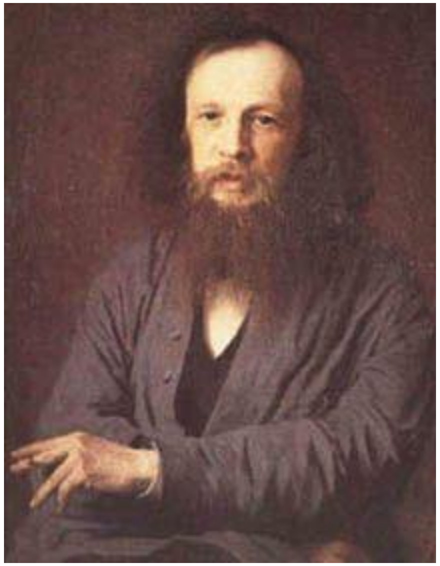
Development of the periodic table



Dimitri Mendeleev (1834-1907)

- 62 elements were known at the time
- developed a method for organizing elements based on their atomic mass
- recognized periodicity in the physical-chemical properties
- rows = periods
- columns = groups

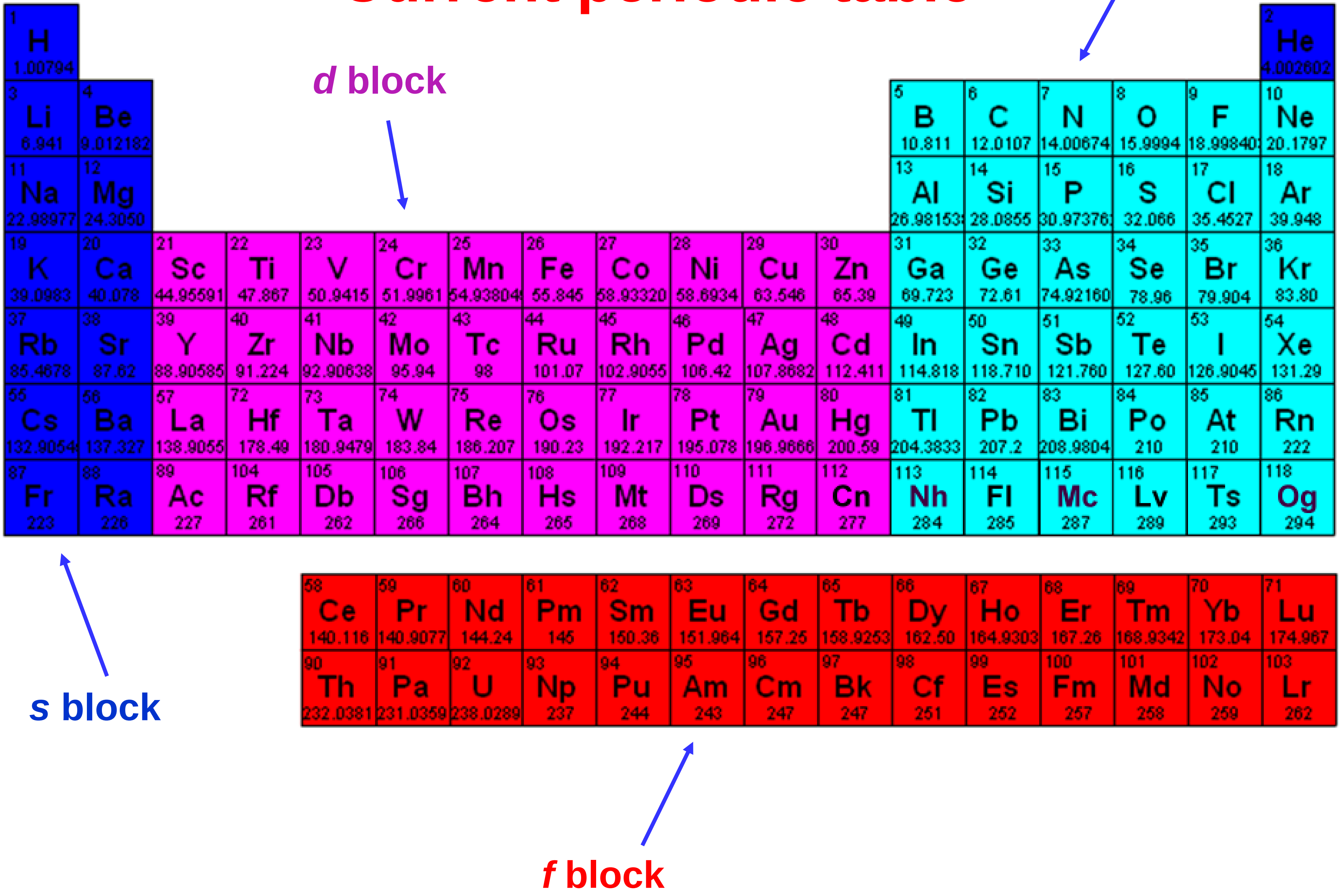
- found the flaws (inversions) in the monotonous order of atomic weights (Co-Ni, Te-I)
- predicted the existence of undiscovered elements and left their places blank in his periodic table

	ekasilicon	Ge
atomic mass	72	72.59
density (g/cm^3)	5.5	5.35
melting point (°C)	high	947
color	grey	grey
density of oxide (g/cm^3)	4.7	4.7
chemistry of oxide	slightly basic	slightly basic
boiling point of chloride	below 100 °C	86 °C ($GeCl_4$)
density of chloride (g/cm^3)	1.9	1.9

ekasilicon: Ge   ekaaluminum: Ga

ekaboron: Sc   ekamanganese: Tc   trimanganese: Re

Current periodic table



p block

d block

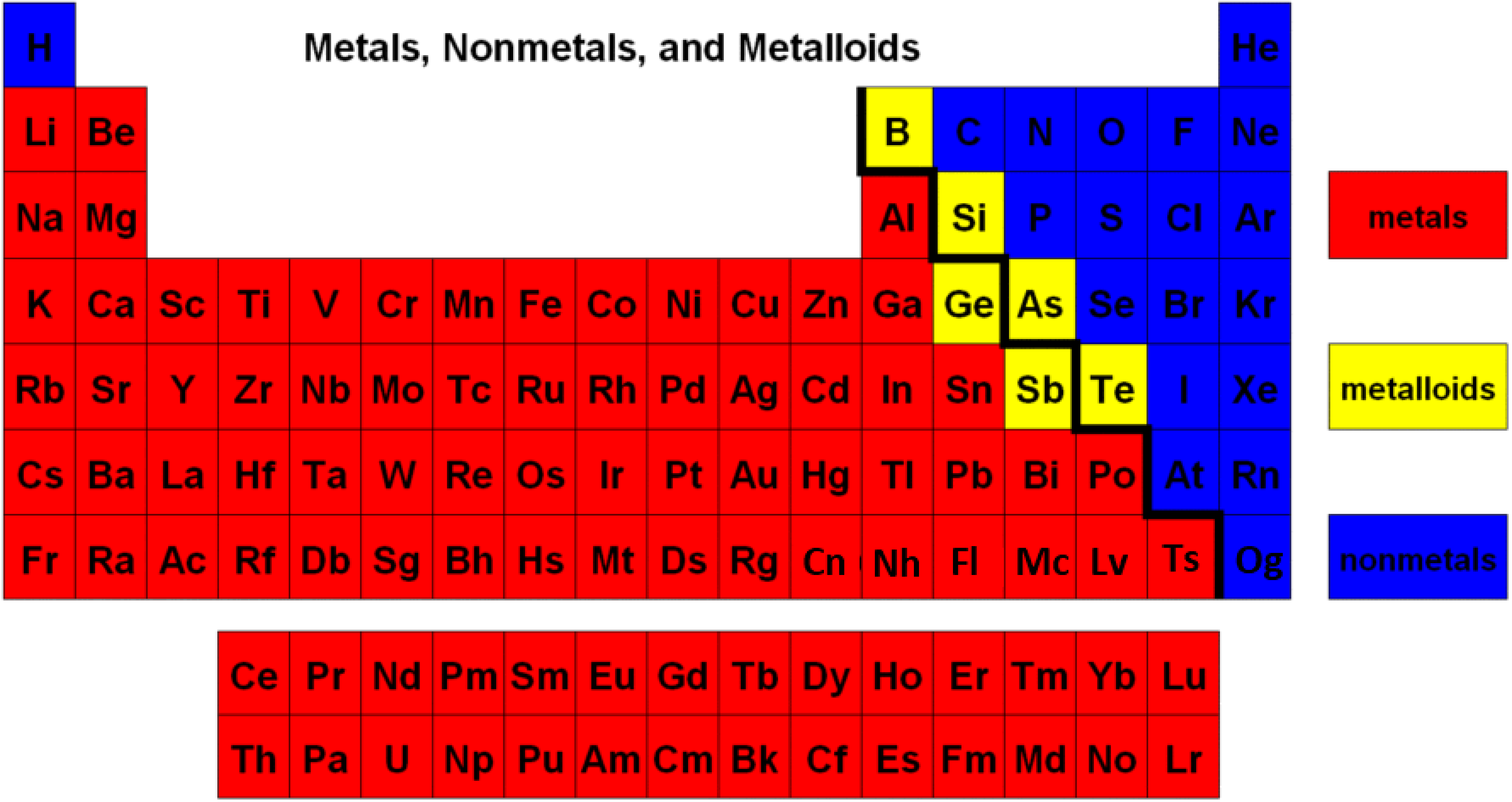
s block

f block

1 H 1.00794																	2 He 4.002602
3 Li 6.941	4 Be 9.012182											5 B 10.811	6 C 12.0107	7 N 14.00674	8 O 15.9994	9 F 18.99840	10 Ne 20.1797
11 Na 22.98977	12 Mg 24.3050											13 Al 26.98153	14 Si 28.0855	15 P 30.97376	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.95591	22 Ti 47.867	23 V 50.9415	24 Cr 51.9961	25 Mn 54.93804	26 Fe 55.845	27 Co 58.93320	28 Ni 58.6934	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.92160	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.90585	40 Zr 91.224	41 Nb 92.90638	42 Mo 95.94	43 Tc 98	44 Ru 101.07	45 Rh 102.9055	46 Pd 106.42	47 Ag 107.8682	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	53 I 126.9045	54 Xe 131.29
55 Cs 132.9054	56 Ba 137.327	57 La 138.9055	72 Hf 178.49	73 Ta 180.9479	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.217	78 Pt 195.078	79 Au 196.9666	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 208.9804	84 Po 210	85 At 210	86 Rn 222
87 Fr 223	88 Ra 226	89 Ac 227	104 Rf 261	105 Db 262	106 Sg 266	107 Bh 264	108 Hs 265	109 Mt 268	110 Ds 269	111 Rg 272	112 Cn 277	113 Nh 284	114 Fl 285	115 Mc 287	116 Lv 289	117 Ts 293	118 Og 294

58 Ce 140.116	59 Pr 140.9077	60 Nd 144.24	61 Pm 145	62 Sm 150.36	63 Eu 151.964	64 Gd 157.25	65 Tb 158.9253	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 231.0359	92 U 238.0289	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

Current periodic table



Metals, Nonmetals, and Metalloids

H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

🟥 metals

🟨 metalloids

🟦 nonmetals

90-95 metals 6-8 metalloids ~20 nonmetals

What is the natural law behind the periodic table?

Answer:

- not obvious in 1869
- early 20th century: atomic structure

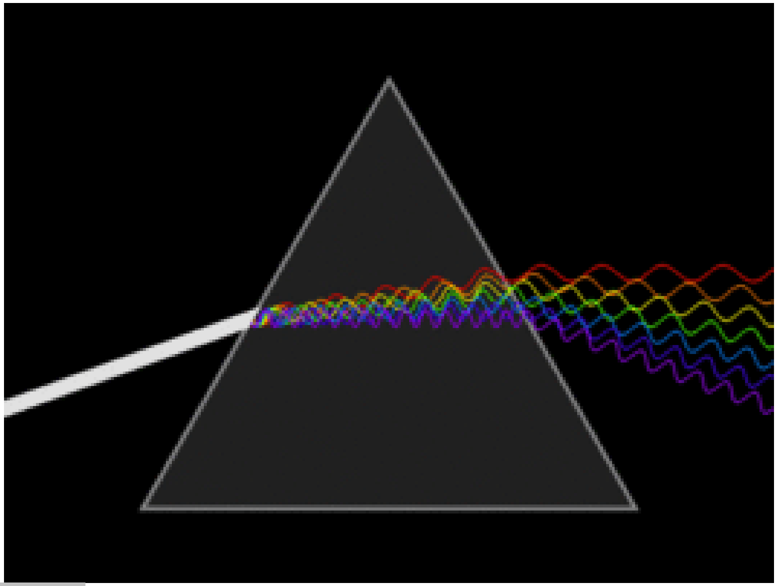
Important experimental information:

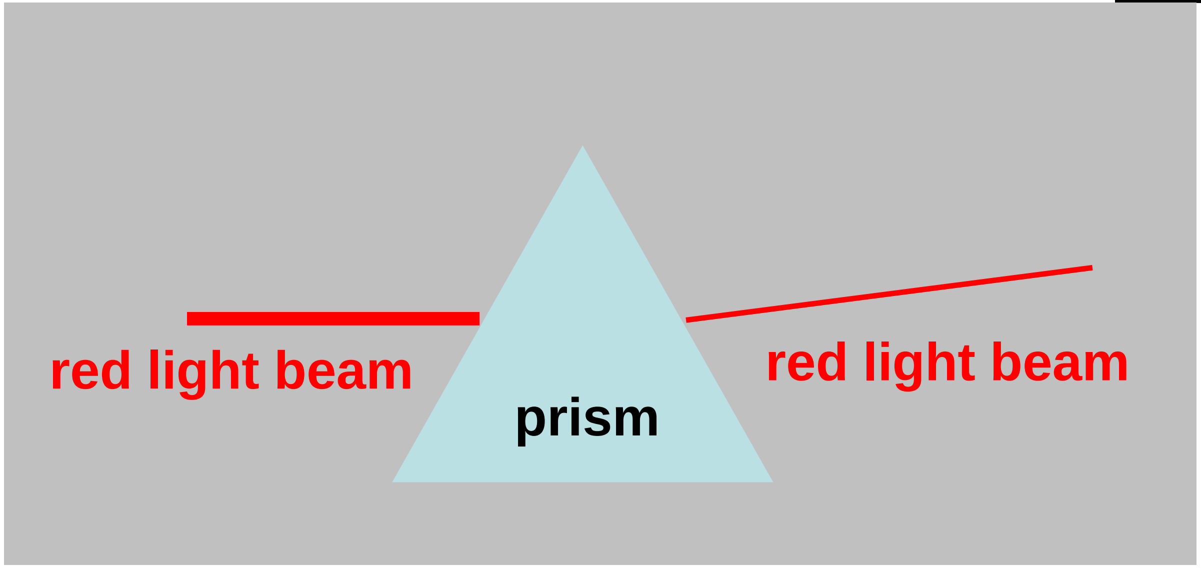
- the properties of light
- the interaction of light with matter (e.g. photoelectric effect)
- studies of light emission (e. g. atomic spectra and blackbody radiation)

LIGHT

light = wave (Hooke, Huygens 17th century)

Newton, 1666: white light = mixture of several different colors, complex wave (polichromatic)





red (or blue or green etc) light = simple wave (monochromatic)

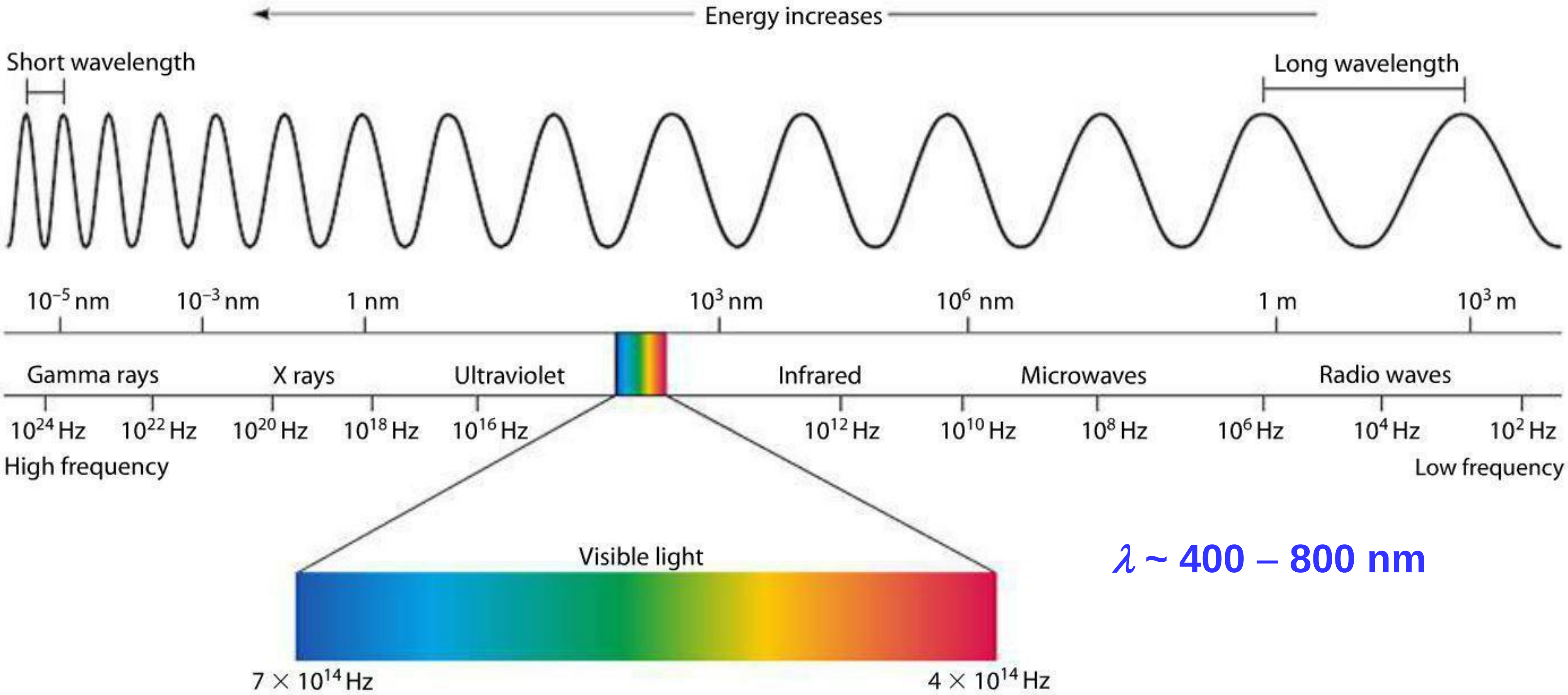
LIGHT WAVES

wavelength, λ : distance over which the wave's shape repeats

frequency, $\nu = \frac{c}{\lambda}$ (where c: speed of light)

number of waves per unit time

Electromagnetic radiation

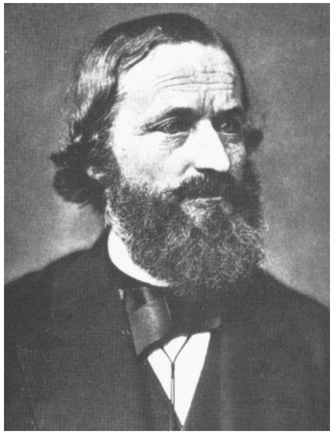


λ ~ 400 – 800 nm

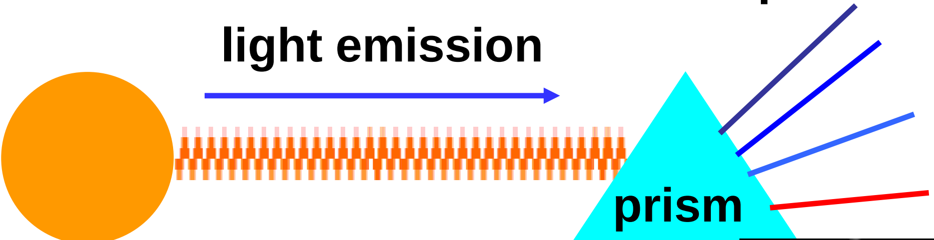
Spectrum analysis = Spectroscopy (1850s)

Robert Wilhelm Bunsen (1811-1899)

- heated elements emit characteristic colors (= spectral lines of given wavelengths)
- linkage between chemical elements and their unique spectral patterns → chemical analysis



Gustav Robert Kirchhoff (1824-1887)



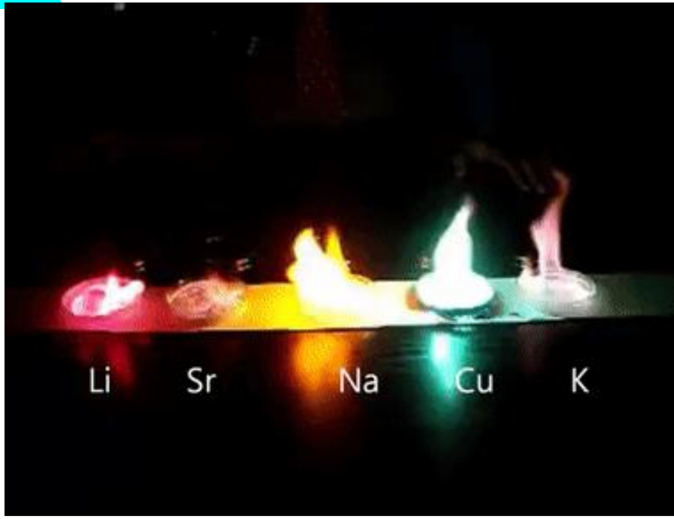
hot sample containing the element

light emission

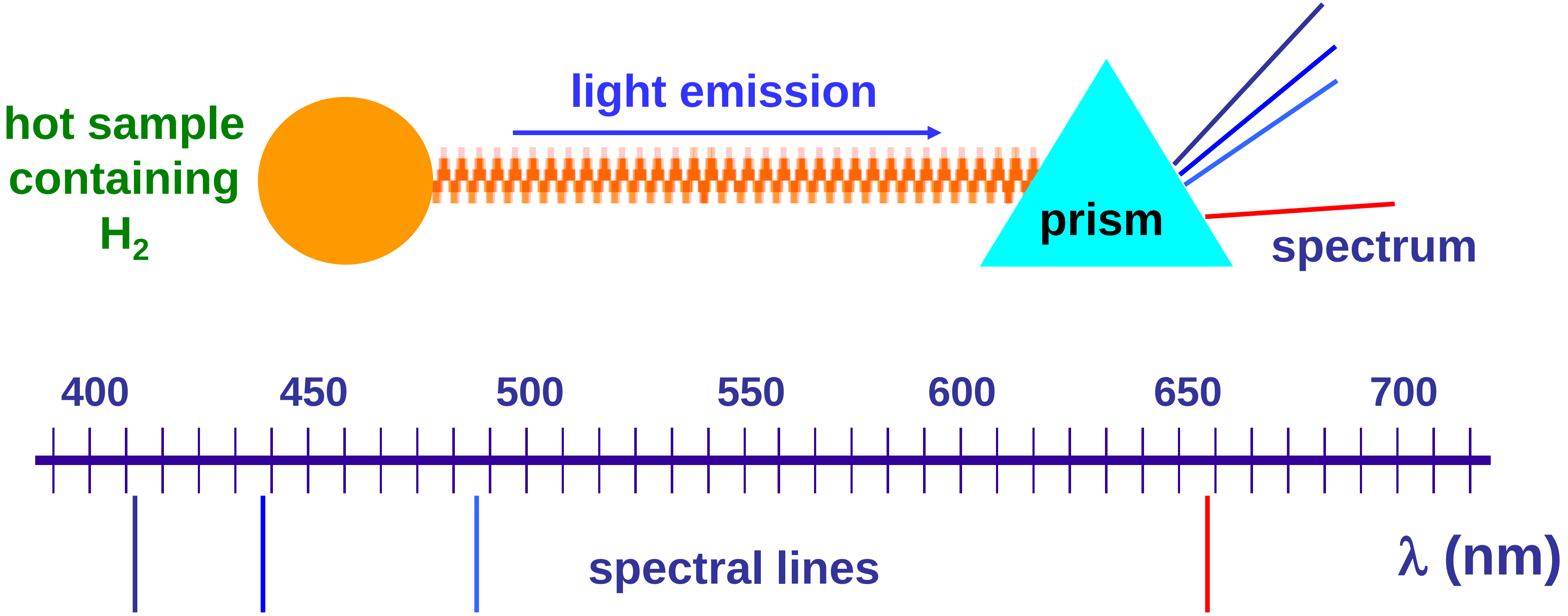
spectrum

prism

- mostly alkali and alkaline earth metals: **FLAME TEST**
- discovery of new elements (Cs, Rb)
- chemical composition of the Sun



ATOMIC SPECTRUM OF HYDROGEN



Balmer-Rydberg equation: empirical expression describing the wavelengths of spectral lines of H

$$\frac{1}{\lambda} = R_H \left[\frac{1}{m^2} - \frac{1}{n^2} \right]$$

n, m: integers $n > m$

$R_H = 1.097 \times 10^{-2}\ \text{nm}^{-1}$ *Rydberg constant*

LINES IN ATOMIC SPECTRA

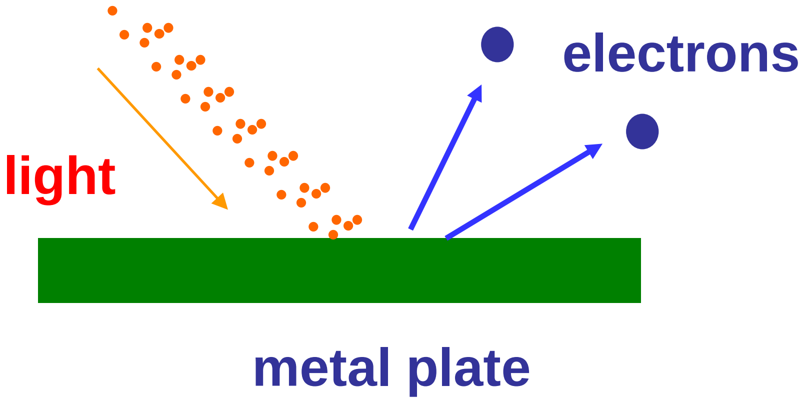
1. Atoms must have discrete (= noncontinuous) possible values for energy

2. Spectral lines arise from transition between these discrete states (higher energy state ⇒ lower energy state: energy is released in the form of light)

3. Lights of different color (i.e. different wavelength) must have different energy

4. Light must contain discrete (= noncontinuous) packets of energy

PHOTOELECTRIC EFFECT

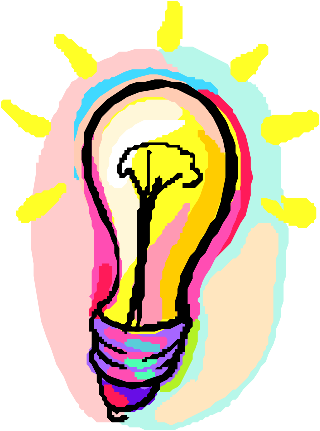


ALBERT EINSTEIN`S INTERPRETATION

Light behaves like a beam of particles (*PHOTONS*)

1921: Nobel prize in Physics

BLACKBODY RADIATION



- hot objects (e.g.: stars, light bulb) emit light
- continuous spectrum (not atomic spectrum!)
- classical physics cannot describe it quantitatively
- Planck assumed that radiation can only be emitted in small packets (quanta)

atomic spectra, blackbody radiation, *photoelectric effect* → light is not just a wave anymore, it can behave as a stream of quanta

AIP

light carries energy

interactions: energy is always transferred in discreet amounts

1 light quantum = 1 PHOTON

energy of a photon $E = h\nu = h\frac{c}{\lambda}$

h: Planck's constant 6.626 × 10^{-34} Js

Max Karl Ernst Ludwig Planck (1858-1947)

1918: Nobel prize in Physics

SUBATOMIC PARTICLES, ATOMIC MODELS

Discovery of subatomic particles:

- Electron, J. J. Thomson, 1897
- Proton, E. Rutherford, 1911
- Neutron, J. Chadwick, 1932

Atomic models

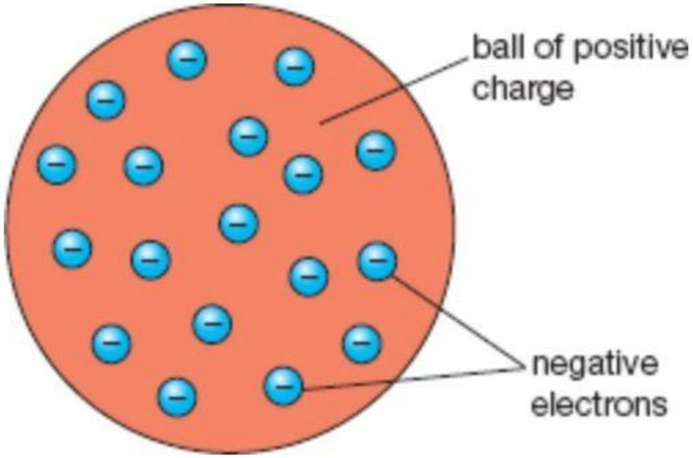
- **J. J. Thompson: pudding model**
- **Rutherford: planetary model**
- Bohr model
- Quantum mechanical model: Heisenberg and Schrödinger

Thomson's atomic model

- **attempted to consolidate the known properties of atoms at the time:**
  1) electrons are negatively-charged particles and
  2) atoms are neutrally-charged.
- **atom: uniform spheres of positively charged matter in which electrons are embedded**
- **each atom has different number of electrons, and they are distributed like plums inside of a pudding**

Joseph John Thomson (1856-1940)

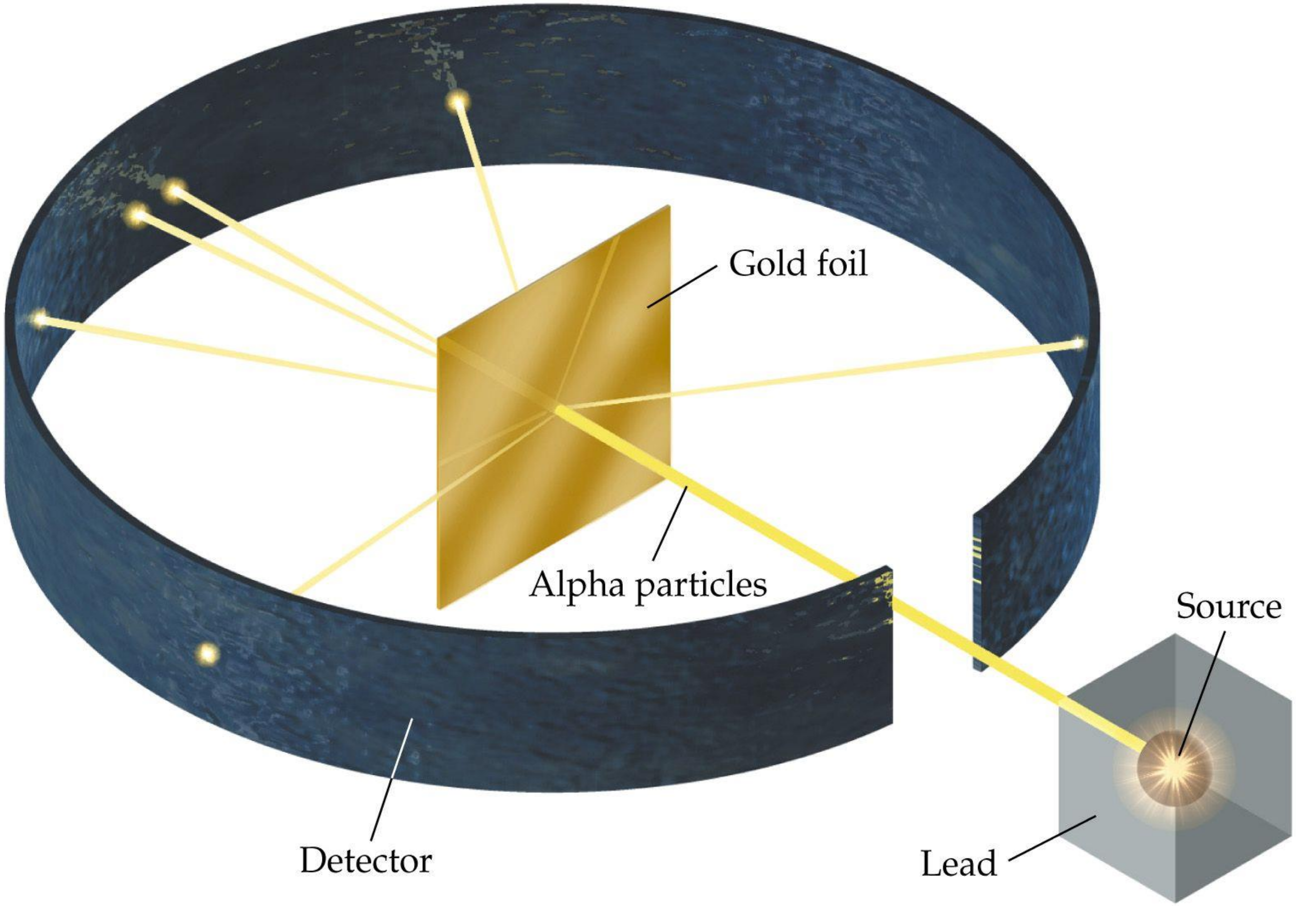
1906: Nobel prize in Physics



Thomson's 'plum-pudding' model of the atom

Gold foil experiment or Scattering experiment (1911)

Discovery of proton and the nucleus (1911)



Hans Geiger and Ernest Marsden

Ernest Rutherford
(1871–1937)

Rutherford’s atomic model

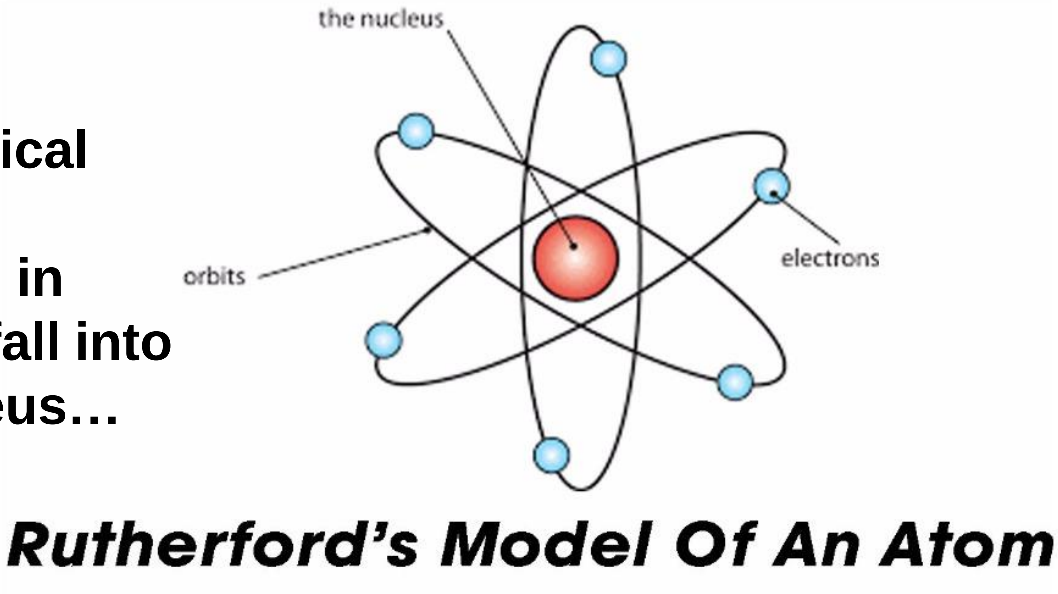
- atomic nucleus is very small, its radius is about 1/100000 of the radius of the atom

  nuclei r ~ 10^{-15} m atoms d ~ 10^{-10} m (1 Å)

- positive charge is concentrated into the nucleus (into a very small volume)

- central nucleus is surrounded by a cloud of orbiting electrons
  → planetary model

- problem: according to classical physics, negatively charged electrons should not remain in orbit, they should instantly fall into the positively charged nucleus…



BOHR MODEL, 1913

1922: Nobel prize in Physics

based on classical mechanics, added further postulates to interpret quantized energy changes

Niels Henrik David Bohr (1885-1962)

Postulate 1: Stationary states in the atom:

Electrons cannot occupy arbitrary orbits, only ones in which their energies remain unchanged in time

Postulate 2: Frequency condition:

Light absorption and emission of photons results from a transition between two stationary states;

$E_2 - E_1 = h\nu$

$$\frac{1}{\lambda} = \frac{E_0}{hc}\left[\frac{1}{m^2} - \frac{1}{n^2}\right] = \boxed{\frac{e^4 m_e}{8\varepsilon_0^2 h^3 c}}\left[\frac{1}{m^2} - \frac{1}{n^2}\right]$$

$$\frac{1}{\lambda} = \boxed{R_H}\left[\frac{1}{m^2} - \frac{1}{n^2}\right]$$

WAVE-PARTICLE DUALITY, 1924

If light can behave as a stream of particles (like matter), can ordinary matter behave like a wave?

$E = mc^2$ and $E = \frac{hc}{\lambda}$ $\rightarrow mc = \frac{h}{\lambda}$

$$\lambda = \frac{h}{mv} = \frac{h}{p}$$

Wavelength of a macroscopic object is unmeasurably small

e.g. car: v = 100 km/h (360 m/s); m = 1 t

$\lambda_{car} = 1.8 \times 10^{-39}$ m

Wavelength of a microscopic object can be of huge importance

e.g. electron in H atom: $v_e = 2.2 \times 10^6$ ms^{-1}; $m_e = 9.1 \times 10^{-31}$ kg

$\lambda_e = 3.3 \times 10^{-10}$ m

Louis-Victor de Broglie (1892–1987)

1929: Nobel prize in Physics

UNCERTAINTY PRINCIPLE, 1927

- **fundamental limit to the precision with which certain pairs of physical properties of a particle (known as complementary variables) can be known**
- **it arises in quantum mechanics simply due to the wave-particle nature of all quantum objects**

Werner Karl Heisenberg (1901–1976)

Complementary variables include:

- **Position and momentum**
- **Energy and duration**

i.e. the momentum and the position of an object cannot be determined precisely simultaneously

Uncertainty principle:

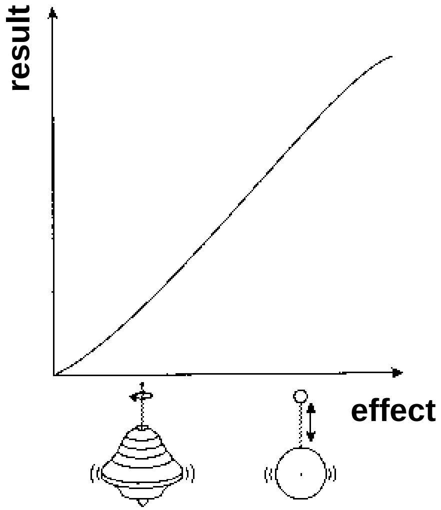
$$\sigma_x \sigma_{mv} \geq \frac{h}{4\pi}$$

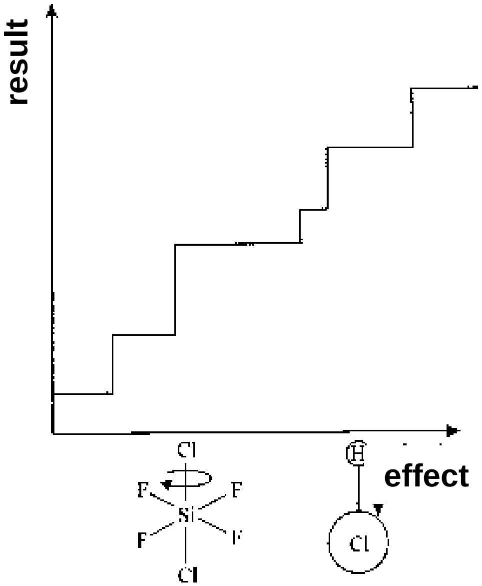
σ_x: **uncertainty of position**

σ_{mv}: **uncertainty of momentum ($p = mv$, momentum)**

QUANTUM MECHANICS – NOVEL PHYSICS

Classical physics: macroscopic world, quantities (functions) are continouos, particle-wave duality and uncertainty principle are valid (but usually irrelevant)





Quantum physics: microscopic world (atoms, subatomic particles), but it can be projected to macroscopic scale, quantities (functions) are quantized, particle-wave duality and uncertainty principle are valid (and are of huge importance)

QUANTUM MECHANICAL MODEL OF THE ATOM

1926, Erwin Schrödinger: WAVE MECHANICS

1933: Nobel prize in Physics

1927, Werner Heisenberg: MATRIX MECHANICS

1932: Nobel prize in Physics

Erwin Rudolf Josef Alexander Schrödinger (1887-1961)

The two different physical approaches were shown to be mathematically EQUIVALENT.

Schrödinger equation

- **the analogue of Newton's law (*F = ma*) for a quantum system**
- **classical mechanics: past and future of a system (when and where was/will be total eclipse of the sun, when will Halley's comet be next visible?)**
- **quantum mechanics: can only predict, nothing about the past**

SCHRÖDINGER EQUATION

- state equation ≡ complicated partial differential equation for Ψ state function
- exact solution (in a closed, analytical form) can only be given for the simplest cases
- e.g.: a free particle with mass of *m*, not subject to a potential
  - its momentum (so its rate) and its energy are known exactly
  - over all space, the probability of finding the particle is identical → the particle can be anywhere and it is unknown how much time it spends in a given position

⇒ the uncertainty principle works!!

- When applying the Schrödinger equation for an atom
  - electron is considered as a wave
  - $\Psi(x,y,z,t)$: wave function
  - Ψ^2 ~ probability of finding an electron in a given volume of space

QUANTUM MECHANICAL MODEL OF THE ATOM

Schrödinger equation can be solved exactly for H atom (and hydrogen-like, one-electron systems)

Solution: quantized energy levels without postulates

Quantized energy levels: quantum numbers, separate orbitals in the atom for the electron

orbit ≡ classical mechanics (the path of an object)

orbital ≡ quantum mechanics

Result: 3 different quantum numbers

principal quantum number (n): defines primarily the size and the energy level of the orbital in the H atom

$n = 1, 2, 3, 4, 5, ...$ (integer)

orbitals with the same n: SHELL

$n = 1 \rightarrow K$ shell $n = 2 \rightarrow L$ shell *(M, N,.......)*

angular-momentum quantum number (l): defines primarily the three-dimensional shape of the orbital in the H atom

$l = 0, 1, 2, \ldots, n-1$ (integer)

orbitals with the same n and l: SUBSHELL

$l = 0 \Rightarrow s$ subshell $l = 1 \Rightarrow p$ $l = 2 \Rightarrow d$ $l = 3 \Rightarrow f$

magnetic quantum number (m_l): defines primarily the orientation of the orbital in the H atom relative to an outside magnetic field

$m_l = -l, -l+1, \ldots, -1, 0, 1, \ldots, l-1, l$ (integer)

total: $2l + 1$ possible values

n, l, and m_l identify a given orbital

The number of orbitals within a subshell:

s: $l = 0$, $m_l = 0$ (1 s orbital)
p: $l = 1$ $m_l = -1, 0, +1$ (3 p orbitals)
d: $l = 2$ $m_l = -2, -1, 0, +1$ +2 (5 d orbitals)
f: $l = 3$ $m_l = -3 \ldots 0 \ldots +3$ (7 f orbitals)

ORBITALS

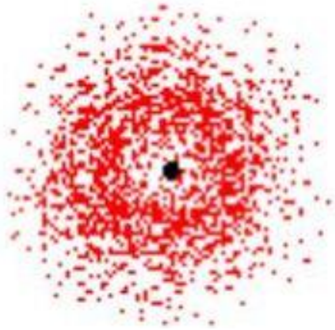
n	l	m_l	subshell	orbital
1	0	0	1*s*	1*s*
2	0	0	2*s*	2*s*
2	1	-1	2*p*	$2p_x$, $2p_y$, $2p_z$
2	1	0	2*p*	
2	1	+1	2*p*	
3	0	0	3*s*	3*s*
3	1	-1	3*p*	$3p_x$, $3p_y$, $3p_z$
3	1	0	3*p*	
3	1	1	3*p*	

ORBITALS

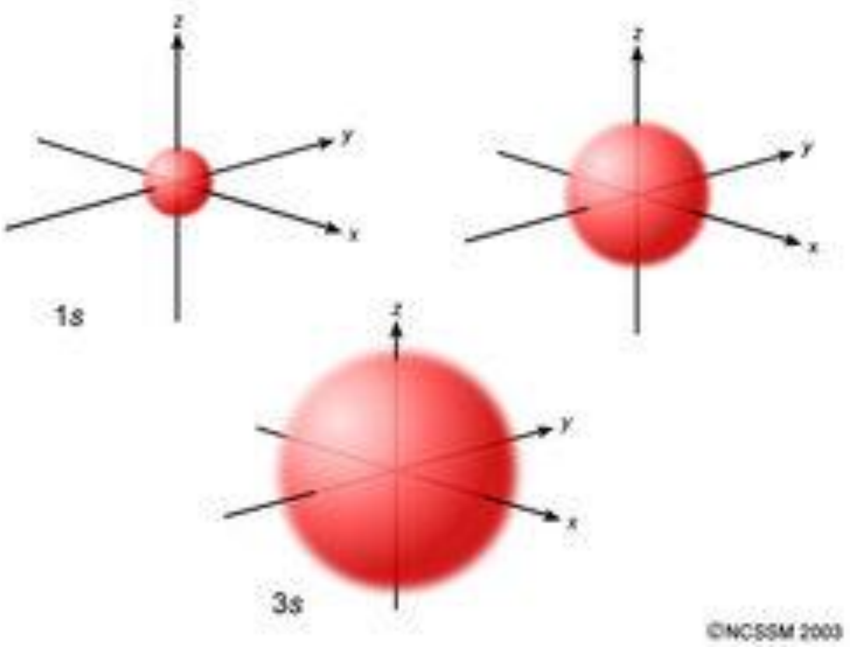
n	l	m_l	subshell	orbital
3	2	-2	$3d$	$3d_{xy}$, $3d_{yz}$, $3d_{xz}$, $3d_{x^2-y^2}$, $3d_{z^2}$
3	2	-1	$3d$	
3	2	0	$3d$	
3	2	+1	$3d$	
3	2	+2	$3d$	
4	0	0	$4s$	$4s$
4	1	-1 – +1	$4p$	$4p_x$, $4p_y$, $4p_z$
4	2	-2 – +2	$4d$	$4d_{xy}$, $4d_{yz}$…
4	3	-3 – +3	$4f$	$4f_{xyz}$, $4f_{xz^2}$, $4f_{yz^2}$…

SPATIAL REPRESENTATION OF ORBITALS

- Only probabilites known for electron movement!!
- Atomic orbitals are usually represented by surfaces, which contain the electron with high probability (90 %) and reflect the symmetry of the orbital.



s orbitals

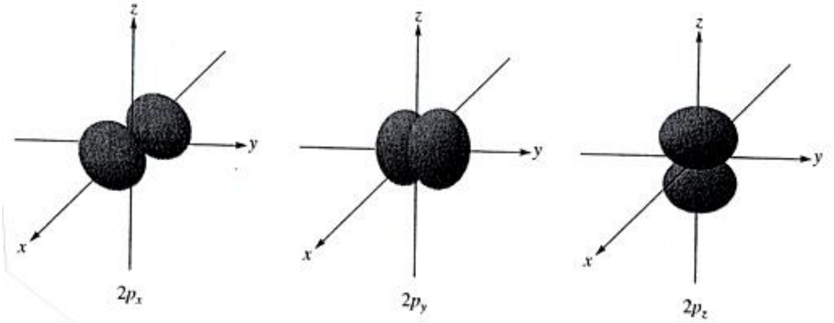


Spherical symmetry:

- probability of finding (Ψ^2) an e^- depends only on distance from the nucleus, not on the direction

p orbitals

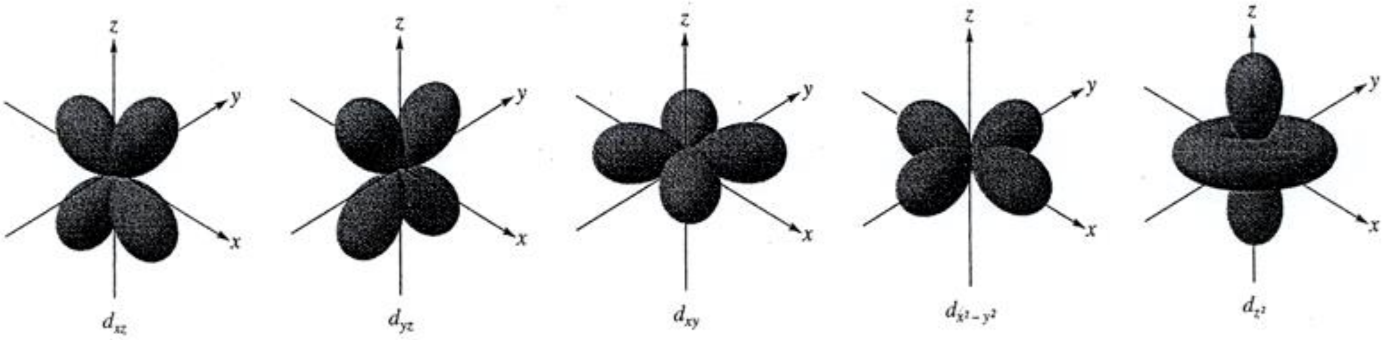
- Ψ^2 depends on distance and direction
- 3 different, perpendicular orbitals: cylindrical symmetry (dumbbell shape) around one of the x, y, z axes



- two lobes with identical electron density separated by a planar node (surface of zero probability; $\Psi^2 = 0$)

d orbitals

- 5 different orbitals
- different directions, and two subtypes: 4 have a shape of a cloverleaf, 1 is like a dumbbell (like a p orbital) + donut-shaped region
- each has 2 nodes (when $n = 1$)

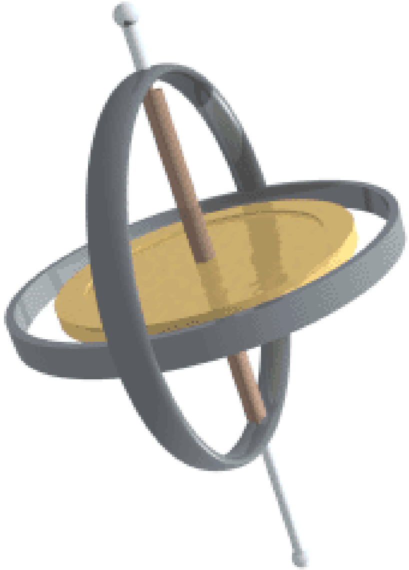


SPIN OF ELECTRONS

- **spin is a quantum mechanical property of the electron**
- **electrons behave as if they were spinning around an axis (like Earth does), the spinning gives rise to a tiny magnetic field and a spin quantum number (m_s):**
- **its value is independent of the other 3 quantum numbers**

possible values: $m_s = -½$ ↓

$m_s = +½$ ↑



MULTIELECTRON ATOMS

Schrödinger equation: cannot be solved exactly for multielectron atoms because the electron-electron repulsion cannot be handled

approximate solutions: orbitals are very similar to the orbitals of the hydrogen atom, but the sizes and energies are different

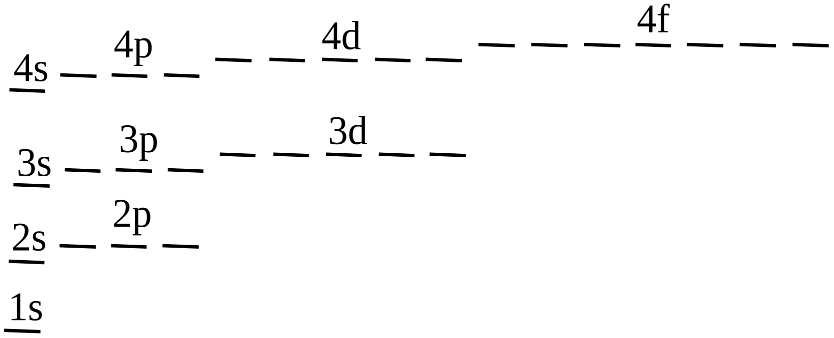
H-atom: energy depends only on *n* → subshells within the same shell have identical energy (degeneration)

***E*(2s) = *E*(2p)**

***E*(3s) = *E*(3p) = *E*(3d)**

Multielectron atoms (Z > 1):

- **energy of orbitals depend on both *n* and *l* → energies of subshells within the same shell are different**



Order of orbital energy levels:

1*s* < 2*s* < 2*p* < 3*s* < 3*p* < 4s < 3d < 4*p* < 5*s* < 4d < 5p < < 6s < 5d < 4f < 6p < 7s

ELECTRON CONFIGURATIONS OF MULTIELECTRON ATOMS

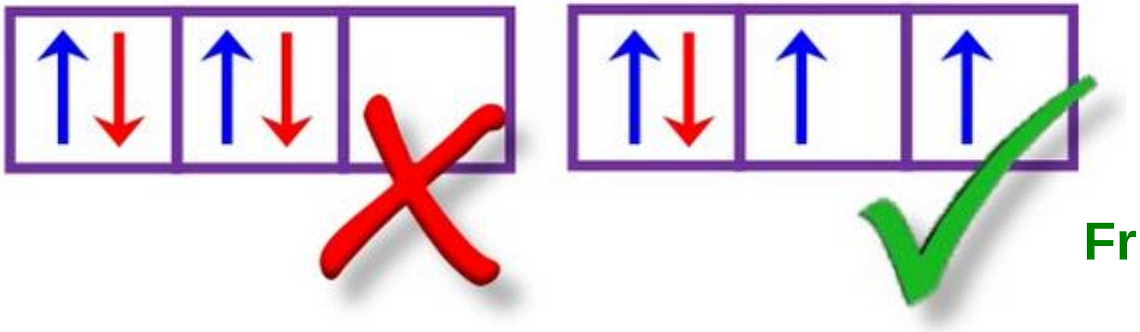
Wolfgang Ernst Pauli (1900 –1958)

Aufbau (German: "building up") principles:

1. Lower-energy orbitals fill before higher-energy orbitals

2. **Pauli exclusion principle:** An orbital can only hold two electrons with opposite spins (= No two electrons in an atom can have the same four quantum numbers)

3. **Hund's rule:** If two or more orbitals with the same energy are available, one electron goes in each until all are half-full. The electrons in the half-filled orbitals all have the same spin quantum numbers.



Friedrich Hermann Hund (1896- 1997)

H: $1s^1$

He: $1s^2$

Li: $1s^2\ 2s^1$

Be: $1s^2\ 2s^2$

B: $1s^2\ 2s^2\ 2p_x^1$

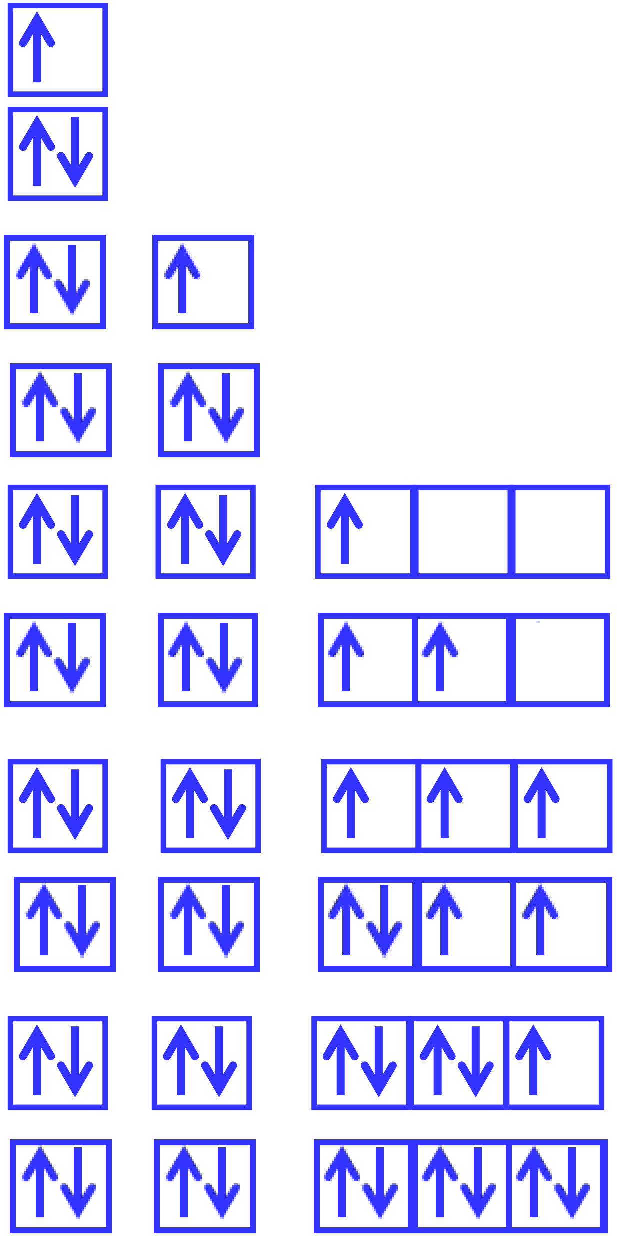
C: $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1$

N: $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$

O: $1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$

F: $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^1$

Ne: $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2$



Na: $1s^2\ 2s^2\ 2p^6\ 3s^1 \equiv [\mathrm{Ne}]\ 3s^1$

Al: $[\mathrm{Ne}]\ 3s^2\ 3p_x{}^1$

P: $[\mathrm{Ne}]\ 3s^2\ 3p_x{}^1\ 3p_y{}^1\ 3p_z{}^1$

Ar: $[\mathrm{Ne}]\ 3s^2\ 3p^6$

K: $[\mathrm{Ar}]\ 4s^1$

Ca: $[\mathrm{Ar}]\ 4s^2$

valence shell

Sc: $[\mathrm{Ar}]\ 4s^2\ 3d^1$

Zn: $[\mathrm{Ar}]\ 4s^2\ 3d^{10}$

Ga: $[\mathrm{Ar}]\ 4s^2\ 3d^{10}\ 4p_x{}^1$

As: $[\mathrm{Ar}]\ 4s^2\ 3d^{10}\ 4p_x{}^1\ 4\mathrm{p}_y{}^1\ 4p_z{}^1$

VALENCE ELECTRONS

- Electrons of the outer shell and the inner, unclosed subshells

  (core electrons: „everything else”; not valence electrons)

- Number of valence electrons primarily determines an element’s chemical properties

- Main group elements (s and p block): group number = number of valence electrons

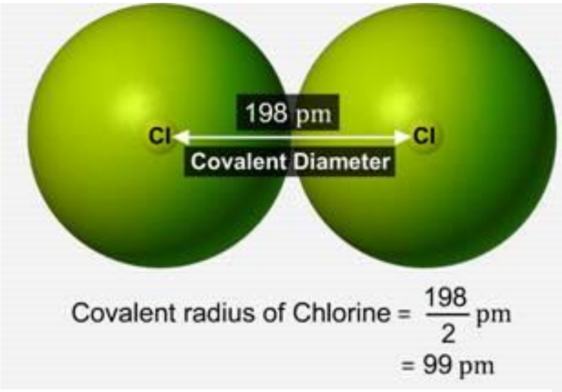
alkali metals: ns^{1}	alkaline earth metals: ns^{2}
boron group: ns^{2} np^{1}	carbon group: ns^{2} np^{2}
nitrogen group: ns^{2} np^{3}	chalcogens: ns^{2} np^{4}
halogens: ns^{2} np^{5}	noble gases: ns^{2} np^{6}

- Transition metals: often group number = number of valence electrons, but there are exceptions!!

ATOMIC RADIUS

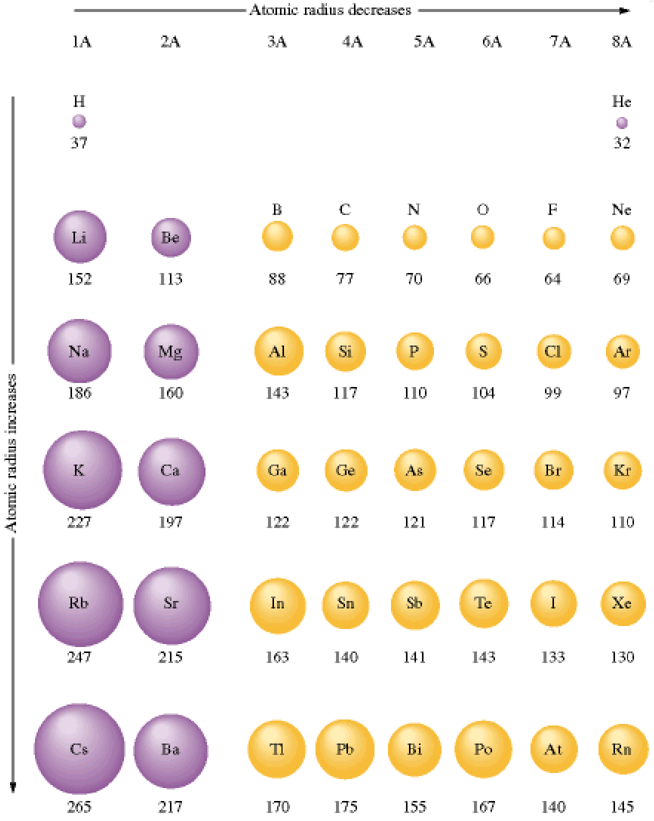
challenges with the definition:

- **bound or not bound?**
- **atoms have no definite border (90 % of electron density denotes the „end” of atomic orbitals)**
- **based on the distance between nuclei in bonds (*covalent radius*)**



decrease in atomic radius

increase in atomic radius, increase in *n*



IONIC RADIUS

Atom > cation $r(M) > r(M^{x+})$ reason: same p^+, less e^-

Atom > anion $r(M) < r(M^{x-})$ reason: same p^+, more e^-

isoelectronic series: species with the same number of electrons and the same electron configuration

N^{3-}, O^{2-}, F^-, Ne, Na^+, Mg^{2+}, Al^{3+}

←

$1s^2\,2s^2\,2p_x^2\,2p_y^2\,2p_z^2$

increase in size; reason: same e^-, increasing p^+ from left to right

$r(Fe) > r(Fe^{2+}) > r(Fe^{3+})$

126 pm 92 pm 79 pm

IONIZATION ENERGY

Definition: removing an electron from a gaseous atom

$$E\,(g) \rightarrow E^{+}(g) + e^{-} \qquad E_i \text{ (kJ/mol)}$$

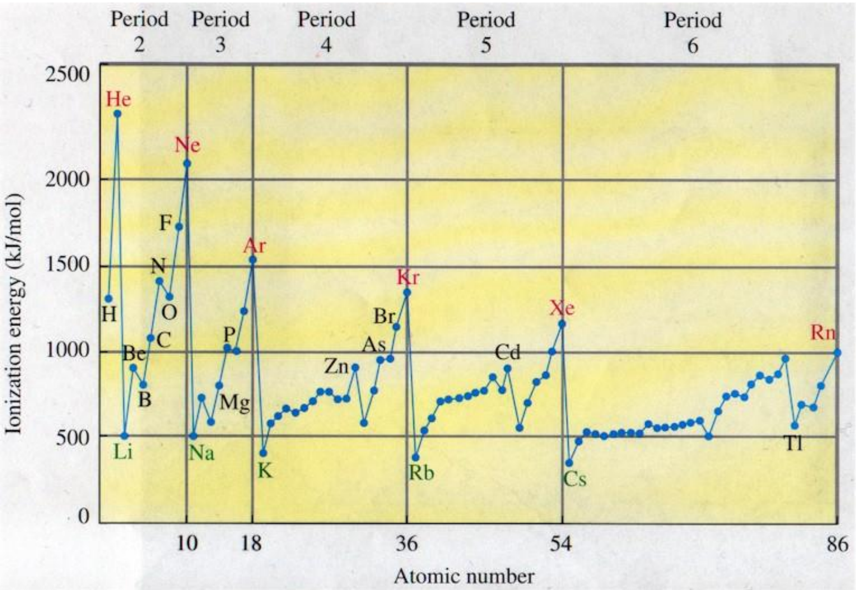
first ionization energy

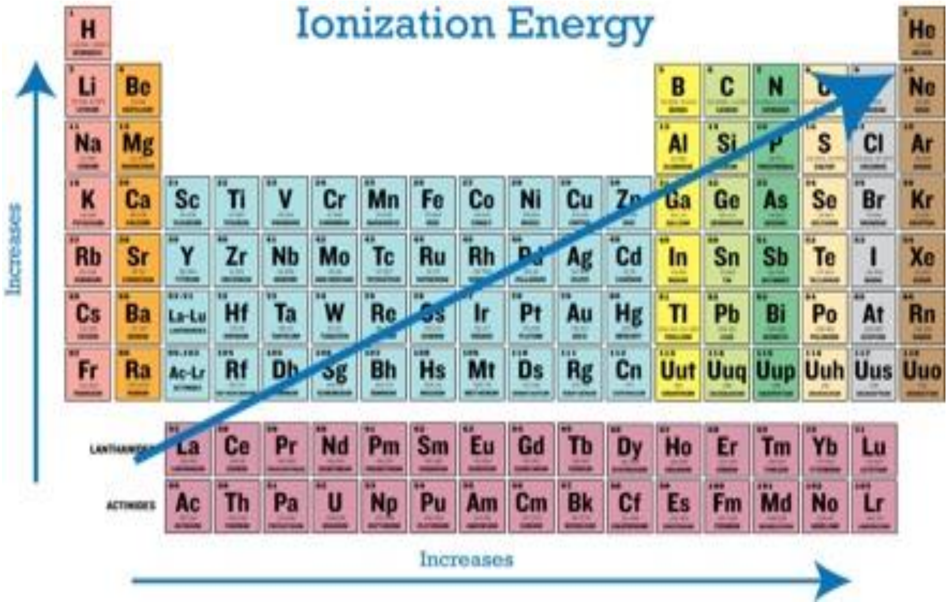
The energy needed to remove an electron from a neutral atom in the gaseous state, always positive

The removal of further electrons (2nd, 3rd etc E_i), require more energy

lowest: 376 kJ/mol (Cs)

highest: 2372 kJ/mol (He)





ELECTRON AFFINITY

Definition: adding an electron to a gaseous atom

$$E\ (g) + e^{-} \rightarrow E^{-}\ (g) \qquad E_{ea}\ \text{(kJ/mol)}$$

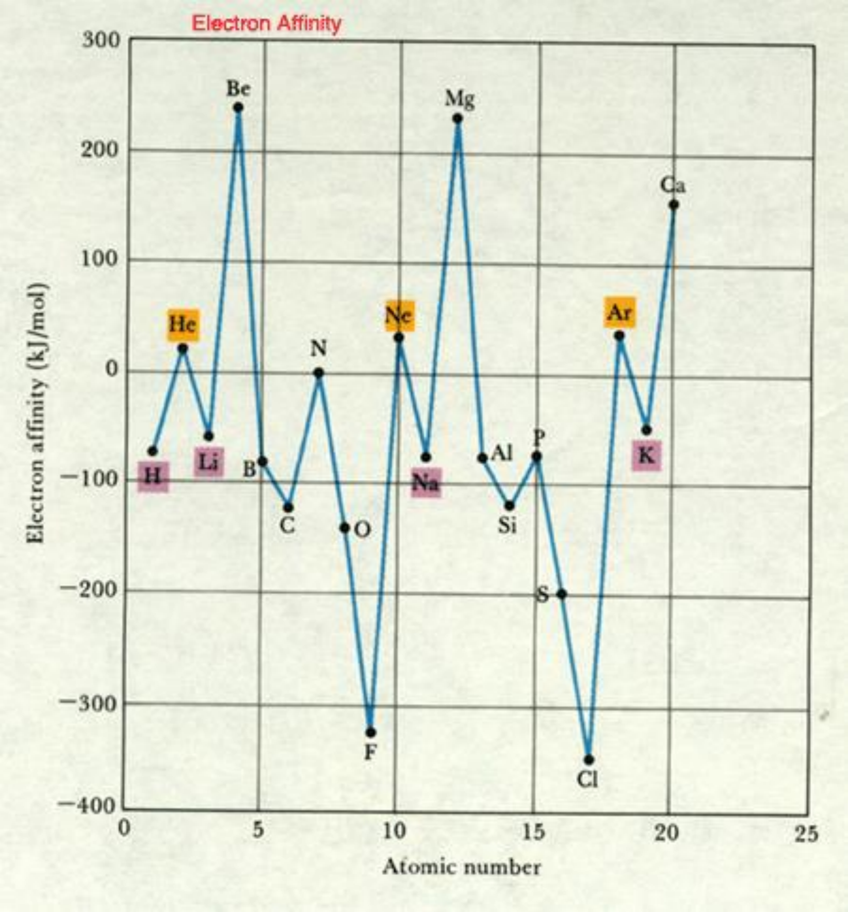
both energy release or absorption is possible

release: ***negative*** E_{ea}

e.g. H –73 kJ/mol

absorption: ***positive*** E_{ea}

e.g. He ~ +35 kJ/mol



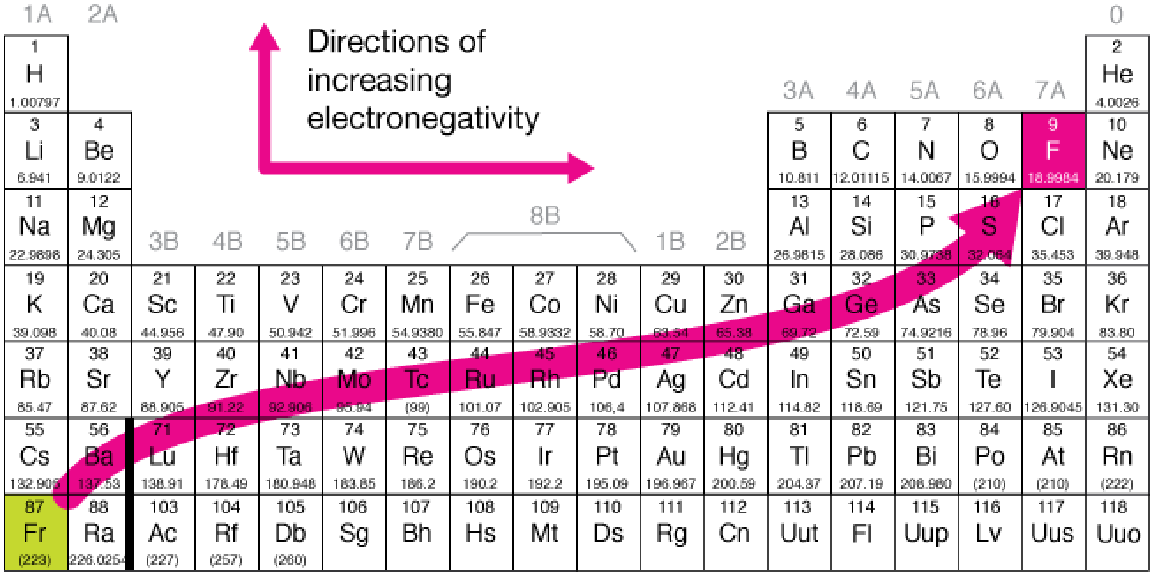
ELECTRONEGATIVITY

a dimensionless number showing the ability of an atom in a molecule to attract the shared electrons in a covalent bond

F: 4.0   O: 3.5   Cl: 3.0   H: 2.1   Al: 1.5   K: 0.8

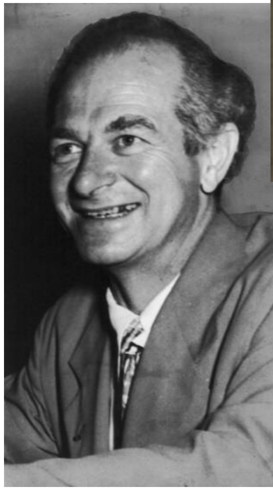
Two main calculation methods:

- **Pauling: based on the bond enthalpies of diatomic molecules**
- **Mulliken: based on the sum of first ionization energies and electron affinities**



1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	0
1 H 1.00797																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.01115	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.179
11 Na 22.9898	12 Mg 24.305											13 Al 26.9815	14 Si 28.086	15 P 30.9738	16 S 32.064	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.08	21 Sc 44.956	22 Ti 47.90	23 V 50.942	24 Cr 51.996	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.70	29 Cu 63.54	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.905	40 Zr 91.22	41 Nb 92.906	42 Mo 95.94	43 Tc (99)	44 Ru 101.07	45 Rh 102.905	46 Pd 106,4	47 Ag 107.868	48 Cd 112.41	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te 127.60	53 I 126.9045	54 Xe 131.30
55 Cs 132.905	56 Ba 137.53	71 Lu 138.91	72 Hf 178.49	73 Ta 180.948	74 W 183.85	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.09	79 Au 196.967	80 Hg 200.59	81 Tl 204.37	82 Pb 207.19	83 Bi 208.980	84 Po (210)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.0254	103 Ac (227)	104 Rf (257)	105 Db (260)	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo

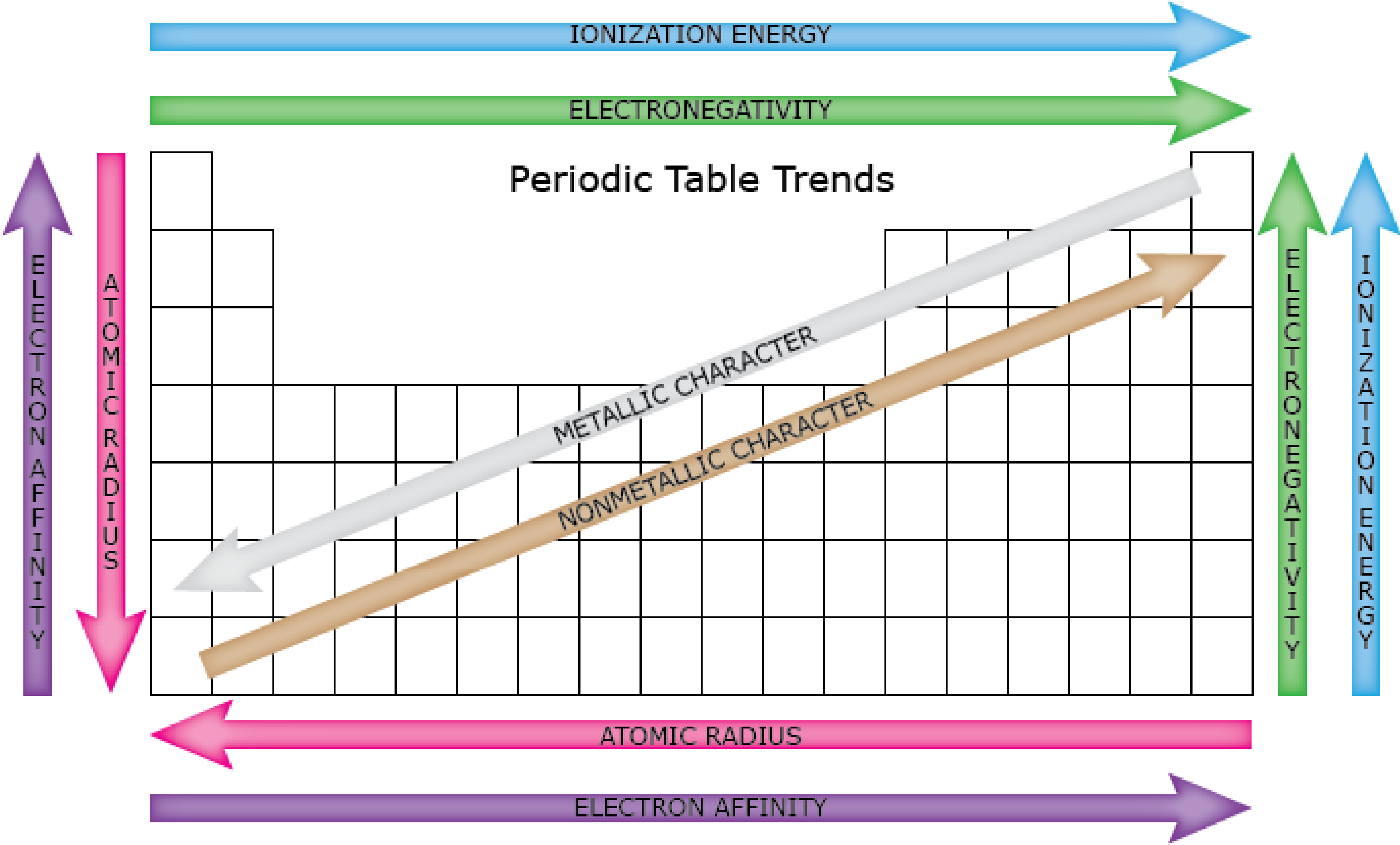
Robert Sanderson Mulliken (1896–1986)



Linus Carl Pauling (1901- 1994)

1954: Nobel prize in Chemistry

Periodic Table Trends



IONIZATION ENERGY

ELECTRONEGATIVITY

ELECTRON AFFINITY

ATOMIC RADIUS

METALLIC CHARACTER

NONMETALLIC CHARACTER

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