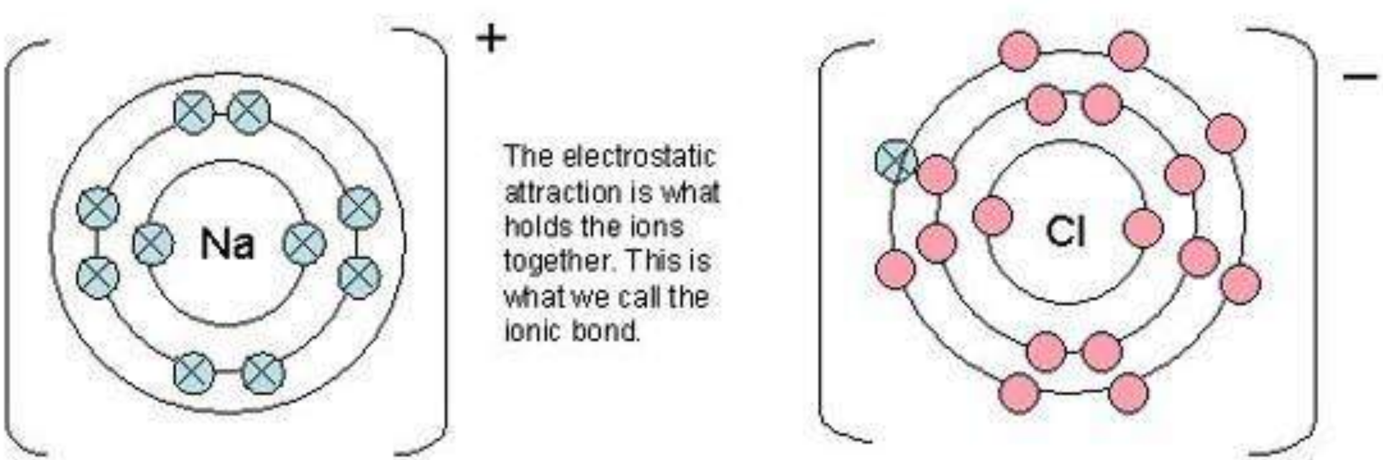
- Occurs between + and - ions.
- Requires electron transfer.
- Large difference in electronegativity required.
- Example: NaCl



- Example: NaCl







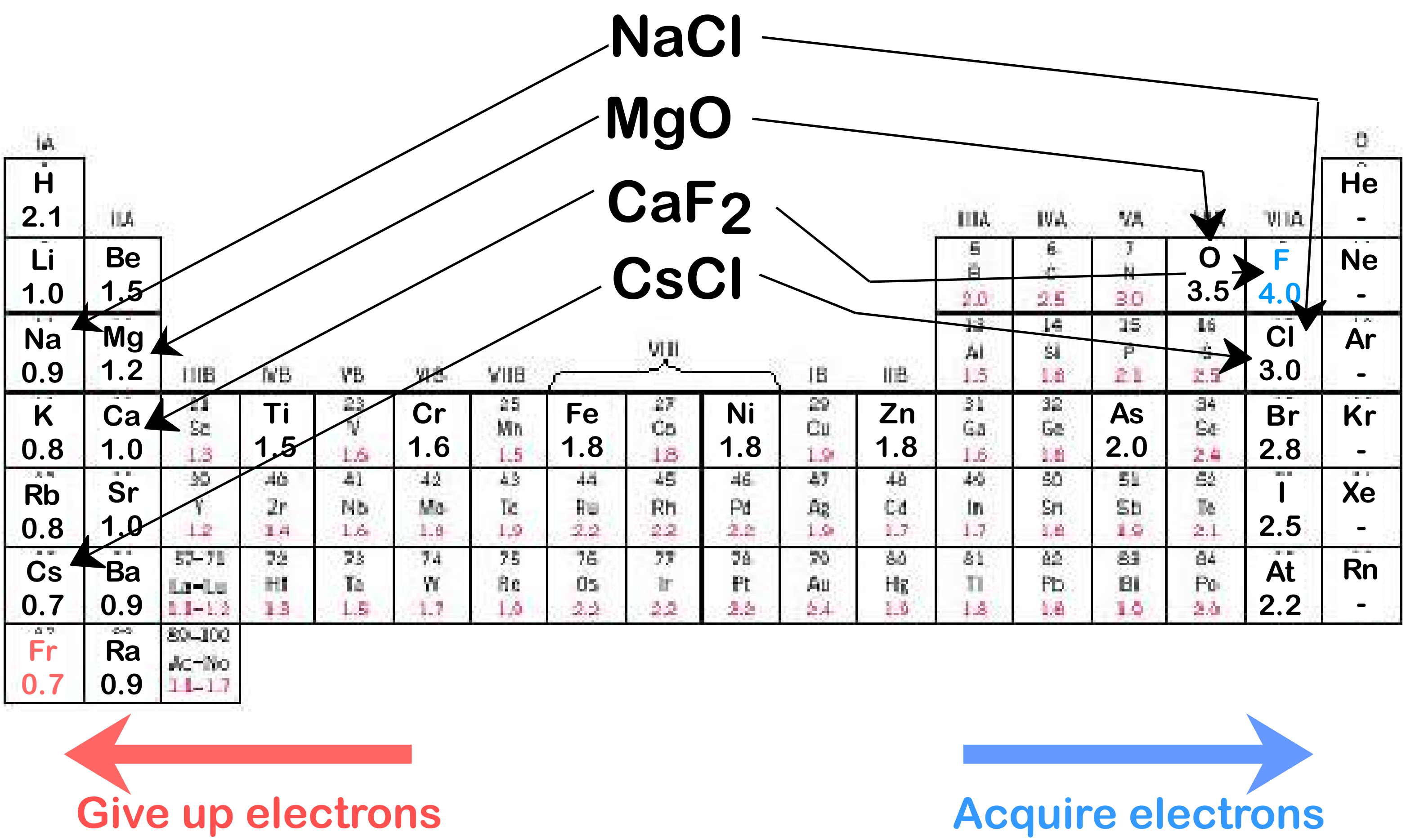
IONIC BONDING

- **<u>Properties</u>** :

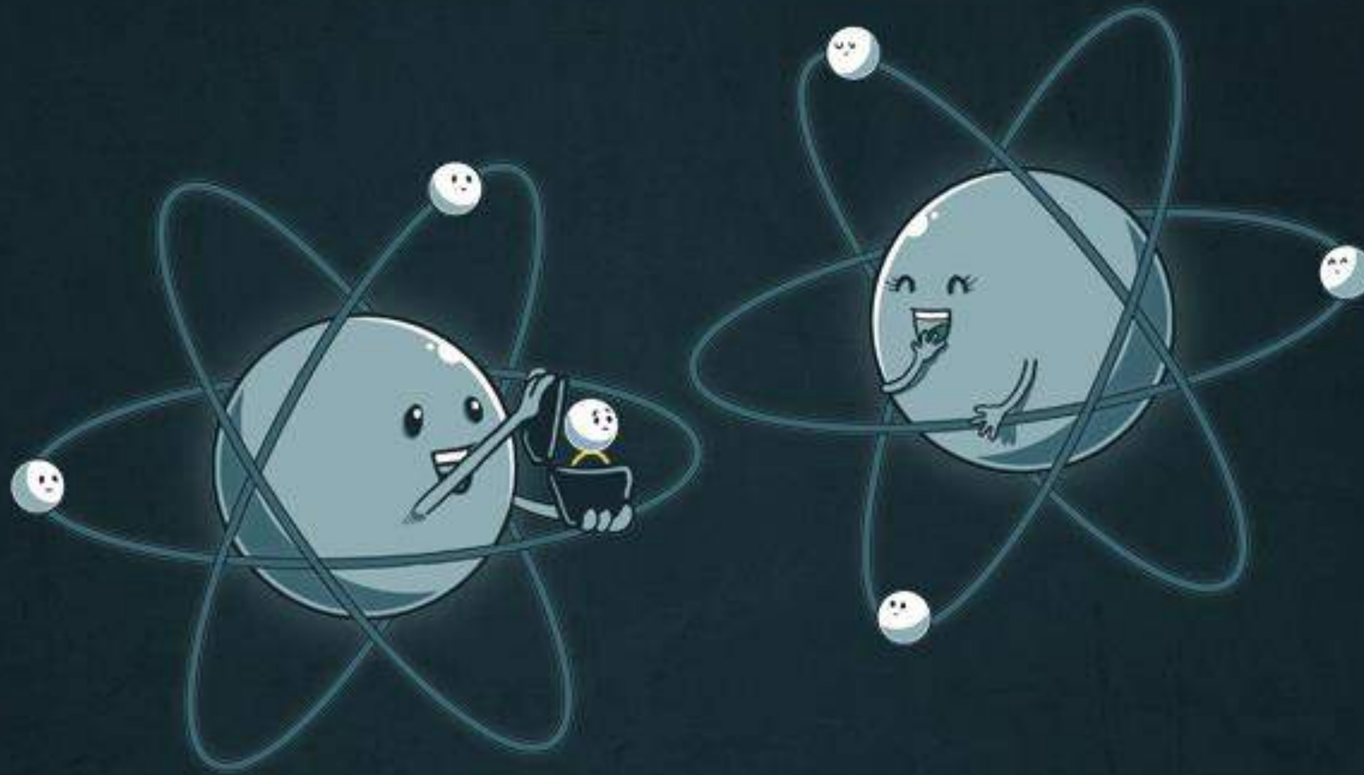
✓ Solid at room temperature (made of ions)
✓ High melting and boiling points
✓ Hard and brittle
✓ Poor conductors of electricity in solid state
✓ Good conductor in solution or when molten

EXAMPLE : IONIC BONDING

- Predominant bonding in Ceramics



TYPES OF CHEMICAL BONDS



#1: IONIC

2) COVALENT BONDING

How is covalent bonding formed??

- Electrons are **shared** to form a bond.
- Most frequently occurs between atoms with similar electronegativities.
- Often found in:

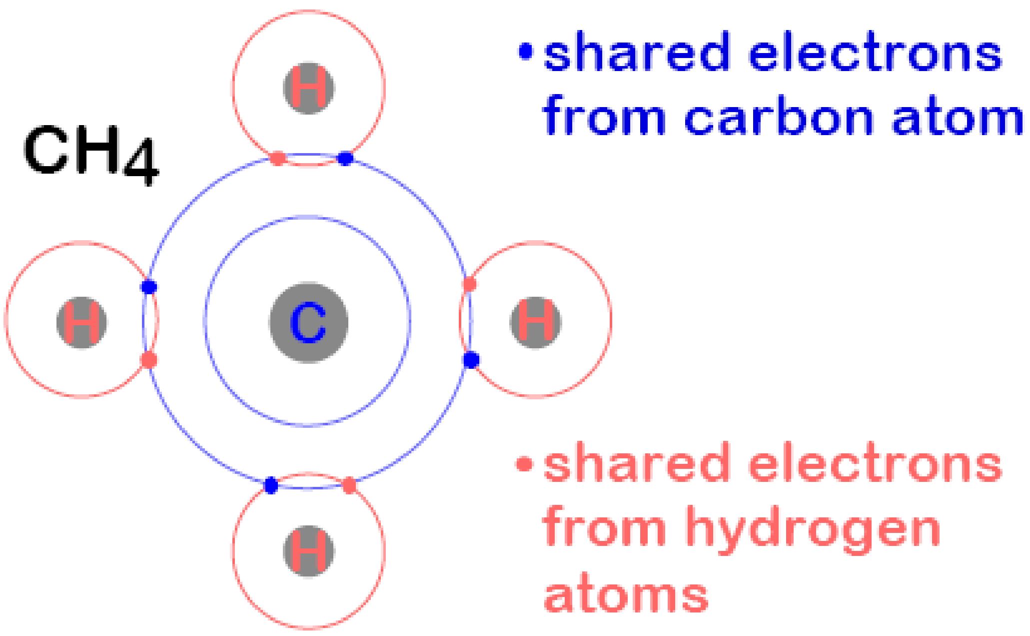
  - Molecules with nonmetals (H_2, Cl_2, F_2, etc)
  - Molecules with metals and nonmetals (aluminum phosphide (AlP))
  - Elemental solids (diamond, silicon, germanium)
  - Compound solids (about column IVA) (gallium arsenide - GaAs, indium antimonide - InSb and silicone carbide - SiC).

- Example: CH_4

C: has 4 valence e, needs 4 more

H: has 1 valence e, needs 1 more

Electronegativities are comparable.

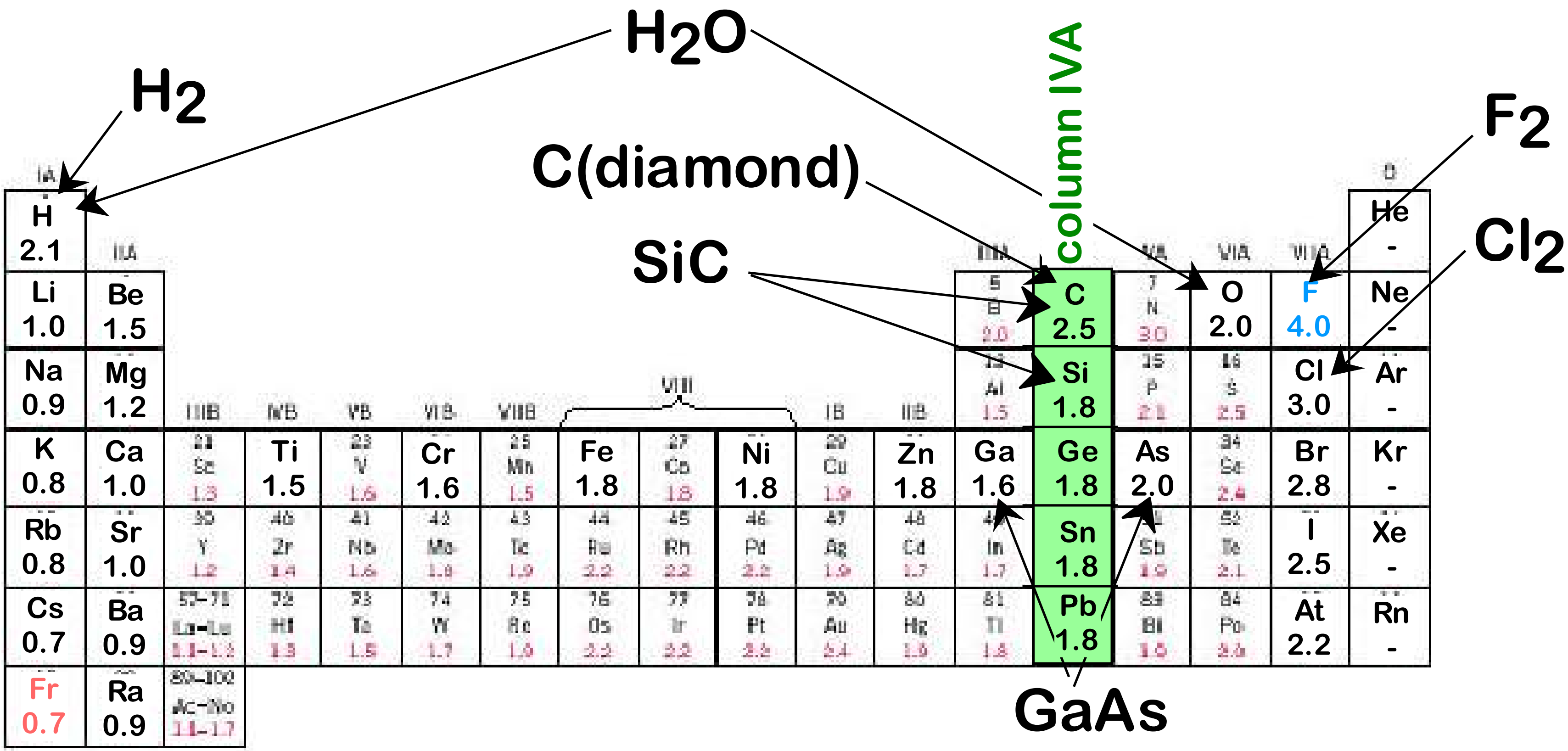


COVALENT BONDING

Properties

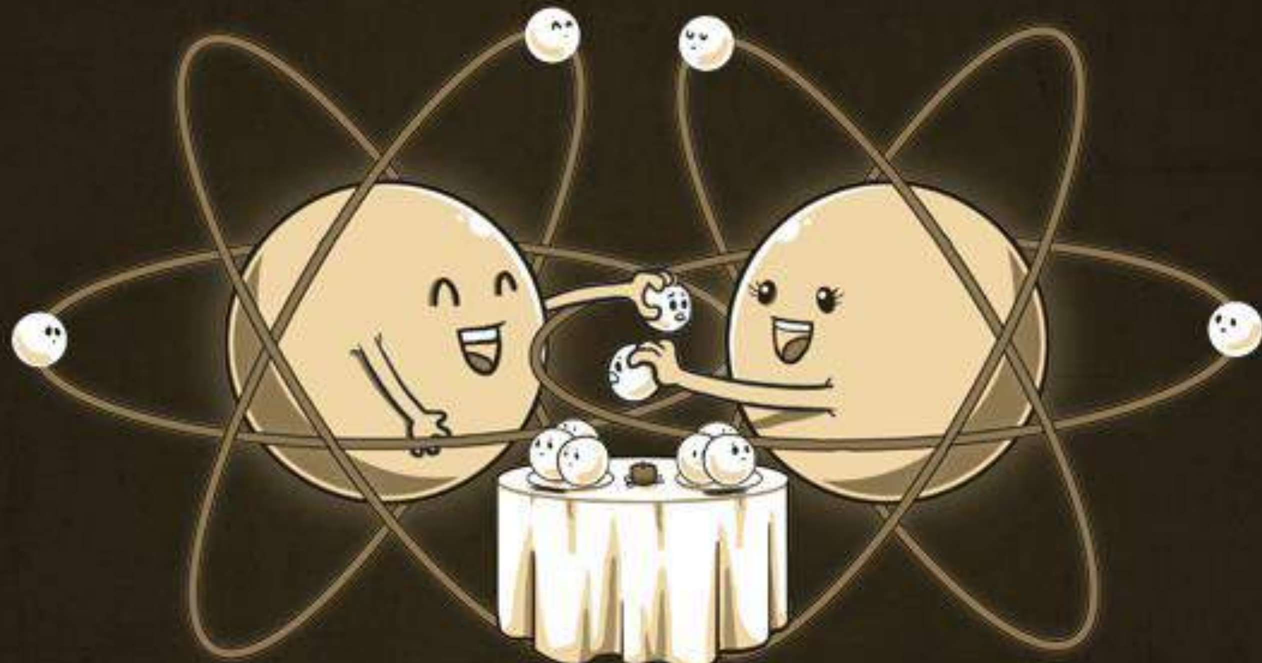
- Gases, liquids, or solids (made of molecules)
- Poor electrical conductors in all phases
- Variable (hard , strong, melting temperature, boiling point)

EXAMPLE : COVALENT BONDING



- Molecules with nonmetals
- Molecules with metals and nonmetals
- Elemental solids
- Compound solids (about column IVA)

TYPES OF CHEMICAL BONDS



#2: COVALENT

3) METALLIC BONDING

How is metallic bonding formed??

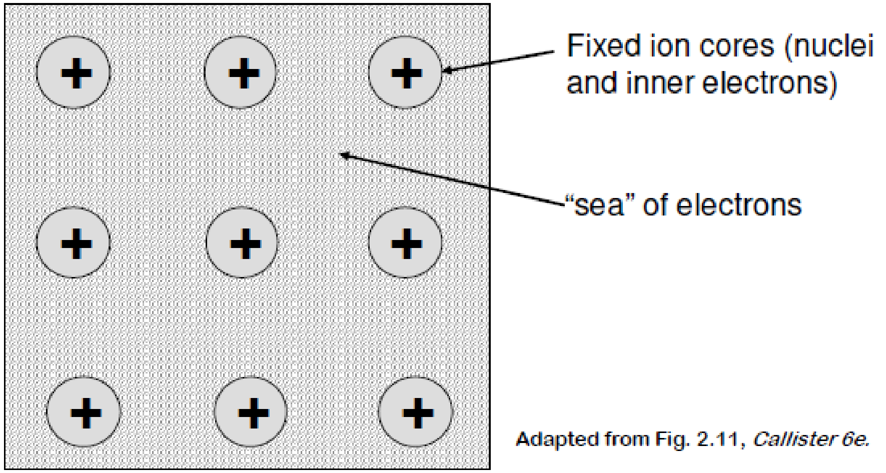
- Occur when some electrons in the valence shell separate from their atoms and exist in a cloud surrounding all the positively charged atoms.

- The valence electron form a ‘sea of electron’

- Found for group IA and IIA elements

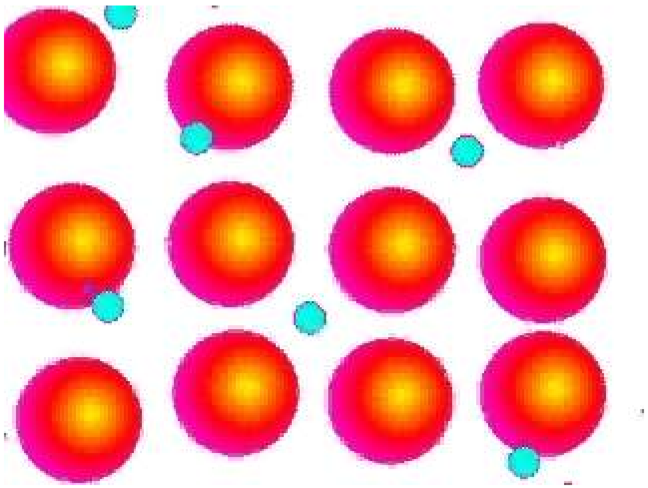
- Found for all elemental metals and its alloy

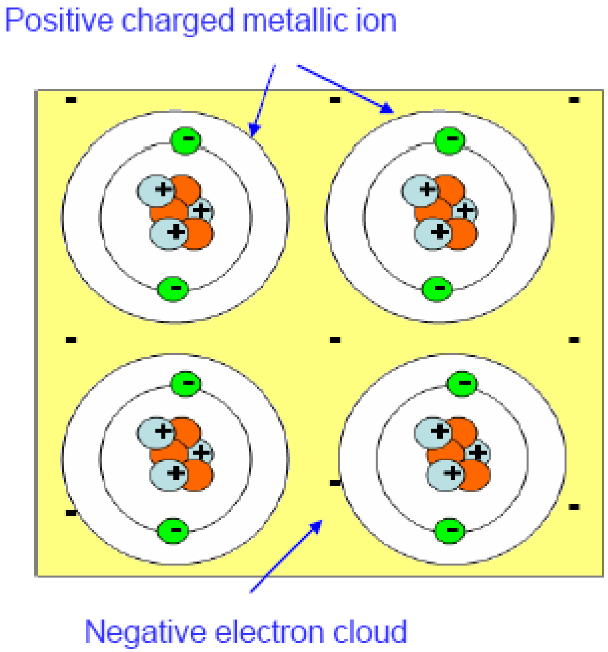
Arises from a sea of donated valence electrons



Adapted from Fig. 2.11, *Callister 6e.*

3) METALLIC BONDING

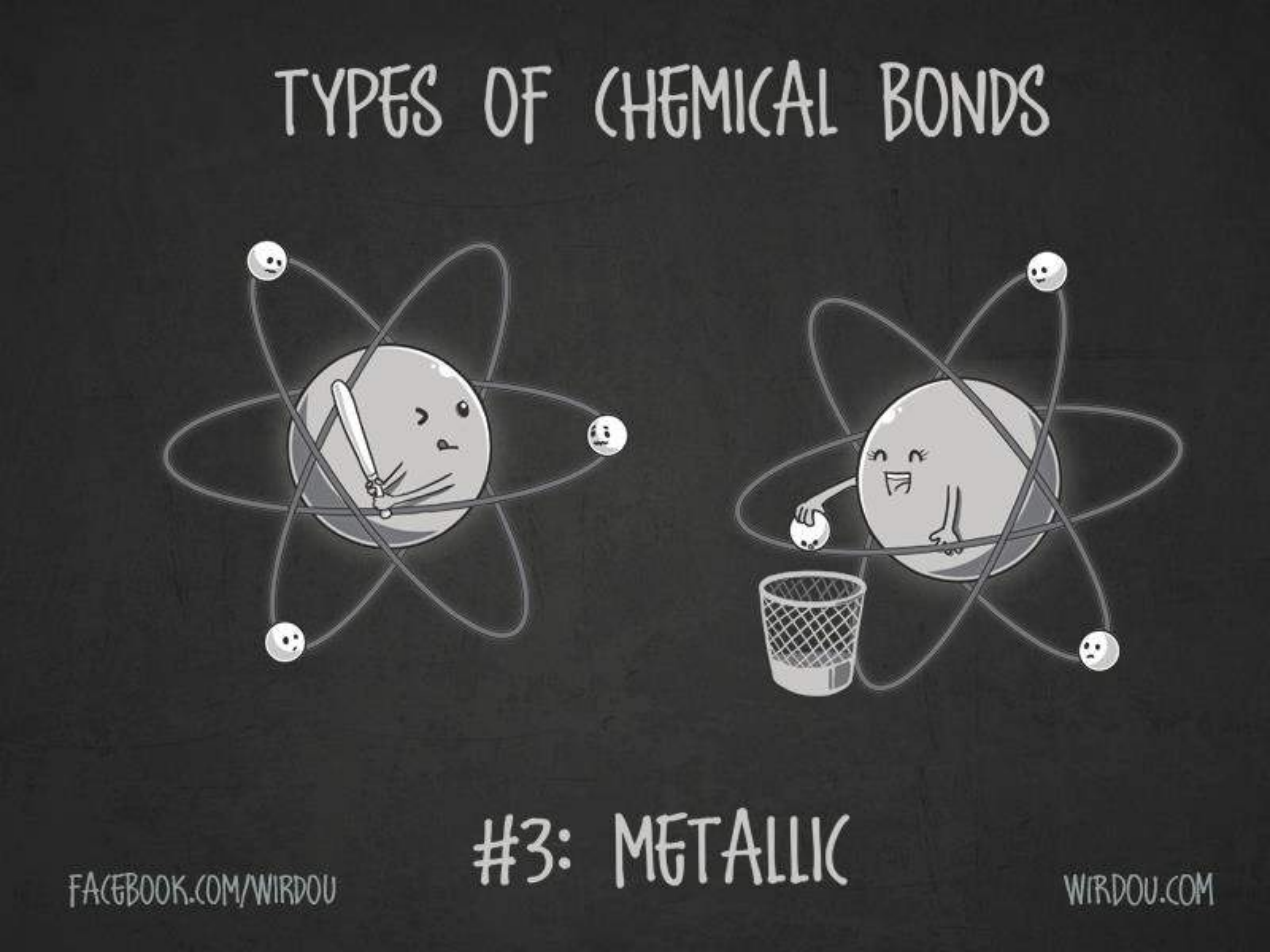




METALLIC BONDING

Properties:

- ✓ Good electrical conductivity-cloud electron are free to move to conduct electricity
- ✓ Good heat conductivity
- ✓ Ductile


TYPES OF CHEMICAL BONDS
#3: METALLIC
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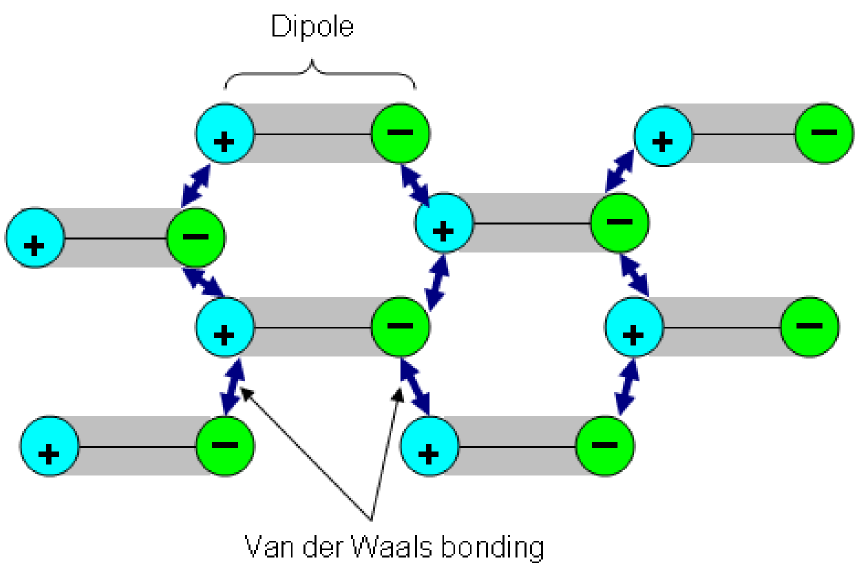
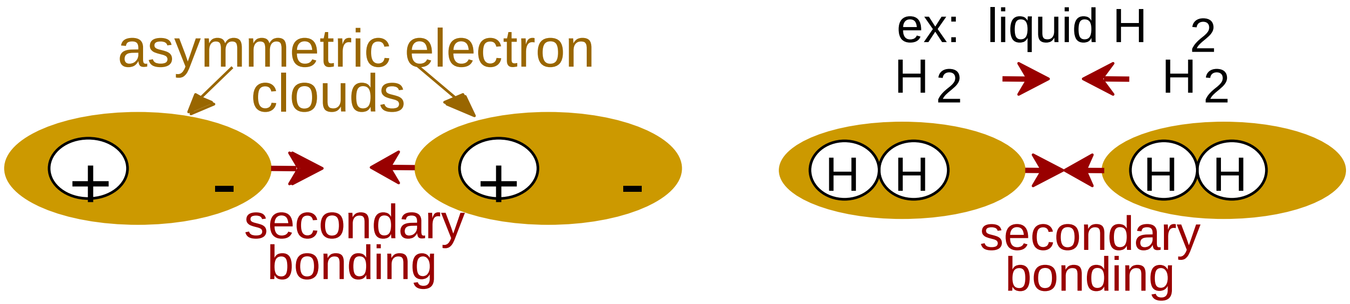
SECONDARY INTERATOMIC BONDING

VAN DER WAALS

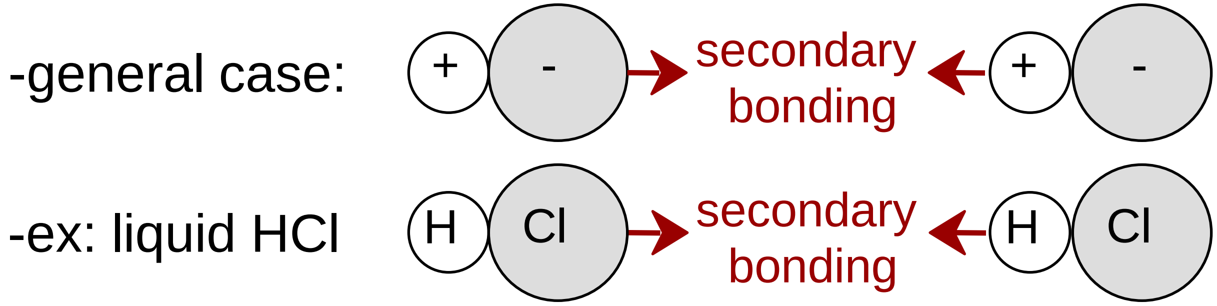
- Arise from atomic or molecular dipoles



- Three bonding mechanism
  - **Fluctuating Induced Dipole Bonds**
    - Eg: Inert gases, symmetric molecules (H_2, Cl_2)

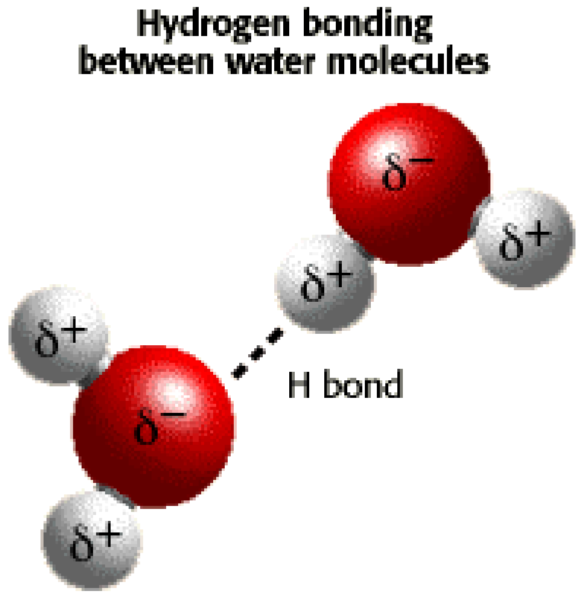


  - **Polar molecule-Induced Dipole Bonds**
    - Asymmetrical molecules such as HCl, HF



- Permanent Dipole Bonds

- Hydrogen bonding
- Between molecules
- H-F, H-O, H-N



Summary of BONDING

Type	Bond energy	Melting point	Hardness	Conductivity	Comments
ionic bonding	Large (150-370kcal/mol	Very high	Hard and brittle	Poor -required moving ion	Nondirectional (ceramic)
Covalent bonding	Variable(75-300 kcal/mol Large – Diamond Small – Bismuth	Variable Highest – diamond (>3550) Mercury (-39)	Very hard (diamond)	poor	Directional (Semiconductors, ceramic, polymer chains)
metallic bonding	Variable(25-200 kcal/mol) Large-Tungsten Small- Mercury	Low to high	Soft to hard	excellent	Nondirectional (metal)
Secondary bonding	Smallest	Low to moderate	Fairly soft	poor	Directional inter-chain (polymer) inter-molecular

**** Directional bonding – Strength of bond is not equal in all directions***

**** Nondirectional bonding – Strength of bond is equal in all directions***

LEARNING OBJECTIVE

You should be able:

- Describe an atomic structure
- Configure electron configuration
- Differentiate between each atomic bonding
- Briefly describe ionic, covalent, metallic, hydrogen and van der waals bonds
- Relate the atomic bonding with material properties