CHEMISTRY

LUANAR - BUNDA

BASIC SCIENCES DEPARTMENT

CHE 111 – 2017/ 2018

ASSESSMENT

- **End of Semester Exam ---- 60%**

- **Continuous (Practicals, Tutorials, Mid Sem. Exam., etc,) --- 40%**

Books on reserve in Library

1. Chang R. (2005). Chemistry, 8th ed.

2. Ebbing D.D. (1990). General Chemistry

3. McMurry J. And Castellion M.E. (2003). Fundamentals of general organic and Biological Chemistry, 4th ed.

Books on reserve

- 4. P.W. Atkins, Clugston M.J., Frazer M.J., and Jones R.A.Y. (1988). Chemistry: Principles and Applications

- 5. Bodner G.M. And Purdue H.L.. (1995). Chemistry an Experimental Science. 2nd ed.
- Prescribed: Peter Mumba. Useful Principles in Chemistry for Agriculture and Nursing students

Atomic theory

- Scientific problems are always checked by experimental work
- If a large number of experiments all give similar results, the results are summarised into a single statement known as a law
- Further experiments are carried out to account for the law. This is done by proposing a hypothesis
- If the results of the experiments agree with the hypothesis, those results are stated into what is called a theory

Atomic theory

- John Dalton (1766-1844)

- 1808: Formulated a precise definition of the building blocks of matter

- All matter is made up of atoms that can not be created or destroyed. i.e., are indivisible by ordinary chemical means

- Atoms of a particular element are identical in size, shape, mass and all other properties and differ from other atoms in these properties

- **A chemical reaction is a union, rearrangement or separation of atoms**

- **This atomic theory was backed by expt. data and so was accepted worldwide**

- **Note: Dalton speculated that an atom could not be divided itself**

- **Scientists were interested in the nature of the atom**

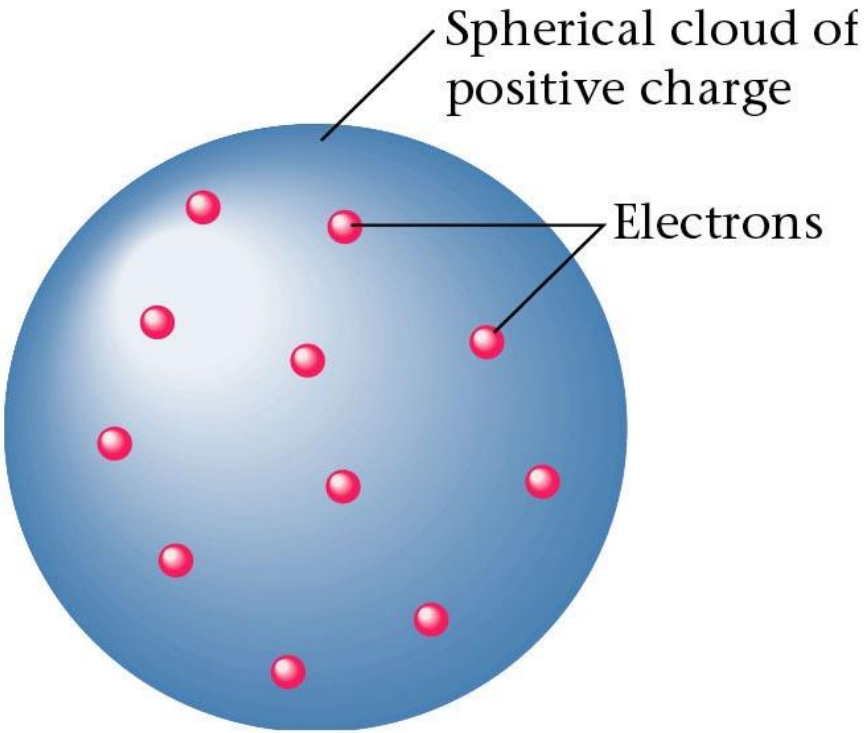
Structure of the atom (Part 1)

- Carried out several experiments
- Results showed that an atom has internal structure
- Has even smaller particles: Subatomic particles
- These expts led to discovery of 3 particles: electrons, protons and neutrons
- Electrons were discovered by J.J. Thomsom in 1897 (1856-1940)

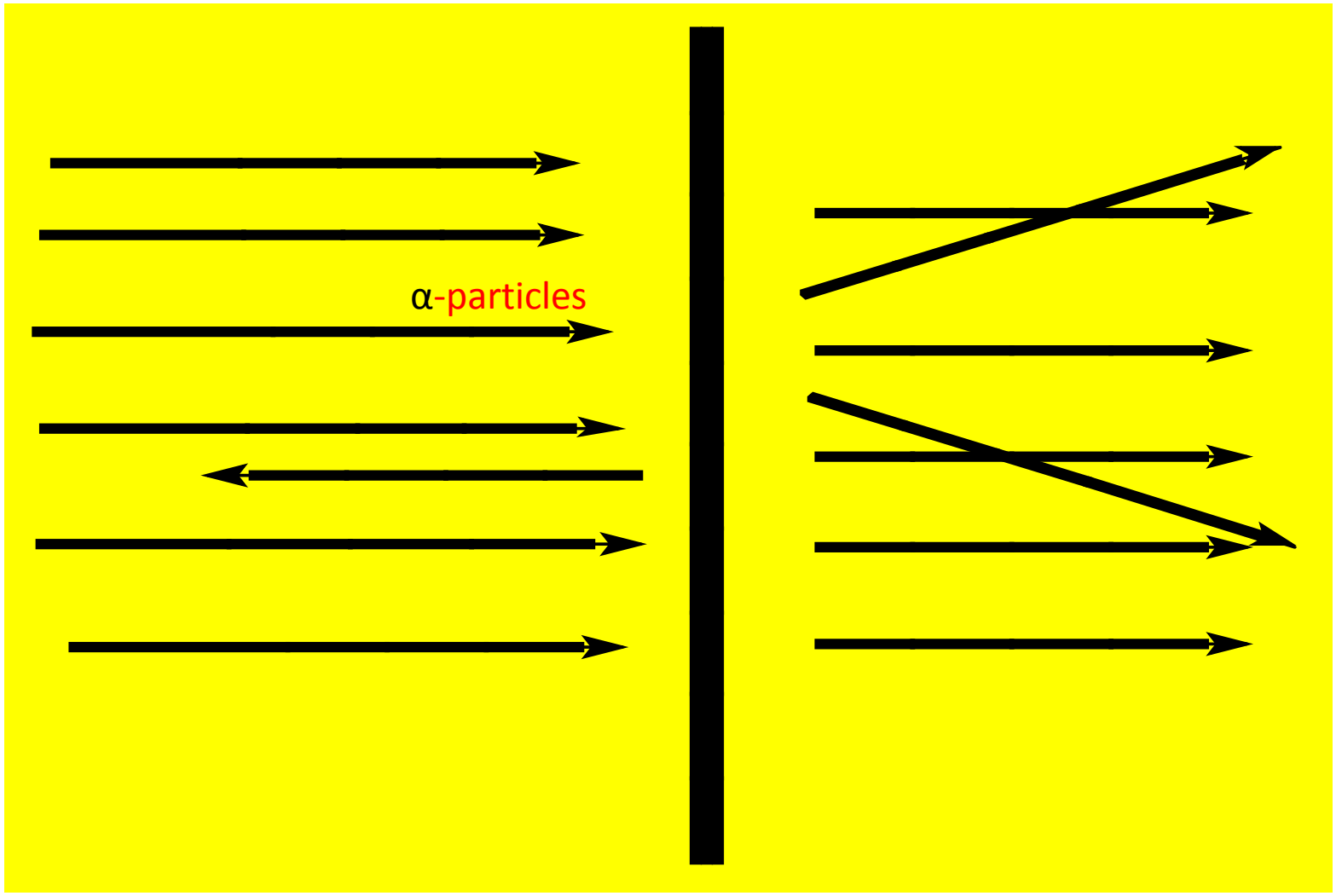
- **Atoms were found to be electrically neutral**
- **It was clear that they must also contain positively charged particles**
- **Experiments showed electrons to be light particles**
- **It seemed proper to associate most of the mass of the atom with the positive charges**
- **Now if most of the mass was with the positive charges, it seemed reasonable that these positive charges must occupy most of the atomic volume**

With this reasoning, J.J. Thomson proposed:

An atom was a uniform sphere of positive charges with the electrons embedded in it so as to maintain electrical neutrality



Rutherford (1908): α-Scattering expt.

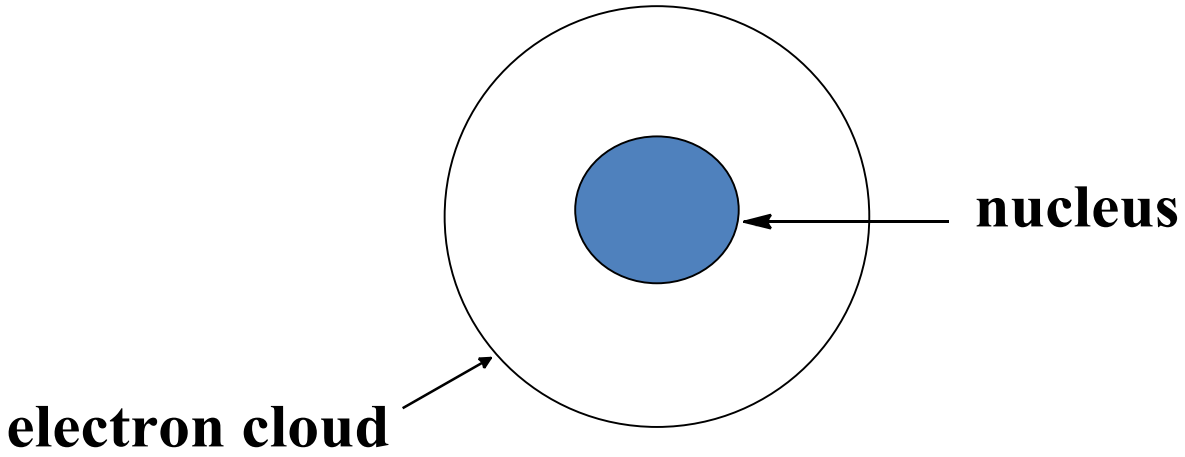


Rutherfords's conclusion

Rutherford thus conceived the nuclear model of the atom

An atom consists of a central nucleus where all the massive positive charges are

Around this, is an empty space where the electrons are



- **The positively charged particles are called protons**
- **Rutherford had predicted that the number of protons was equal to the atomic number (Z)**
  - **Atomic No. = No. of protons**

Scientists had known that H had one proton and He had 2 protons

So they said the ratio of mass to H/He = 1:2

Expts showed the ratio H/He = 1:4

So they said there must be other particles contributing to the mass of He

In 1932, Chadwick performed experiments similar to that of Rutherford and discovered particles that were neutral with mass slightly higher than that of protons

Called them neutrons

He was found to have 2 neutrons and H had none

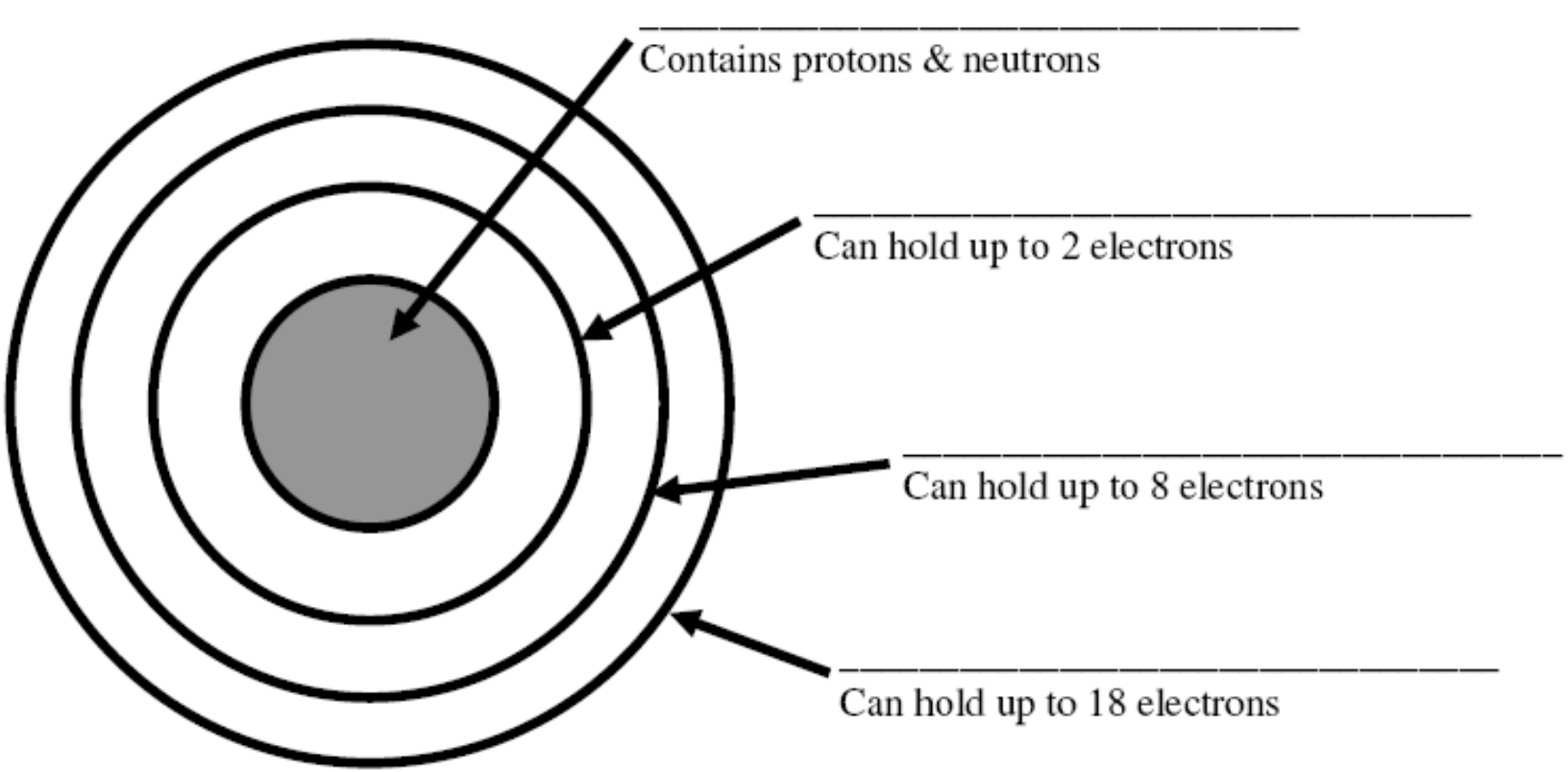
So this explained the 1:4 ratio for mass of H to He

No. of protons (Z) + No. of neutrons (N) give the total mass No. (A)

i.e, $A = Z + N$

The question was "How are the electrons arranged outside the nucleus?"

Electrons are arranged outside the nucleus in shells



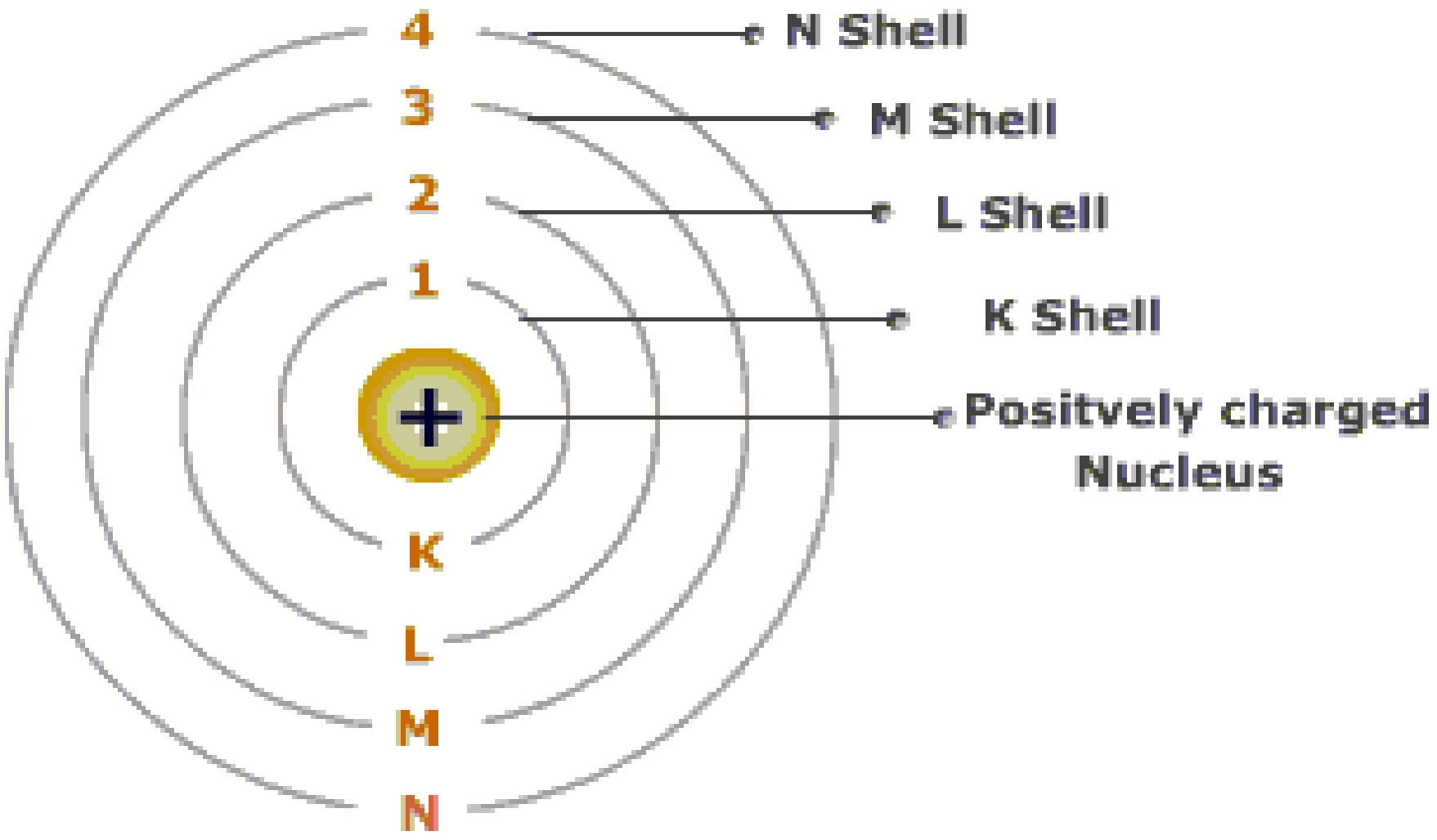
Adapted from http://**www.sciencespot.net/Media/atomsfam.pdf**

Created by G.Baker
www.thesciencequeen.net

we can deduce the No. of electrons in the different shells and get:

Shell No	Maximum number of electrons
1	2
2	8
3	18
4	32
5	50

- **Electrons are added to the shells one at a time across a row**
- **The innermost shell is called the K shell**
- **The second one is called the L shell**
- **The third one is called M shell**
- **So we have K, L, M,N, O shell, etc, in order of increasing size**


4
3
2
1
K
L
M
N
N Shell
M Shell
L Shell
K Shell
Positvely charged
Nucleus

Shell No	Shell Symbol	Maximum number of electrons
1	K	2
2	L	8
3	M	18
4	N	32
5	O	50

More about the atom

- An atom has components (subatomic):
- Electron: Negatively charged
- Charge = -1.60206×10^{-19} Coulombs
- Mass = 9.108×10^{-28} g
- Unit of charge (1.60206×10^{-19} C) is represented by the symbol "e" so -e for electron
- So an electron has a charge of -1

- **Proton:** Positively charged
- Charge = $+1.60206 \times 10^{-19}$ Coulombs
- Mass = 1.672×10^{-24}g
- Unit of charge is represented by the symbol "e" so, +e for proton
- So a proton has a charge of +1
- **Neutron:** Neutral
- Mass = 1.675×10^{-24}g

The atomic unit is called the Atomic mass unit (amu)

This is derived to measure atomic mass

1 amu = 1.66×10^{-24}g

We can thus convert masses of electron (in g), proton, etc, to their masses in amu as follows:

$$Mass\ of\ an\ electron = \frac{9.108 \times 10^{-28}}{1.66 \times 10^{-24}}$$

$$= 0.000549\ amu$$

Similarly:

Mass of proton = 1.00725 amu

Mass of neutron = 1.00904 amu

Note: Mass of proton = ~2000 x that of electron

Atomic mass units are rounded off to whole numbers called mass numbers (A)

So mass of electron = o amu, proton= 1 amu, neutron= 1 amu

Mass No (A) = No. protons (Z) + No. Neutrons (N)

ie : A = N + Z

Atomic weights

- An atom is a very small particle
- Difficult to determine its absolute mass
- Possible to determine the mass of an atom relative to another
- What you get is the **Relative Atomic Mass** of the atom **(RAM)**
- If a formula of a compound is known, it is possible to determine the RAMs of the atoms involved and vice versa
- Problem of early 19^{th} century was how to determine atomic weights and formulae *simultaneously* since both were unknown

Dalton's approach:

"If there is no information to the contrary, make the simplest possible assumption and pursue its consequences"

Dalton's "rule of greatest simplicity" : If two elements form only one compound with each other, a molecule of this compound contains only one atom of each element

- **Dalton carried out several experiments, one involved the reaction of hydrogen with oxygen**

- **In order to determine RAMs of atoms, Dalton assigned a mass of 1 unit to the simplest atom, hydrogen and obtained masses of other atoms relative to H.**

- **In 1805, water was the only known compound of H and O, so its formula was taken as HO (by the ''rule of greatest simplicity'')**

- **Dalton found that 8g of O reacted with 1g of H**

- **The assumption that the formula for water was HO, forced Dalton to conclude that the RAM of O was 8 on a scale which took the mass of H as unity (1)**

- **Note: Dalton's assumptions were wrong in some cases and so gave false conclusions**
- **Note: This method of determining RAMs involves ratio of masses**
- **In this case, 8 is the RAM of O relative to that of H**
- **It does not say anything about the actual mass of O**

- **The relative masses of atoms are usually referred to as atomic weights**

- **The atomic weight scale was originally based on the mass of 1 unit for H**
- **Later, and up to now, Carbon -12 was/is taken as a standard**

- **For technical reasons, it is the most convenient practical standard in determining RAMs**

Definition of atomic weight

The weighted mean of the masses of the isotopes of the element relative to the mass of a carbon-12 taken as 12.0 exactly

E.g., Chlorine has 2 isotopes, ^{35}Cl with a relative abundance of 75.5% and ^{37}Cl with an abundance of 24.5%

$$Ram\ of\ Cl = \frac{35 \times 75.5}{100} + \frac{37 \times 24.5}{100} = 35.49$$

Exercise

Gallium consists of 60% ^{69}Ga and 40% ^{71}Ga. Calculate average atomic weight of Ga

By the beginning of the last century it has been possible to express the mass of an atom in grams

Eg., Given that 1g of hydrogen contains 6 x 10^{23} H atoms. Calculate the mass of one Hydrogen atom

$$6 \times 10^{23}\ atoms\ weigh\ 1\ g$$

$$\therefore Weight\ of\ 1H\ atom = \frac{1\ H\ atom \times 1g}{6 \times 10^{23} H\ atoms}$$

$$= 1.67 \times 10^{-24}\ g$$

Note: this is the wt of a proton

We don't usually express masses of atoms in grams for two reasons:

Figures such as 1.67×10^{-24} are cumbersome as the atoms are very light

Chemists knew masses of different atoms relative to each other long before they knew their masses relative to standard gram

Since 1g of hydrogen contains 6×10^{23} H atoms and mass of one hydrogen atom is 1 amu, it follows that:

$$1g = 6 \times 10^{23} amu$$

The actual figure is 6.0226×10^{23}

6.0226×10^{23} is called Avogadro's constant (or Avogadro's number)

Elements and compounds

- A substance which contains only one kind of atom is called an element
- Or
- A substance all of whose atoms have the same Atomic No.

	Hydrogen	Chlorine	Oxygen	Sodium	Potassium	Helium
Symbol	H	Cl	O	Na	K	He

- **Elements are indicated by symbols that are a combination of letters**
- **The first letter is always capitalized**
- **The second or third letters are lower case**
- **E.g. Co is the symbol for the element Cobalt whereas CO is molecular formula for carbon monoxide**

- **An element cannot be decomposed into two or more distinct substances by chemical means**

- **Two or more elements can combine chemically to make a compound**

- **A compound is substance which contains more than one kind of atom**

	Water	Hydrogen Chloride	Methane	Carbon dioxide	Glucose
Mol. Formula	H_2O	HCl	CH_4	CO_2	$C_6H_{12}O_6$

- A compound may be decomposed chemically into two or more distinct elements

Pure substances and mixtures

- **Both elements and compounds are pure substances**
- **E.g., Cane sugar consists of only sugar molecules, so it's a pure substance**
- **Common salt (NaCl) consists of only sodium chloride "molecules". It's also a pure substance**
- **A pure substance has a definite or constant composition**

- **As such, a pure substance has the same properties, e.g., Mps, Bps, density, chemical behaviour, etc**

- **Now if salt is added to sugar, we have a mixture**
- **Mixtures do not have definite or constant composition like pure substances**

- **You can vary the composition of a mixture in any way you like**
- **The properties of a mixture depend on the composition of the mixture**
- **E.g. 2g salt in 50ml water and 5g salt in 50 ml water will have different densities**

Types of mixtures

- Two types exist
- **Heterogeneous:** Where the particles of each component are visibly present
- **E.g.** Salt + black pepper or sand + sugar

- **Homogeneous (also called a solution):** Where particles of one component are evenly dispersed amongst those of the other
- **E.g.** Salt in water: You only see one liquid layer

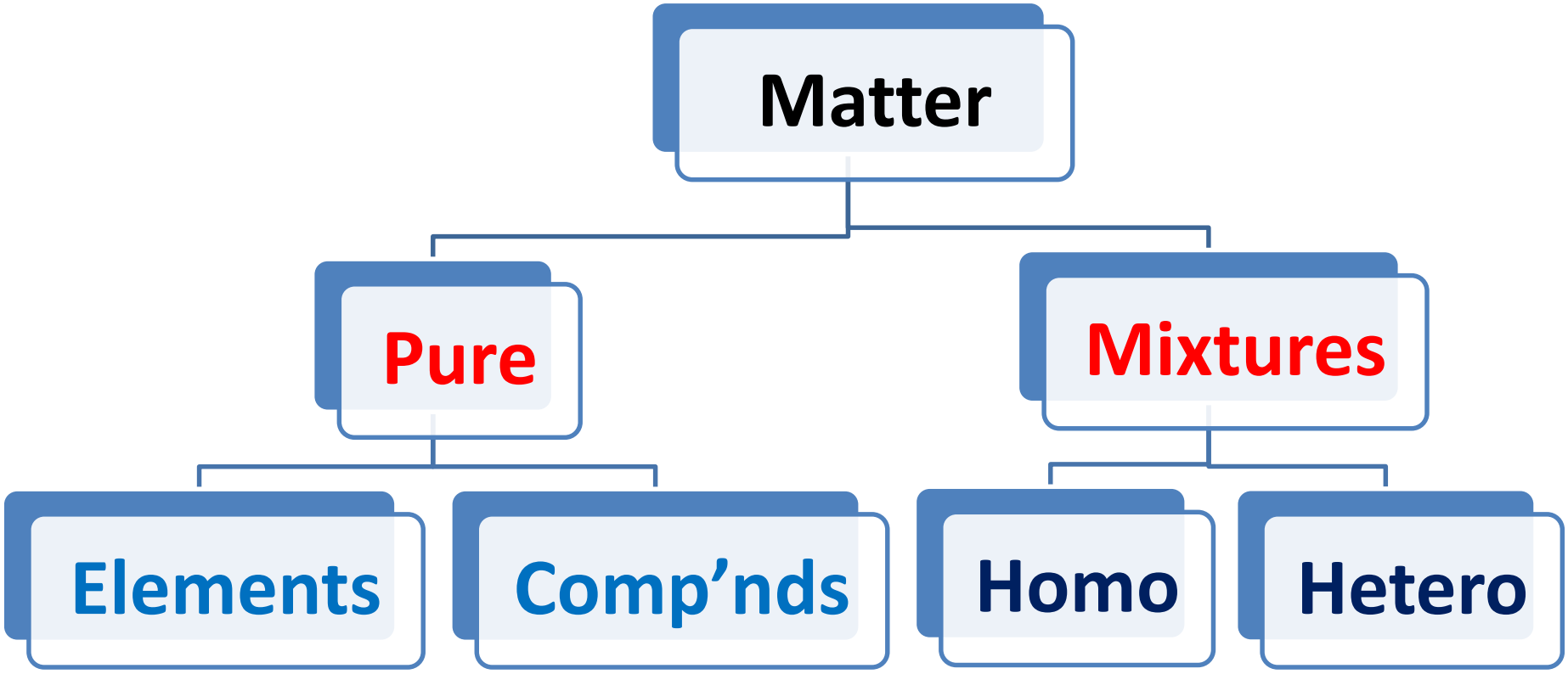
Types of solutions

- **Gaseous**: e.g., Air
- **Liquid**: e.g., Salt + water or ethanol + water
- **Solid**: e.g., Brass (Zinc + Copper)

- Liquid solutions are most common

- In preparing liquid solutions, usually, a small quantity of one substance is added to a large quantity of another

- **The smaller quantity is called the solute**
- **The larger quantity is called the solvent**
- **The ratio of the quantity of solute to the quantity of solvent is called the concentration of the solution**


Matter
Pure
Mixtures
Elements
Comp'nds
Homo
Hetero

Structure of the atom (Part 2)

- **Classical:** Physics developed before 1900

- **Eg.** Newton's Laws of motion

- In the 19th Century, Some physicists believed that the theoretical structure of physics was complete

- **Various experimental results obtained in late 19th century could not be explained by classical physics**

- **Led to the development of another branch of physics**

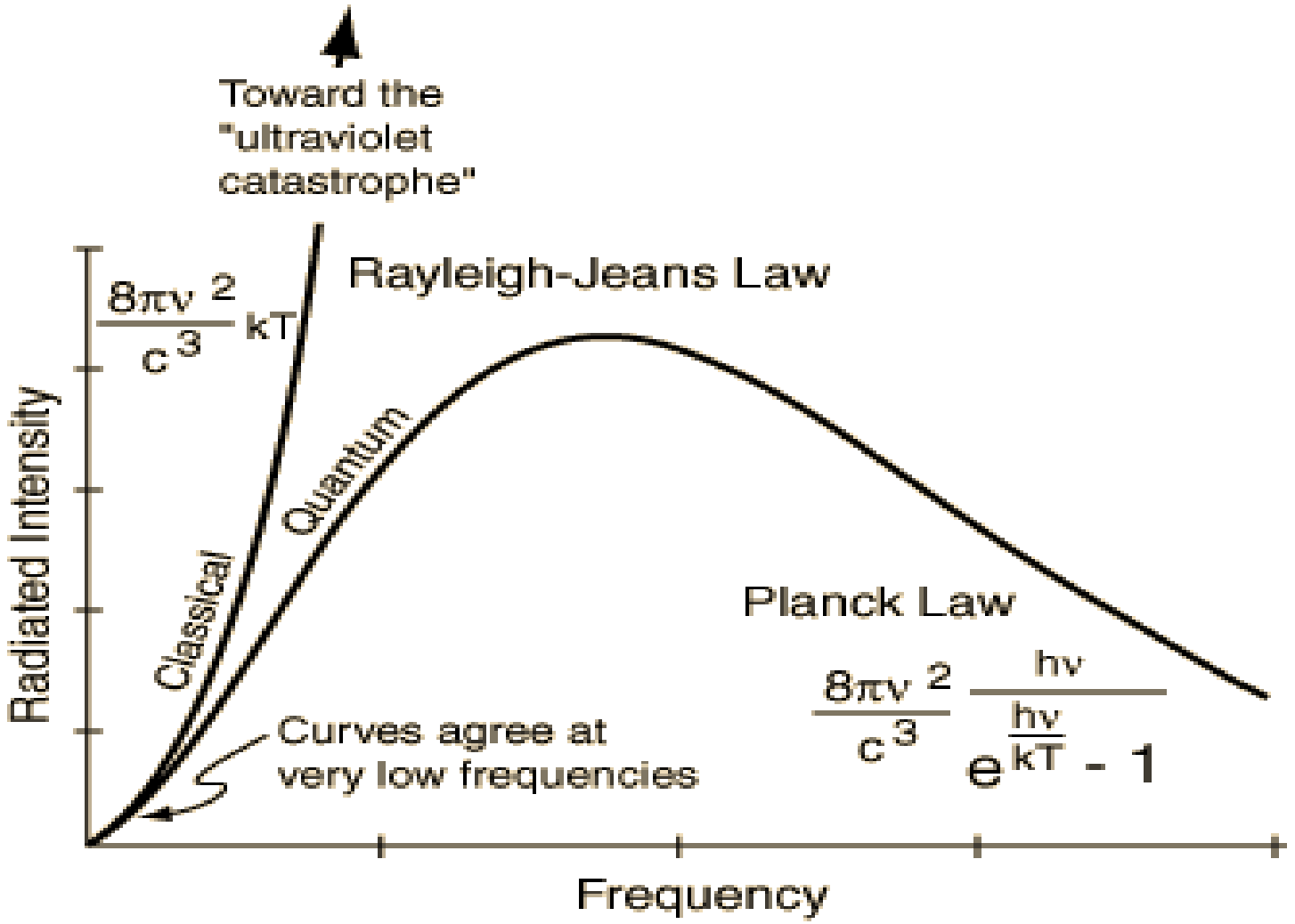
- **Failure of classical physics:** **Inability to explain the observed frequency distribution of radiant energy emitted by a hot solid**

- **Experiments showed that radiation from a heated solid was distributed at different frequencies**

- **As T increases, the maximum in energy shifts to higher freqs.**

- **When a metal rod is heated it first glows red, then orange yellow, then white, then blue-white**

In 1900, German physicist, Max Planck derived an equation that fitted exactly to the observed curves



- **Planck assumed that the radiating atoms can emit energy only in discrete quantities or in form of a bundle**

- **He called the smallest bundle of energy a quantum of energy (L. *Quantum*= how much)**

- **This then gave rise to another branch of physics, quantum physics and quantum theory**

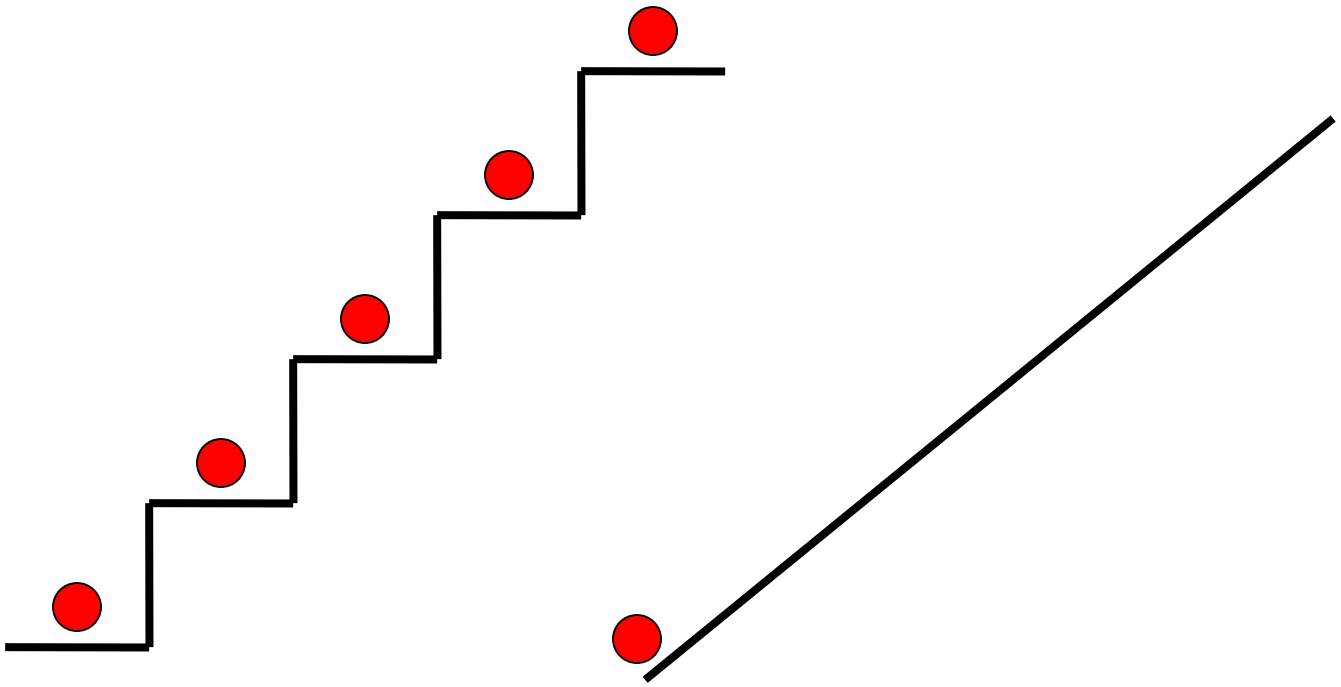
- **Note: A quantum of energy is the amount of energy required to move an electron from one energy level to another**

- **What do we learn from quantum physics?**

- **In Classical physics: energy takes on a continuous range of values and a system can lose or gain any amount of energy**

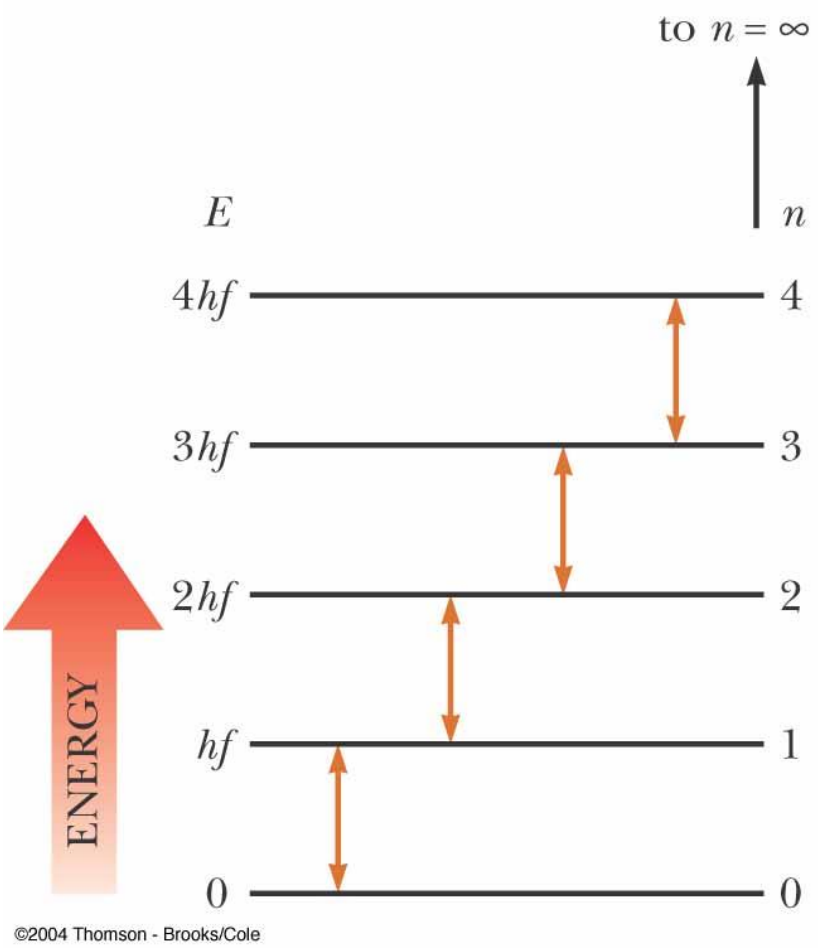
- **In quantum physics: energy is quantized, meaning that the energy can take on only certain values**

Energy can have only certain values (quantities), not in between, instead of a continuum of values. This is like energy existing on stairs of a staircase instead of at any energy on a ramp.



Energy – level diagram

- An energy-level diagram shows the quantized energy levels and allowed transitions
- Energy is on the vertical axis
- Horizontal lines represent the allowed energy levels
- The double-headed arrows indicate allowed transitions



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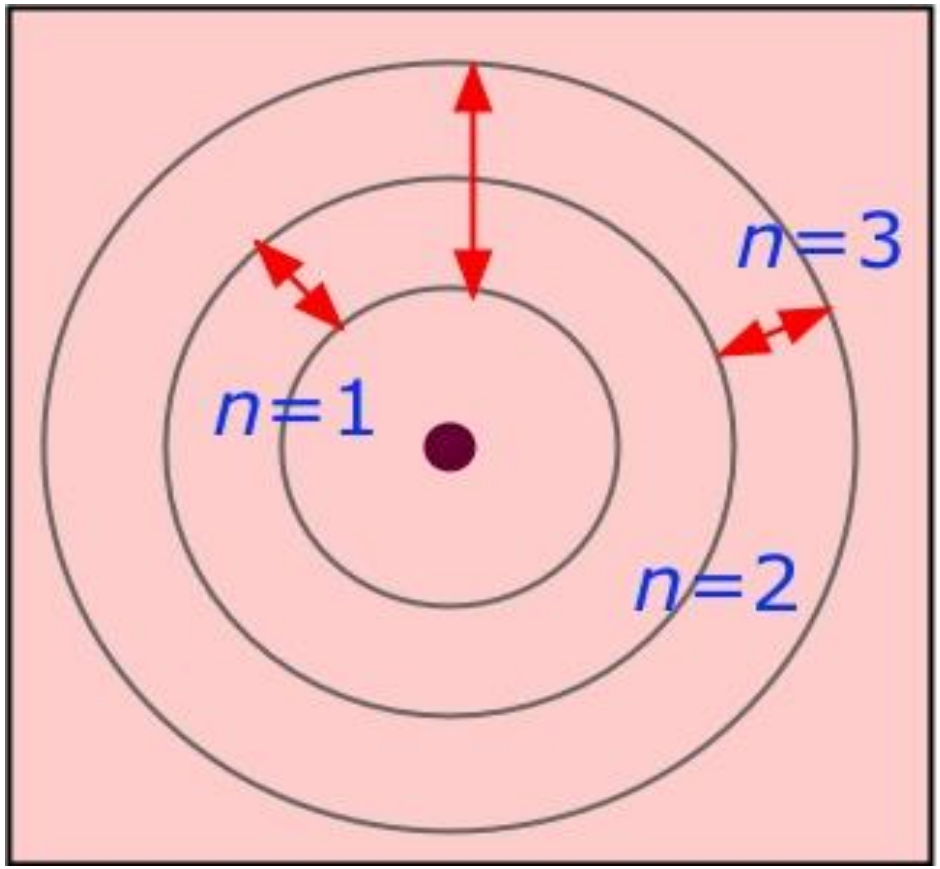
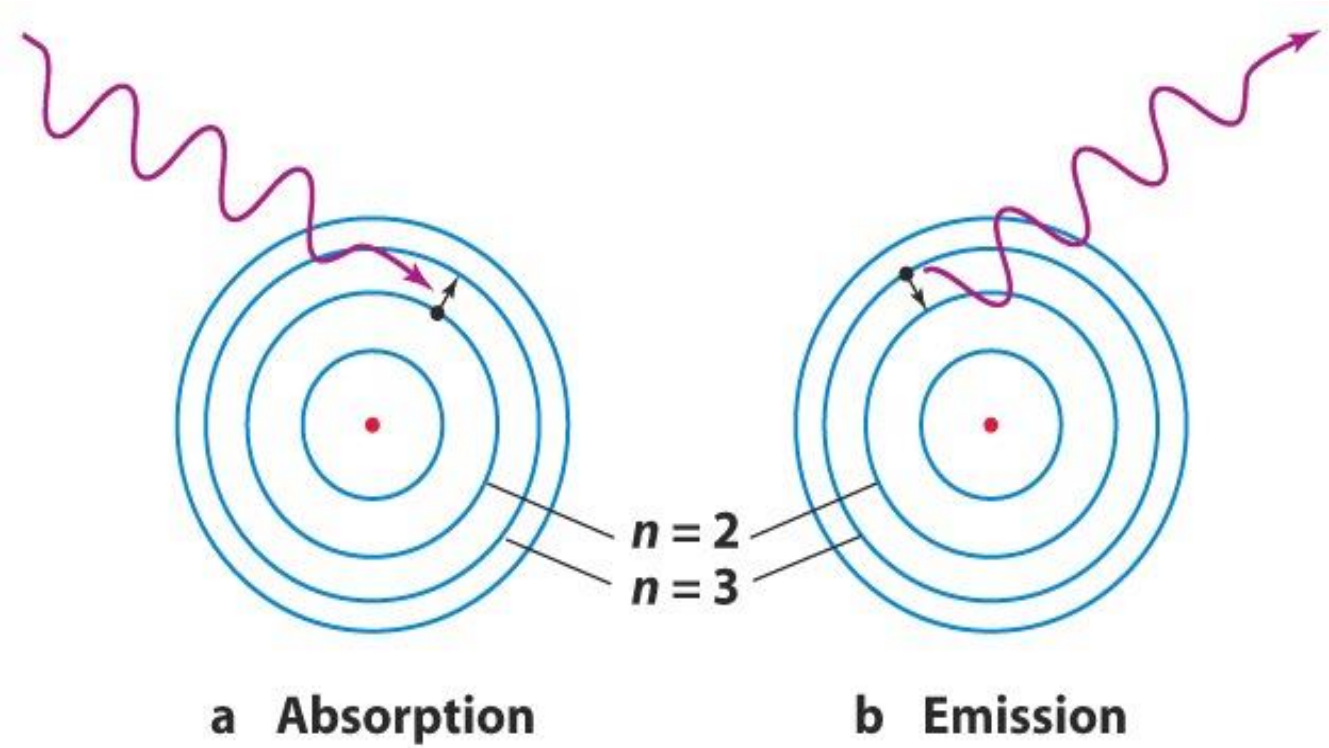
Neils Bohr's (1913)theory

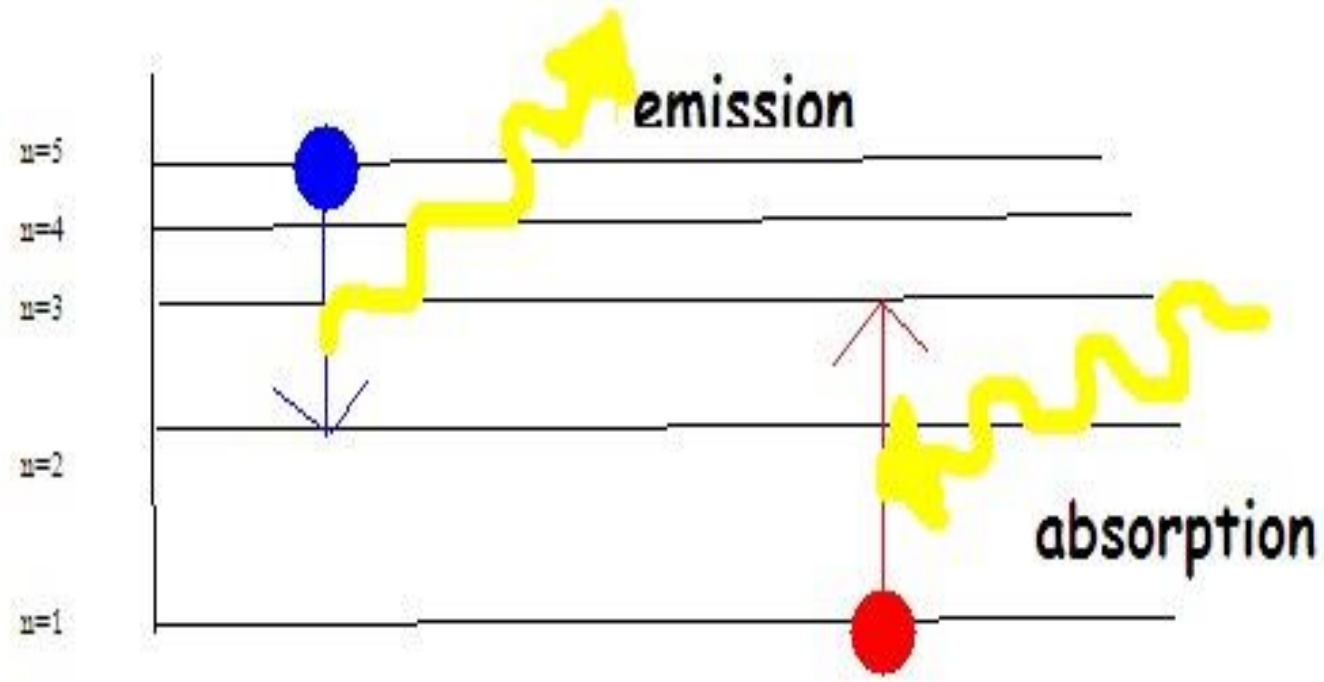
- An electron moves in circular orbit with a fixed energy
- An electron in hydrogen can be found only in certain orbits
- Since each orbit has a particular energy associated with it, Bohr's restriction meant that the energies of the electron in the orbits are quantized
- To explain the line spectrum, Bohr stated that an electron can only move from one energy level to another


n=3
n=1
n=2

Electron moving from lower energy to higher energy level absorbs energy

Electron moving from higher to lower energy level releases energy





Absorption and emission of photons as a result of energy level change of electrons

- The energy released (or absorbed) is governed by an integer called Principal quantum No. (n)
- Principal quantum numbers are 1,2,3,4, etc
- Least energy level is called ground state
- All the other levels are called excited states
- Electrons move round the nucleus as a cloud
- The region of space in which the electron spends at least 95% of its time is called an orbital
- Orbitals are arranged round the nucleus in shells
- Each shell has a definite No. of orbitals

For each "n" value there are n^2 orbitals

Shell No. (pqn, n)	Shell symbol	No. Orbitals (n^2)	Max No. of electrons ($2n^2$)
1	K	1	2
2	L	4	8
3	M	9	18
4	N	16	32
5	O	25	50

- There is 1 orbital in the K shell, 4 orbitals in the L shell, 9 orbitals in the M shell etc.
- The 4 orbitals in the L shell are not of same energy, so too the 9 orbitals in the M shell (except H)
- The L shell has 2 energy levels
- Lower energy has one orbital, higher has 3 orbitals (degenerate)
- (Degenerate orbitals are orbitals that have the same energy.)
- The M shell has 3 energy levels

- **The lowest has 1 orbital, the second has 3 and the highest has 5 orbitals**
- **Within a shell:**
- **Lowest energy level consist of one orbital and is called the s orbital**
- **The next energy level consist of 3 degenerate orbitals called p orbitals**
- **The third energy level consist of 5 degenerate orbitals called d orbitals**

- **The 4^{th} energy level consist of 7 degenerate orbitals called f orbitals**
- **Order of orbital energies: s<p<d<f**
- **How do electrons go into the various orbitals?**

- **Pauli exclusion Principle**
- **States: There can never be more than 2 electrons in any atomic orbital. This means an orbital can have 0, 1 or 2 electrons but no more than 2**

Shell symbol	Shell No. (Principal quantum No. n)	Orbitals	No. of Orbitals	Total No. Orbitals (n^2)	Max. No. of electrons ($2n^2$)
K	1	S	1	1	2
L	2	S P	1 3	4	8
M	3	S P d	1 3 5	9	18
N	4	S P d f	1 3 5 7	16	32
O	5	S, p, d, f, g		25	50

Order of electron filling

- **1s<2s<2p<3s<3p<4s<3d<4p<5s<4d<5p<6s<4f<5d, etc**

- **In describing ground state electron configuration, the guiding idea is that the *total energy* of the atom is as low as possible. To determine these configurations, we use the Aufbau Principle as a guide**

- **Aufbau ("building up") Principle: a scheme used to reproduce the electron configurations of the ground states of atoms by successively filling subshells with electrons in a specific order (the building-up order).**

Note that the sum of the superscripts equals the number of _electrons_ in the atom.

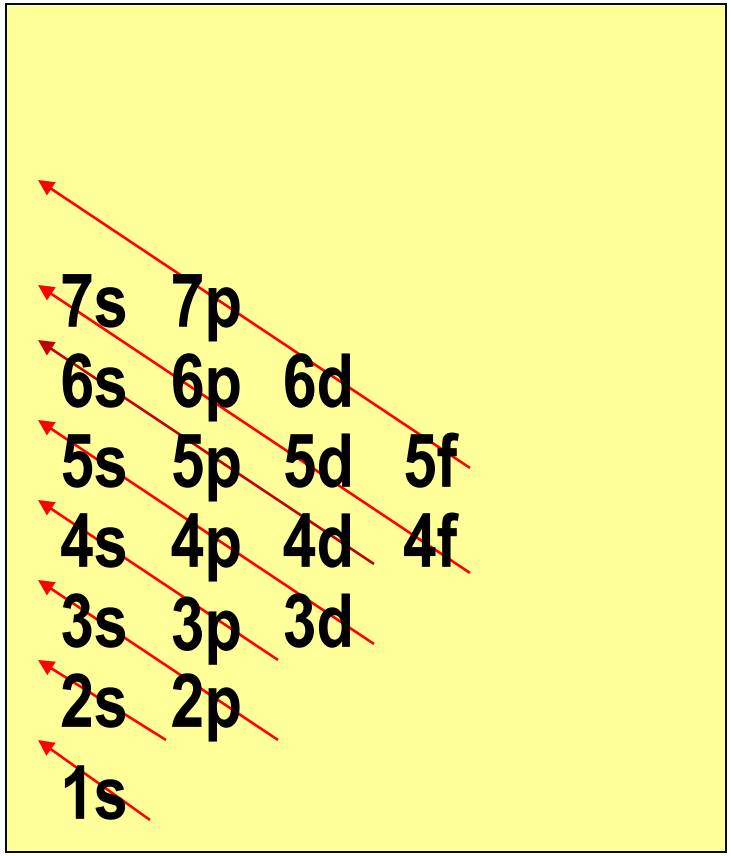
Electron configuration: the arrangement of electrons of an atom in its ground state into various orbitals around the nuclei of atoms.

Aufbau ("building up") Principle

(some exceptions do occur).

Sub-level maxima:

s = 2 e-
p = 6 e-
d = 10 e-
f = 14 e-
…



Electronic structures

- **Show the n values and the orbitals occupied together with the No. of electrons in them**

- **Simply put: they show the distribution of electrons in the various orbitals**
- **E.g, in H, the electron goes in the s orbital of Principal Quantum No. 1**

- **So electronic structure for H = $1s^1$, for He=$1s^2$**

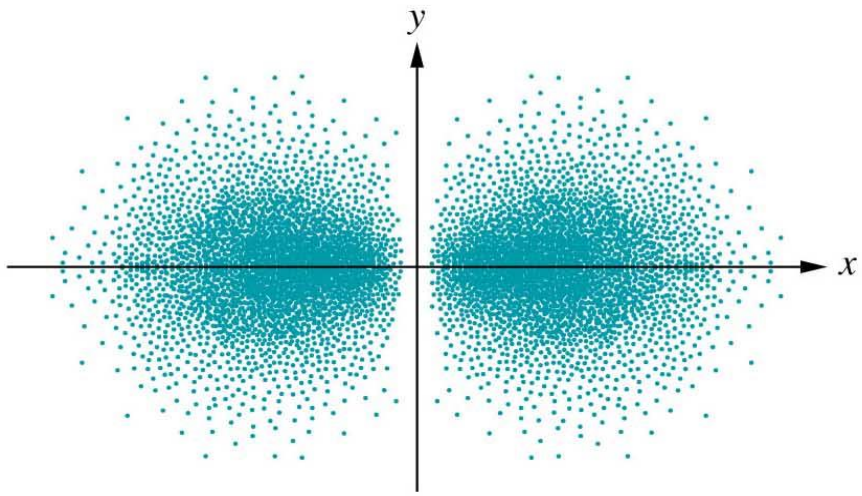
- **Note: the sum of superscripts = Total No. of electrons = Atomic No. for a neutral atom**

- **So if you know the Atomic No., You can write the electronic structure for any atom.**

More on electronic structures

- In carbon there are 2 electrons in the p orbitals
- i.e, $1s^2 2s^2 2p^2$

- Remember there are three p orbitals
- How exactly do the two electrons go in the three degenerate p orbitals?
- Let's first look at the shapes of these 3 orbitals



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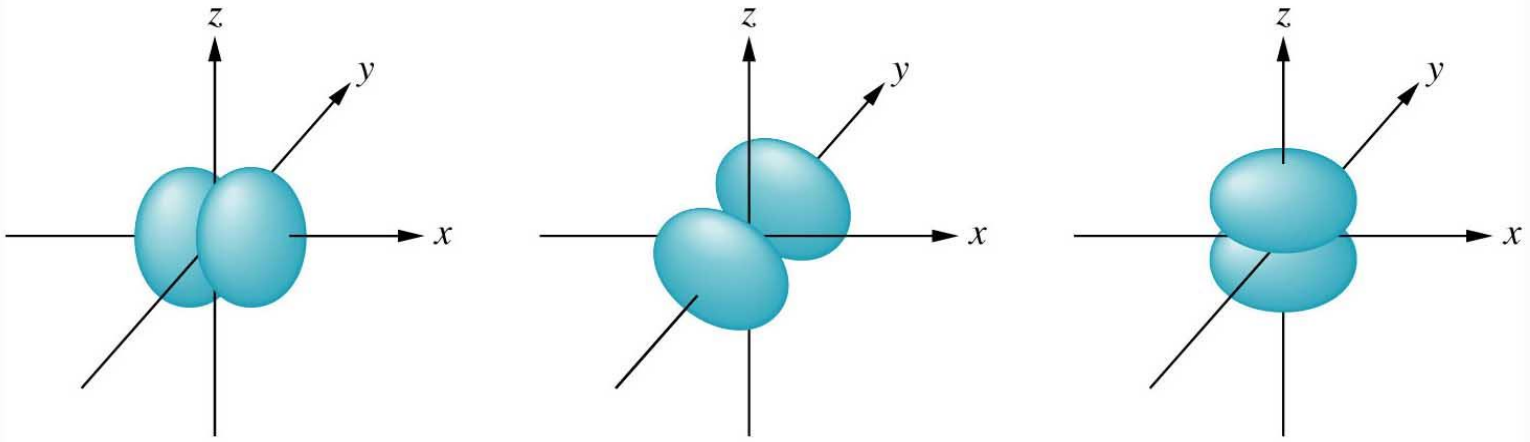


$2p_x$ orbital $2p_y$ orbital $2p_z$ orbital

B

- **There are 2 possibilities in which the two electrons can go in the 3 orbitals**

either $1s^2 2s^2 2p_x^{\,2} 2p_y 2p_z$

or $1s^2 2s^2 2p_x^{\,1} 2p_y^{\,1} 2p_z$

- **Hund's Rule (Rule of maximum multiplicity)**
- **Two electrons will avoid occupying the same orbital as long as there are other degenerate orbitals still empty**

- **So the correct structure is** $1s^2 2s^2 2p_x^{\,1} 2p_y^{\,1} 2p_z$

Paramagnetism and diamagnetism

- Atoms such as H ($1s^1$), O ($1s^2 2s^2 2p_x^{11} 2p_y^1 2p_z^1$), etc have unpaired electrons in their structures
- Such atoms tend to be attracted to a magnet
- Are said to be **paramagnetic**

- Atoms such as He ($1s^2$), Be ($1s^2 2s^2$) have all electrons paired
- Such atoms tend to be repelled by a magnet
- Are said to be **diamagnetic**

Exercise

1. Element X and Y have atomic numbers 54 and 98 respectively.

   a. Write electron configuration of X and Y

   b. Identify which is paramagnetic and which is diamagnetic

The Periodic Table

- **The Periodic Table arranges elements in order of increasing Atomic No.**
- **If elements are arranged in order of increasing Atomic No., elements with similar electronic structures recur periodically**

- **The periodic recurrence of similar electronic structures leads to periodic recurrence of properties (Chemical and physical)**

- **Elements in a Periodic Table are arranged in 2 ways**
- **Horizontal rows are called Periods**
- **Such elements are arranged in order of increasing Atomic No.**
- **Elements with similar electronic structures are arranged in vertical columns called Groups or Families**

- **Elements in groups 1a and 2a all fill the s orbital. They are called s-block elements**

- **Elements in groups 3a-8a fill the p orbitals. They are called p-block elements**

- **Elements in groups 1b-8b fill the d orbitals. They are called d-block elements (transition elements)**

- **Elements at the bottom (two blocks) fill the f orbitals. They are called f-block elements (also called rare earth or inner transition elements)**

- **Elements in the Periodic Table are also grouped according to the type of orbital filled**

- **Representative elements (main group elements): Group 1a-7a: Have incompletely filled s or p orbitals of highest principal quantum No.**

- **Noble gases (except He): Have completely filled p orbitals**

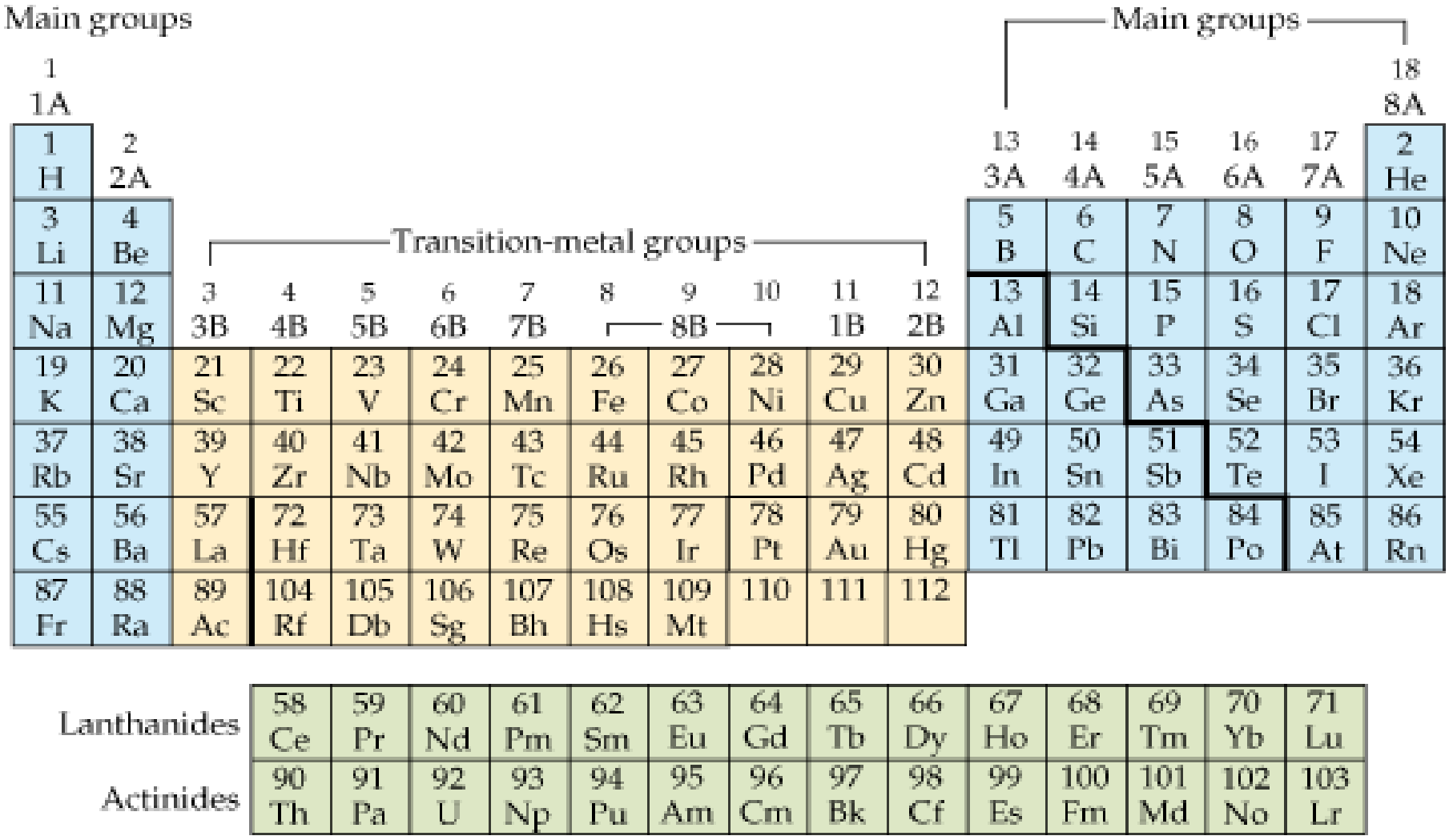
- **Transition elements (groups 1b, 3b-8b): Have incompletely filled d orbitals**

- **Group 2b elements are usually excluded from transition elements**

- **F-block elements have incompletely filled f orbitals**

- **Sum of electrons in the highest pqn = group no. for that element**

- **Highest pqn in electronic structure= period in which that element appears**


Main groups
Main groups
Transition-metal groups
1 1A
2 2A
3 3B
4 4B
5 5B
6 6B
7 7B
8
9 8B
10
11 1B
12 2B
13 3A
14 4A
15 5A
16 6A
17 7A
18 8A
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
57 La
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
89 Ac
104 Rf
105 Db
106 Sg
107 Bh
108 Hs
109 Mt
110
111
112
Lanthanides
58 Ce
59 Pr
60 Nd
61 Pm
62 Sm
63 Eu
64 Gd
65 Tb
66 Dy
67 Ho
68 Er
69 Tm
70 Yb
71 Lu
Actinides
90 Th
91 Pa
92 U
93 Np
94 Pu
95 Am
96 Cm
97 Bk
98 Cf
99 Es
100 Fm
101 Md
102 No
103 Lr

Group 1a elements

- **Have one outer shell electron, i.e., the s^1 structure (except H)**

- **Are called ALKALI elements (H is an exception)**
- **<u>Physical properties</u>**
- **Are malleable**
- **Soft metals**
- **Conductors of heat and electricity**

- **Chemical properties**

- Ist 3 react violently with water
- E.g., $2Na + 2H_2O \rightarrow 2Na^+OH^- + H_2(g)$

- React with hydrogen to form hydrides
- E.g., $2Li + H_2 \rightarrow 2LiH$

- React with oxygen to from oxides
- E.g., $2Li + 1/2O_2 \rightarrow Li_2O$
- $4Na + O_2 \rightarrow 2Na_2O$

Group 2a elements

- Have 2 outer shell electrons, i.e., ns^2 structure
- Are called ALKALINE EARTH elements
- <u>Physical properties</u>
- Harder than alkalis
- Denser than alkalis
- Conductors of heat and electricity
- Malleable

- **Chemical properties**

- **Lose the 2 outer shell electrons**
- **Ca + energy → Ca^{2+} + $2e^-$**

- **At High temp, they react with water**
- **$Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2(g)$**

- **Be : No reaction**
- **Mg : Reacts at high temp**
- **Ca, Sr: React with cold water**

- **React with hydrogen to form hydrides**

- **Eg., Ba $+H_2 \rightarrow BaH_2$**

Group 7a Elements

- Have 7 outer shell electrons, i.e., the ns^2np^5 structure
- Are called halogens
- Physical properties
- In nature under normal conditions, exist as diatomic molecules
- F_2, Cl_2 are gases at room temp

- Br_2 is liquid
- I_2 is solid
- At is radioactive
- **Chemical properties**
- React with hydrogen to form hydrides
- $F_2 + H_2 \rightarrow 2HF$

- **Form halides when they react with other halogens**

- **F_2, Cl_2, Br_2, I_2**

- **BrF, ClF, BrI, IF_5, BrF_3, ClF_3, IF_7**

Group 8a Elements

- Have 8 outer shell electrons, i.e., ns^2np^6 structure
- Except He which has s^2 structure
- Called rare or noble gases
- Exist in atmosphere in small amounts, hence rare
- Tend to be inert, hence inert gases

- **Exist in nature as single-atomed molecules**

- **This suggests that they are un reactive and have no tendency to combine with themselves or with other atoms**

- Electron configurations show that gp 8a elements have completely filled outer shells
- Indication of great stability
- In early 60s a number of compounds were formed
- E.g., XeF_6, XeF_4, KrF_4, $Xe(OH)_6$, XeO_3

- Since compounds of some have been detected, the name "inert" has been replaced by the name "noble"

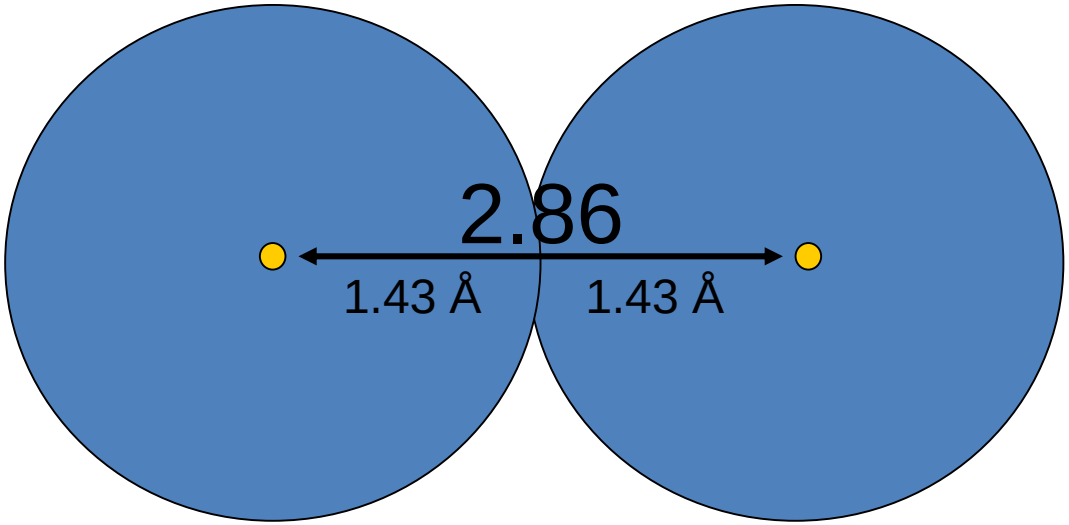
Some atomic properties and the periodic table

- **Atomic radii (Sizes of atoms)**
- **Ionization energy (IE)**
- **Electron affinity**
- **Electronegativity**

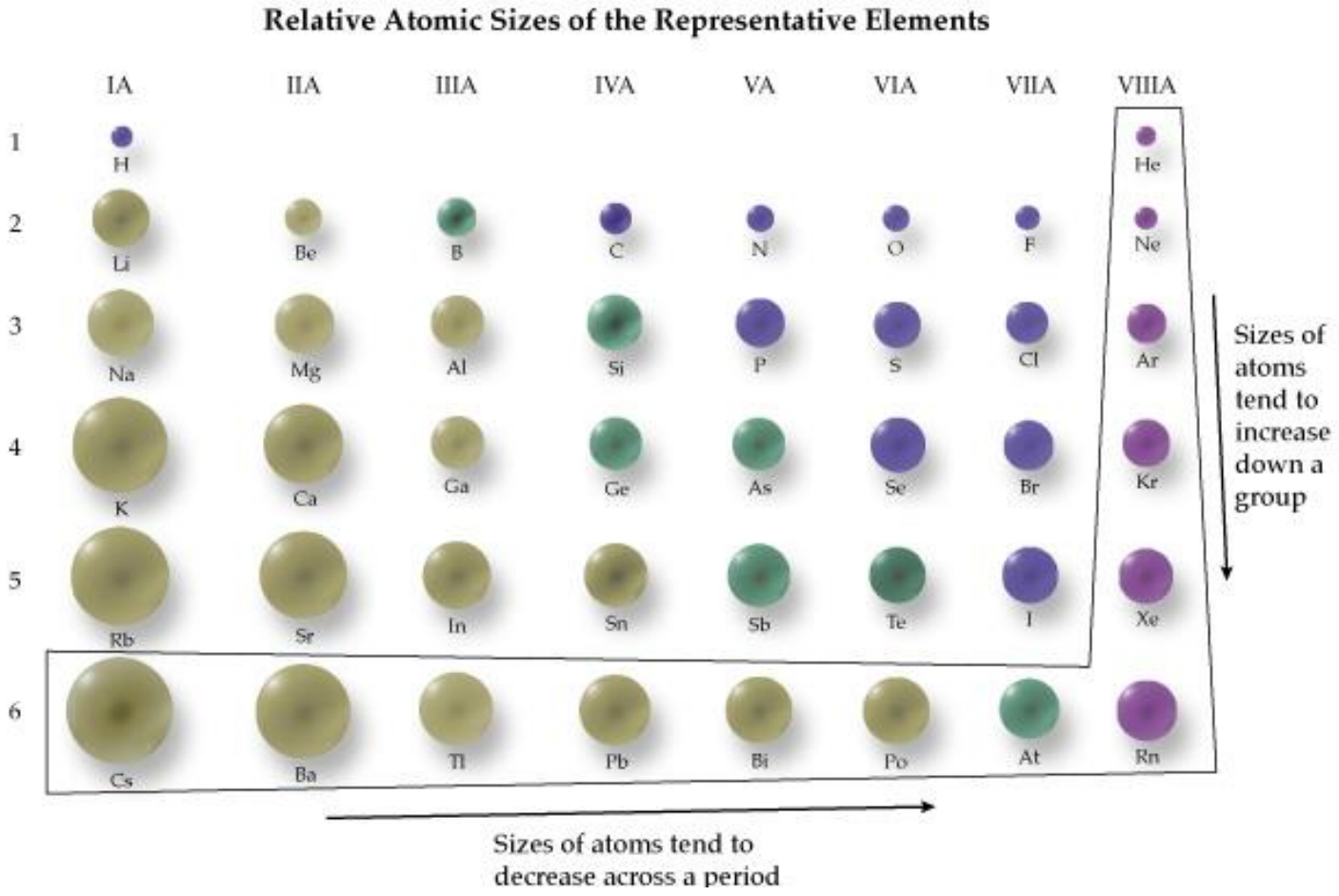
Atomic radius

- **Radius is the distance from the center of the nucleus to the “edge” of the electron cloud**

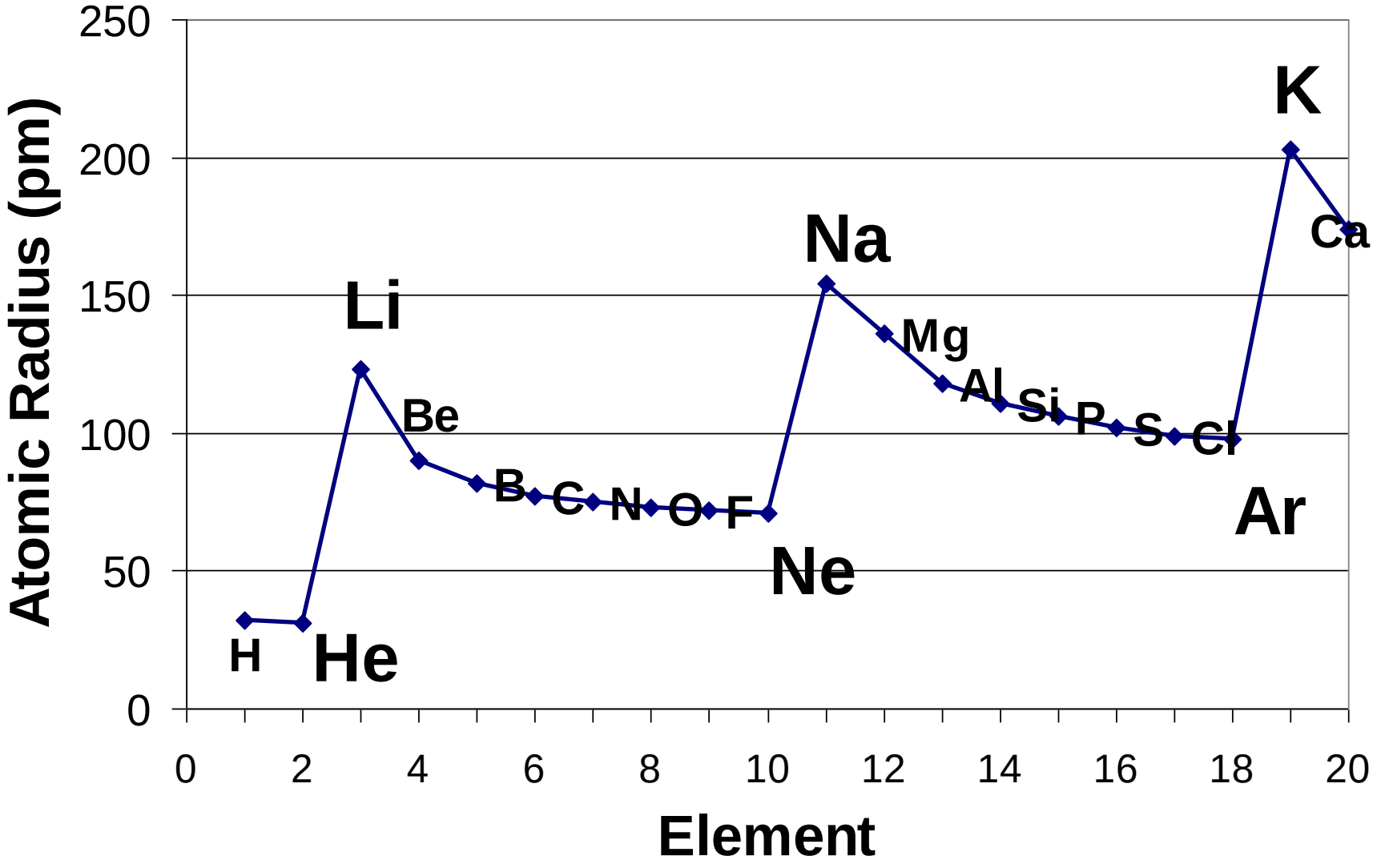
- **Since a cloud’s edge is difficult to define, scientists define covalent radius, or half the distance between the nuclei of 2 bonded atoms.**



In general, atomic size increases from top to bottom within a group, and decreases from left to right across a period.



Atomic Radius versus atomic number



- **Atomic radii increase as we go down a group**

  - **Why?**

- **With each step down the family, we add an entirely new energy level (shell) making the atoms larger with each step.**

- **Atoms get smaller as you go across a period.**

  - **Why?**

- **Each step adds a proton and an electron (and 1 or 2 neutrons).**
- **Electrons are added to existing energy levels.**
- **The nucleus is more positive and the electron cloud is more negative.**
- **The effect is that the more positive nucleus has a greater pull on the electron cloud.**
- **The increased attraction pulls the cloud in, making atoms smaller as we move from left to right across a period.**

- **Effective nuclear charge:**

- **Is the pull that an electron "feels" from the nucleus.**

- **The closer an electron is to the nucleus, the more pull it feels.**

- **As effective nuclear charge increases, the electron cloud is pulled in tighter.**

- **Shielding effect:**

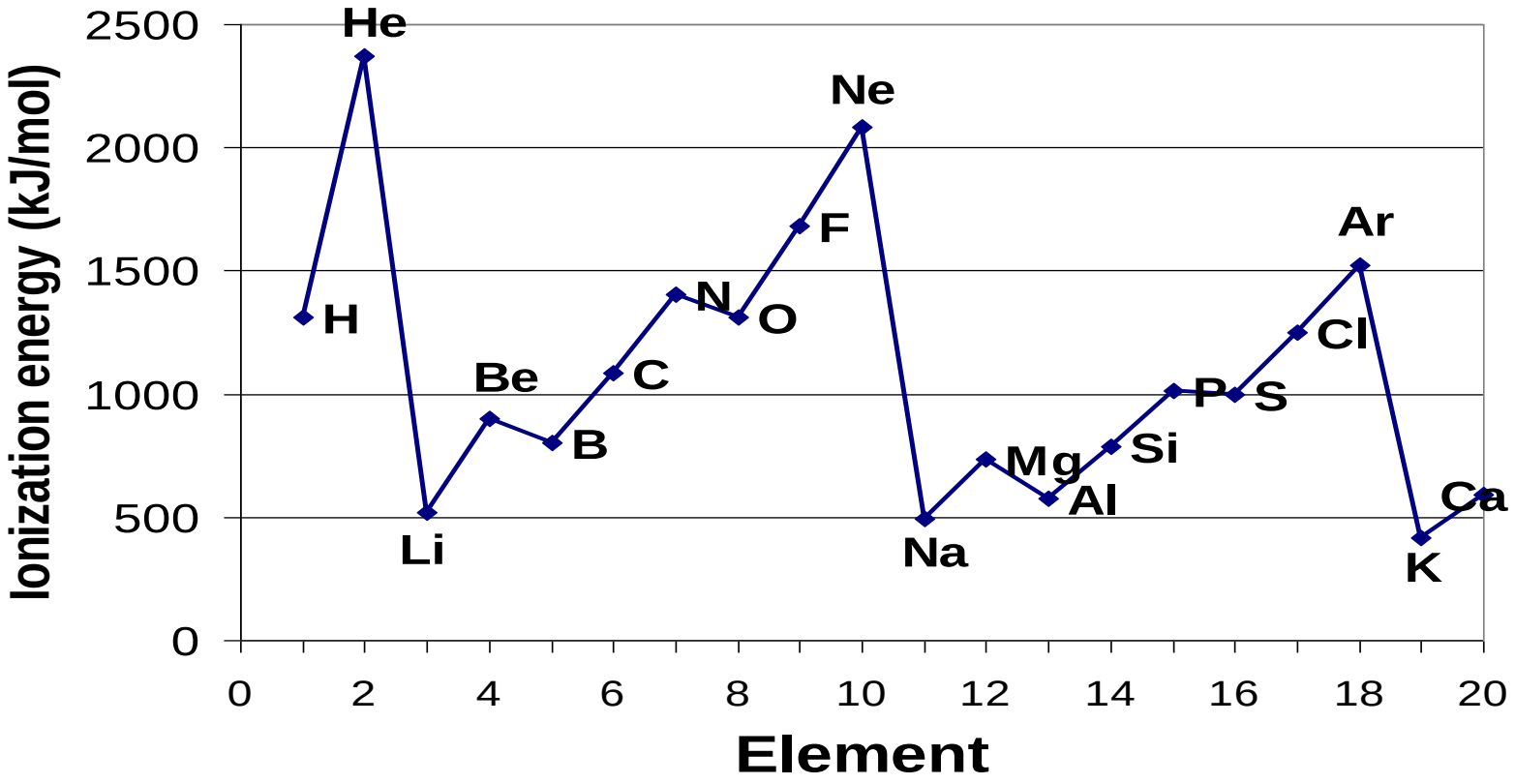
- **As more Principal Energy Levels are added to atoms, the inner layers of electrons shield the outer electrons from the nucleus.**

- **The effective nuclear charge (enc) on those outer electrons is less, and so the outer electrons are less tightly held.**

Ionization energy (IE)

- **Minimum energy required to remove an electron from a gaseous atom**
- **Within a period, the 1st IE increases from gp 1a to gp 8a**
- **Due to ENC which increases from left to right**
- **Within a group 1st IE decreases from top to bottom**
- **Sizes of atoms increase down a group**

Ionisation energy versus atomic number



Note: **An irregularity occurs between Group 2A and 3A**

Another irregularity occurs between Groups 5A and 6A

- **The smaller the IE the greater the metallic character**

- **Electron affinity**

- **The energy released when an electron is added to a gaseous atom**

- **$F(g) + electron \rightarrow F^- + energy$**

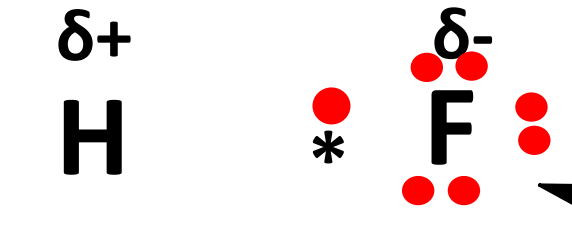
- **There is not much data on electron affinities**

- **Generally, electron affinities tend to increase across a period**

- Non metals tend to have higher values than metals

- **Electronegativity**
- This is a measure of the electron attracting power of an atom in a bond
- It is evident in heteronuclear molecules, e.g. HF
- In HF, the atoms have different electron attracting powers
- F attracts electrons more than H

- Electrons spend more time closer to F than to H
- i.e.,



Low electron density

High electron density

- So there is an electron density shift toward F
- The HF molecule is said to be polar
- The molecule is polarized

- **Electronegativities increase from left to right across a period**
- **Due to increasing Effective Nuclear Charge**
- **Electronegativities decrease down a group**
- **In the Periodic table the most electronegative elements are on the top right hand corner and the least, are on the bottom left hand corner**
- **Elements with low values tend to be metals**
- **Elements with high values tend to be non metals**