Solid State Electronic

Code:OPE 406

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What is the solid electronic?

Study of solids armed with quantum theory.
Focus largely on the particles living inside solid
Part of condensate matter physics (quantum fluids, soft biological matter, liquid crystal, …)

Why solid state is so important?

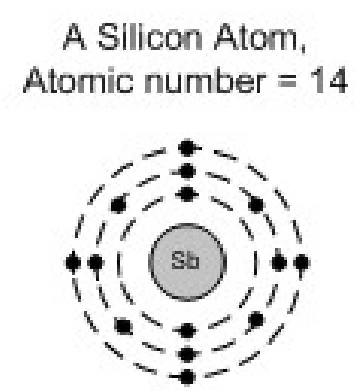
Transistor
Integrated circuit
Magnetic hard drive
Solid state laser
CCD
Solar cell
Superconducting magnate (MRI)

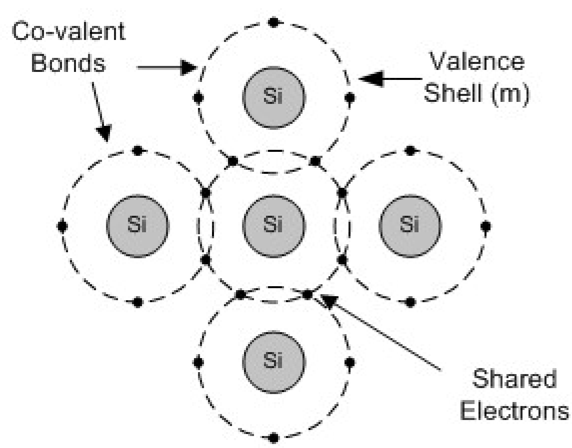
What we will learn in Solid state

1 Atomic structures
2 Solid and semiconductor
3 Photons, surface and energy levels
4 Electronic Materials
5 Silicon Semiconductor
6 Probability density
7 Allwed and Forbidden energy band
8 Electron static in solid and heisenberg
9 Intrinsic Semiconductor
10 Intrinsic semi. conductivity
11 Extrinsic semiconductor
12 Extinsic semi Fermi level
13 Extrinsic semi. conductivity
14 Hall effect and type current
15 Type of junction
16 Pn juction
17 Pn juction hetero breakdown
18 Transistor
19 Optoelectronics devices
20 LEDs
21 Photodetectors
22 Solar cell
23 Introduction to semiconductor manufacturing
24 Doping processes
25 Oxidation
26 Cleanroom design and contamination control

Atomic Structure

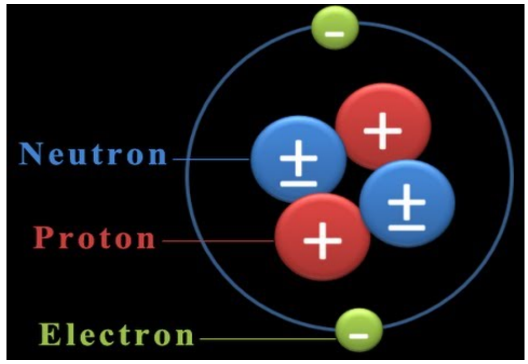
- To understand how semiconductors work, you must first understand a little about how electrons are organized in an atom. The electrons in an atom are organized in layers. These layers are called shells. The outermost shell is called the valence shell.
- The electrons in this shell are the ones that form bonds with neighboring atoms. Such bonds are called covalent bonds. Most conductors have just one electron in the valence shell. Semiconductors, on the other hand, typically have four electrons in their valence shell.
- If all the neighboring atoms are of the same type, it's possible for all the valence electrons to bind with valence electrons from other atoms. When that happens, the atoms arrange themselves into structures called crystals. Semiconductors are made out of such crystals, usually silicon crystals.
- Here, each circle represents a silicon atom, and the lines between the atoms represent the shared electrons. Each of the four valence electrons in each silicon atom is shared with one neighboring silicon atom. Thus, each silicon atom is bonded with four other silicon atoms.





Atomic Structure

Atomic number (Z) number of proton
Atomic mass (A)   (N Proton + N neutron)
Net charge



Atomic weighted Average
Atomic weight
Average Atomic mass

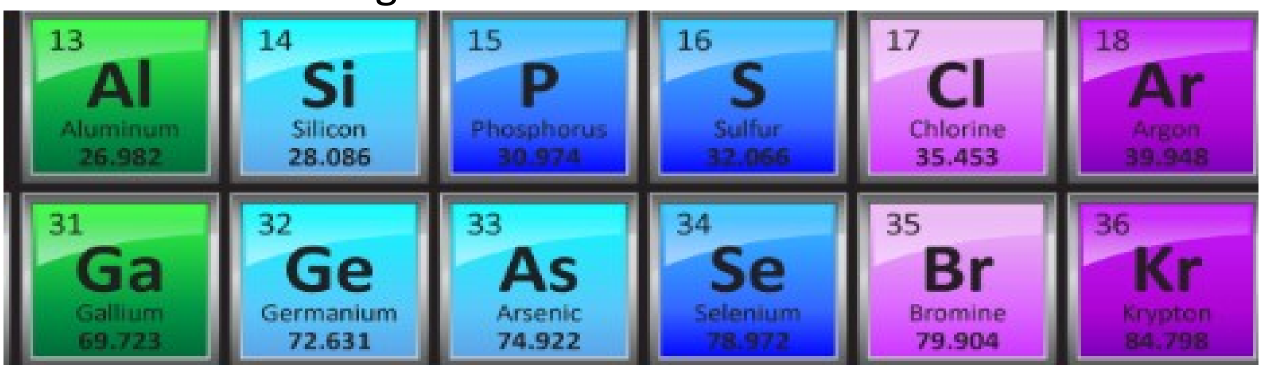


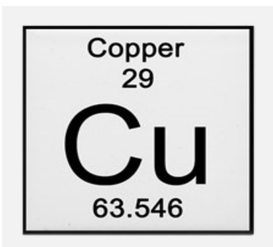
Table Fundamental particles of atom and their characteristics

Particle	Symbol	Mass/ kg	Actual Charge / C	Relative charge
Electron	*e*	$9.109\,389 \times 10^{-31}$	$-1.602\,177 \times 10^{-19}$	–1
Proton	*p*	$1.672\,623 \times 10^{-27}$	$1.602\,177 \times 10^{-19}$	+1
Neutron	*n*	$1.674\,928 \times 10^{-27}$	0	0

Chemistry has been defined as the study of matter in terms of its structure, composition and the properties. As you are aware, matter is made up of atoms, and therefore an understanding of the structure of atom is very important. You have studied in your earlier in High school that the earliest concept of atom (smallest indivisible part of matter) was given by ancient (600-400 BC) Indian and Greek philosophers. At that time there were no experimental evidence. The origin of the concept of atom was based on their thoughts on ‘What would happen if we continuously keep dividing matter’.

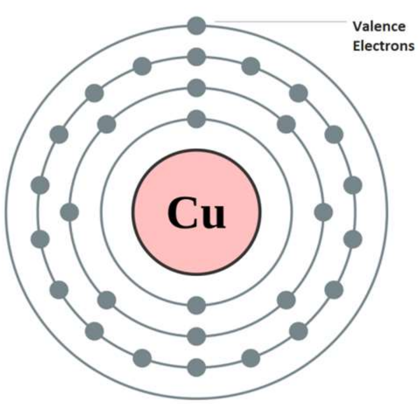
John Dalton revived the concept of atom in the beginning of nineteenth century in terms of his atomic theory which successfully explained the laws of chemical combination. Later experiments showed that the atom is not indivisible but has an internal structure. In this lesson you will learn about the internal structure of an atom which will help you to understand the correlations between its structure and properties. You would learn about these in the later lessons.

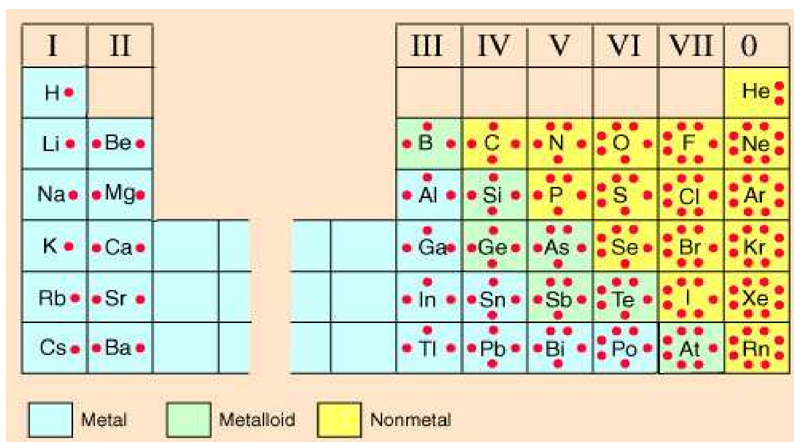
Atomic mass unit (amu) : A weighted Average of the mass for all the Isotopes of a certain element

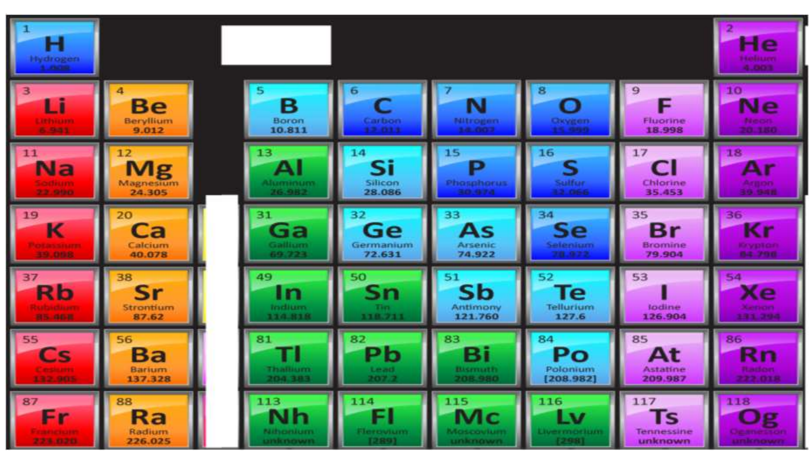


(63X0.69)+(65X0.29)=63.62

Why the number is not same in periodic table?







Atomic Structure

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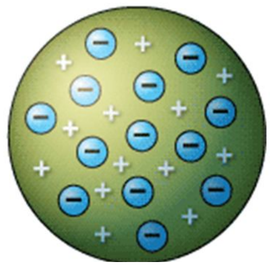
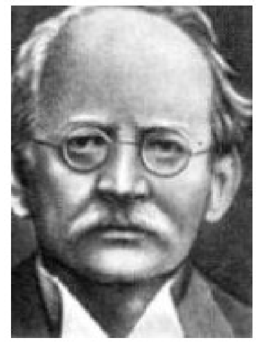
- 2600 years ago Democritus Atomos (uncuttable)
- 1805 Dalton first experiment (particles in different shape) (Atoms)

J.j. Thomson Atomic Model

- Plum Pudding Model or Blueberry Muffin Model
- Proposed in 1904
- Negatively charged electrons (raisins or blueberries) are surrounded by a positively charged "pudding" (or muffin)



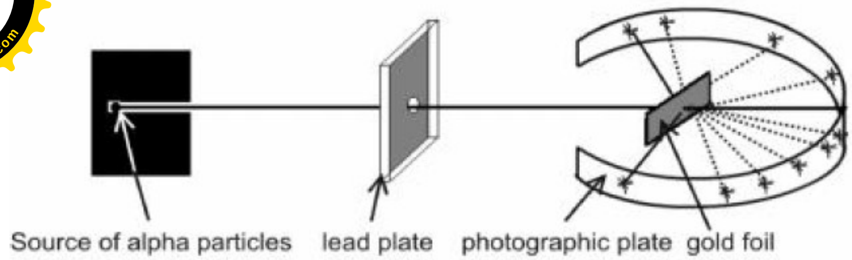
Joseph John

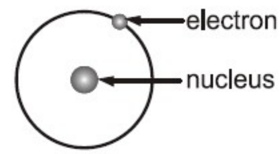


J.J.Thomson
(1856-1940)
Won Nobel prize in Physics in 1906

Ernest Rutherford
(1871-1937)
Won Nobel prize in Chemistry in 1908

Ernest Rutherford performed an experiment called 'Gold Foil Experiment' or 'α- ray scattering experiment' to test the structure of an atom as proposed by Thomson. In this experiment a beam of fast moving alpha particles (positively charged helium ions) was passed through a very thin foil of gold. He expected that the alpha particles would just pass straight through the gold foil and could be detected by a photographic plate. But, the actual results of the experiment were quite surprising. It was observed that most of the α- particles did pass straight through the foil but a number of particles were deflected from their path. Some of these deflected slightly while a few deflected through large angles and about 1 in 1000 α- particles suffered a rebound





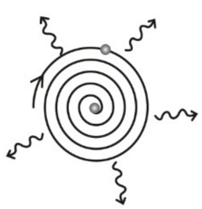
These results led Rutherford to conclude that :

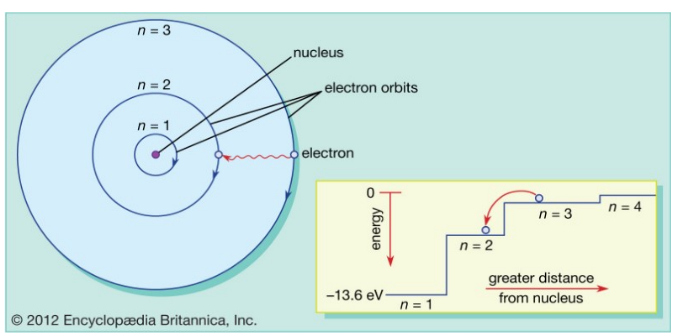
1- the atom contained some dense and positively charged region located at the center of the atom that he called as **nucleus**.

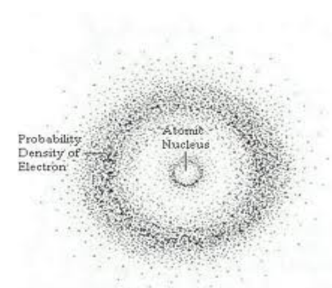
2- all the positive charge of the atom and most of its mass was contained in the nucleus.

3- the rest of the atom must be empty space which contains the much smaller and negatively charged electrons

However, there was a problem with the Rutherford's model. According to the Maxwell's theory of electromagnetic radiation, a charged particle undergoing acceleration would continuously emit radiation and lose energy. Since the electron in the atom is also a charged particle and is under acceleration, it is expected to continuously lose energy. As a consequence, the electron moving around the nucleus would approach the nucleus by a spiral path (Fig. 3.5) and the atom would collapse. However, since it does not happen we can say that the Rutherford's model failed to explain the stability of the atom







The next attempt to suggest a model for atom was made by **Neils Bohr-** a student of Rutherford. Danish physicist who won the Nobel Prize for his work on atom structures and who called for global peace. This model used the concept of quantization of energy of electrons in the atom. Since this fact was suggested by line spectrum of hydrogen atom it is worthwhile to understand the meaning of a spectrum. For this we begin with the understanding of the nature of an electromagnetic radiation

Valence Electron

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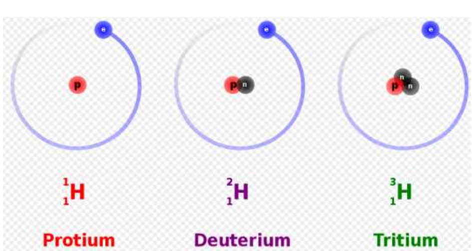
Valence electrons are quite important. They are the electrons that exist in the outermost shell/cloud of an atom at the highest energy level. When atoms interact chemically through bonding, it's their outer shell electrons that are involved. The number of outer shell or valence electrons influences how the element will react by gaining or losing electrons to create stability. The periodic table of elements is arranged to reflect this. Elements with the same number of valance electron

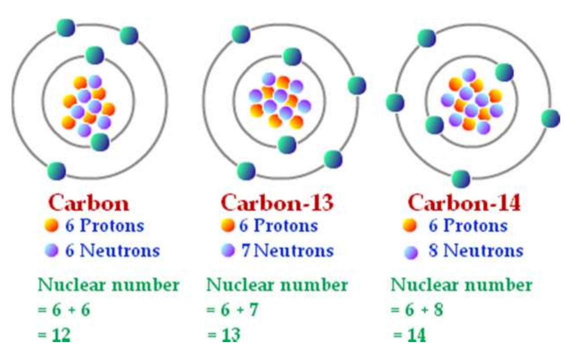
Elements in the periodic table are indicated by SYMBOLS. To the left of the symbol we find the **atomic mass (A)** at the upper corner, and the **atomic number (Z)** at the lower corner.

$$^{A}_{Z}\text{Symbol}$$

Examples: $^{1}_{1}\text{H}$ $^{12}_{6}\text{C}$ $^{23}_{11}\text{Na}$

The number of electrons in a neutral atom (that is, the atomic number) gives the element its unique identity. **No two different elements can have the same atomic number**. The periodic table is arranged by order of increasing atomic number, which is always an integer. In contrast to the atomic number, **different forms of the same element can have different masses**. They are called **isotopes**. The following are representations for some of the isotopes of hydrogen and carbon.





Quantum numbers

- Quantum numbers are Identification numbers for the electrons in Atoms… each electron has 4 quantum numbers make it unique
- 1- principle Q number (n) higher N = higher energy Specifies the energy of an electron and the size of the orbital (the distance from the nucleus of the peak in a radial probability distribution plot). All orbitals that have the same value of *n* are said to be in the same shell (level). For a hydrogen atom with $n=1$, the electron is in its *ground state*; if the electron is in the $n=2$ orbital, it is in an *excited state*. The total number of orbitals for a given *n* value is n^2

n	1	2	3	4	…
shell	K	L	M	N	

- 2- Azimuthal Angular Momentum Q number (l) Specifies the shape of an orbital with a particular principal quantum number. The secondary quantum number divides the shells into smaller groups of orbital called subshells (sublevels). Usually, **a letter code is used to identify *l*** to avoid confusion with : l=n-1 spdf sub shell

n	*l*
1	s
2	s p
3	s p d
4	s p d f
5	s p d f g

- 3-magnetic Q number (ml) Specifies the orientation in space of an orbital of a given energy (*n*) and shape (*l*). This number divides the subshell into individual orbitals which hold the electrons; there are $2l+1$ orbitals in each subshell. This equation will not give you the value of ml, but the number of possible values that ml can take on in a particular orbital. Thus the *s* subshell has only one orbital, the *p* subshell has three orbitals, and so on.

s	p	d	f	<---name of subshell
1	3	5	7	<--- number of orbitals in that subshell
2	6	10	14	<--- number of electrons that will fit into that subshell.

4-Spin Q number (ms) Specifies the orientation of the spin axis of an electron. An electron can spin in only one of two directions (sometimes called *up* and *down*). Pauli exclusion principle