Welcome to Chemistry Standards 101!

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Chemistry

Study of matter and the changes it undergoes

Why study chemistry?

- Chemistry is all around us
- Chemistry is the central science- connects other areas of science
- Chemistry helps you understand the world around you
- Makes you a more informed consumer
- Chemistry is central to cooking
- Teaches you how to solve problems, utilize the scientific method
- Helps you understand current events
- Chemistry is fun!

Chemistry Standards

- CHEM.A.2: Atomic Structure and the Periodic Table
  - CHEM.A.2.2: Describe the behavior of electrons in atoms.

- This section includes predicting ground state electron configuration, predicting characteristics of an atom or ion based on its location on the periodic table, explaining relationship between the electron configuration and the atomic structure of an atom or ion, relating quantized energy levels to atomic emission spectra

Atoms Are Ancient and Empty

Atoms are

- ancient
  - —origin of most atoms goes back to birth of universe
- mostly empty space

Elements heavier than hydrogen and much of the helium were produced in the interiors of stars.

The Elements

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- Atom: The fundamental unit of an element.

The term “element” is used when referring to macroscopic quantities.

The term “atom” is used when discussing the submicroscopic.

The Elements

Atoms:

- make up all matter around us
- to date, 116 distinct kinds of atoms—

  90 found in nature, remainder synthesized

Element

any material consisting of only one type of atom

Protons and Neutrons

Protons:

- carry a positive charge—same quantity of charge as electrons
- are about 1800 times as massive as an electron
- have the same number of protons in the nucleus as electrons surrounding the nucleus of an electrically neutral atom
- the atomic number is the number of protons in each element listed in the periodic table.

Protons and Neutrons

Neutrons:

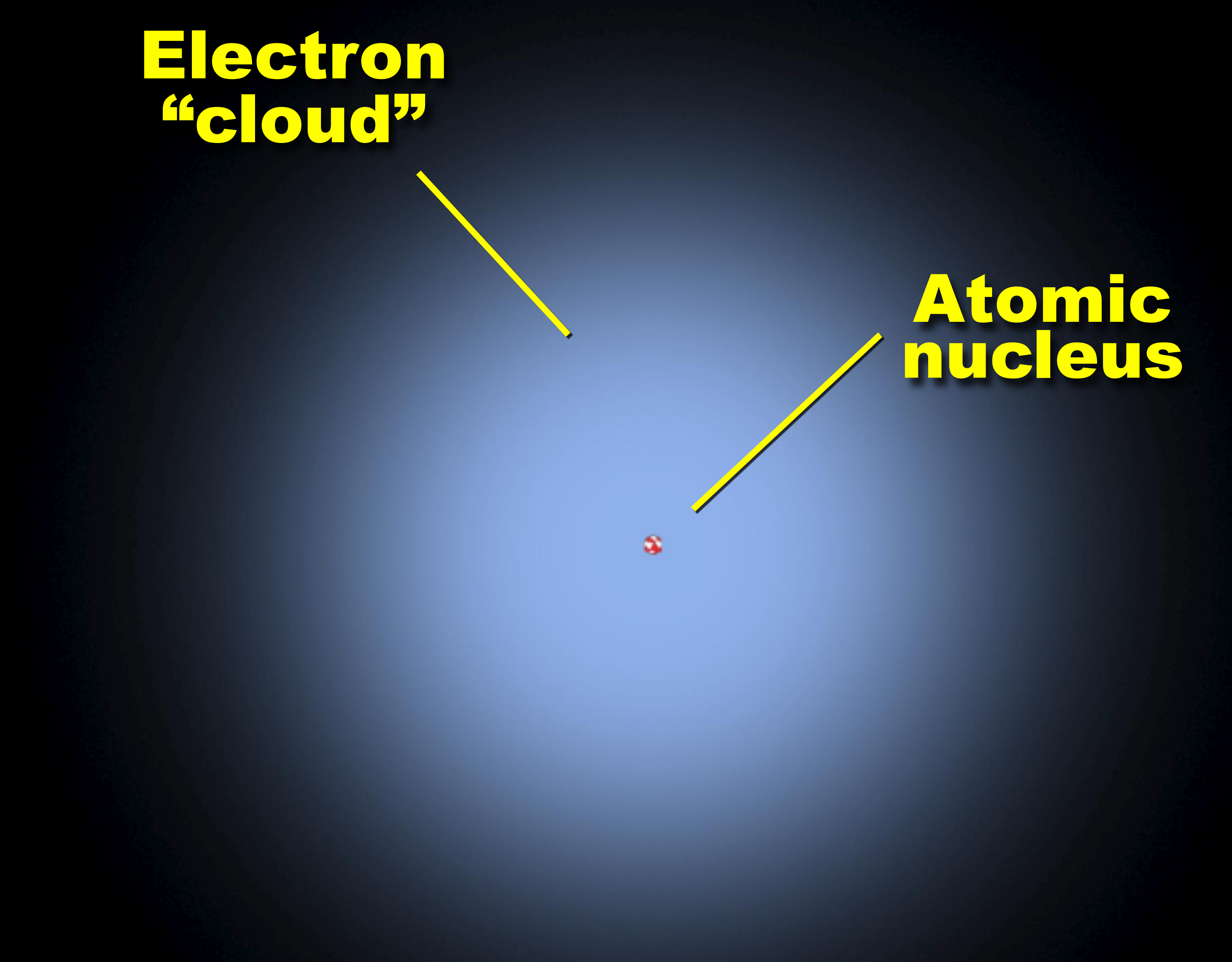
- accompany protons in the nucleus
- have about the same mass as protons but no charge, so are electrically neutral

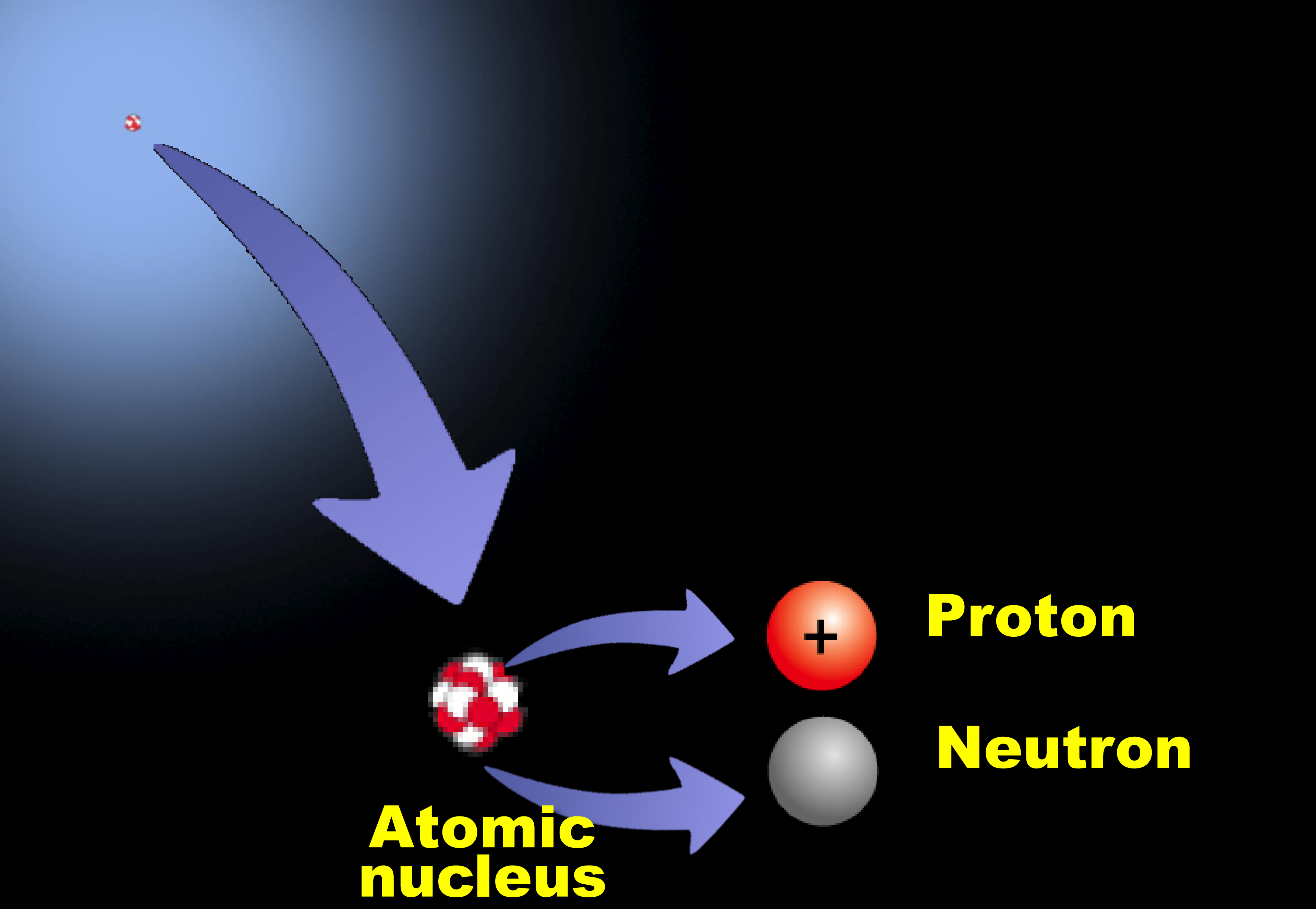
Both protons and neutrons are *nucleons* which means they reside within the nucleus of the atom.

Electrons

Electrons:

- are identical
- repel electrons of neighboring atoms
- have electrical repulsion that prevents atomic closeness
- located outside the nucleus
- in a neutral atom, the number of electrons is equal to the number of protons (atomic number)


Electron
“cloud”
Atomic
nucleus


+
Proton
Neutron
Atomic nucleus

The Periodic Table

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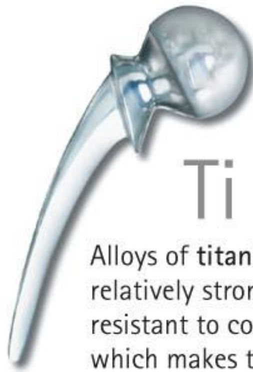
- It is NOT something to be memorized.
- Instead, we learn how to READ the Periodic Table.
- A chemist uses the Periodic Table much like a writer uses a dictionary. NEITHER need be memorized!

Ag

If this **silver** mug were filled with boiling water, the handle would quickly become too hot to handle because silver is one of the best conductors of heat.

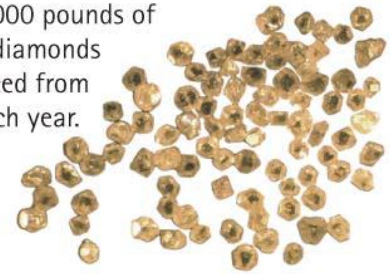


Ti

Alloys of **titanium** are relatively strong and resistant to corrosion, which makes them useful for hip implants.

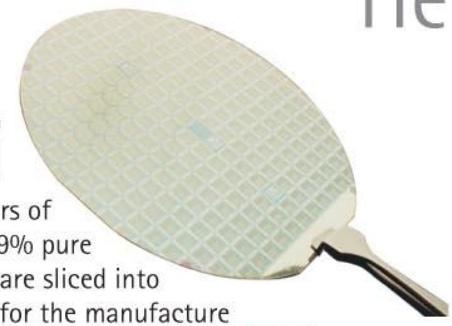
About 50,000 pounds of synthetic diamonds are produced from **carbon** each year.

C



Helium is formed underground as a by product of radioactive decay.

He



Si

Cylinders of 99.9999% pure **silicon** are sliced into wafers for the manufacture of integrated circuits.

Zn

Zinc has a low melting point and is commonly used in making coins.

Hg

Mercury freezes at −40°C and is a liquid at room temperature.

Br

Bromine is a dark orange liquid that readily vaporizes at room temperature.

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 Uun	111 Uuu	112 Uub						

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
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

Metal Metalloid Nonmetal

The Periodic Table

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- Each vertical column is called a “group.”

The Periodic Table

- The elements are highly organized within the Periodic Table.
- Each vertical column is called a “group.”
- Each horizontal row is called a “period.”

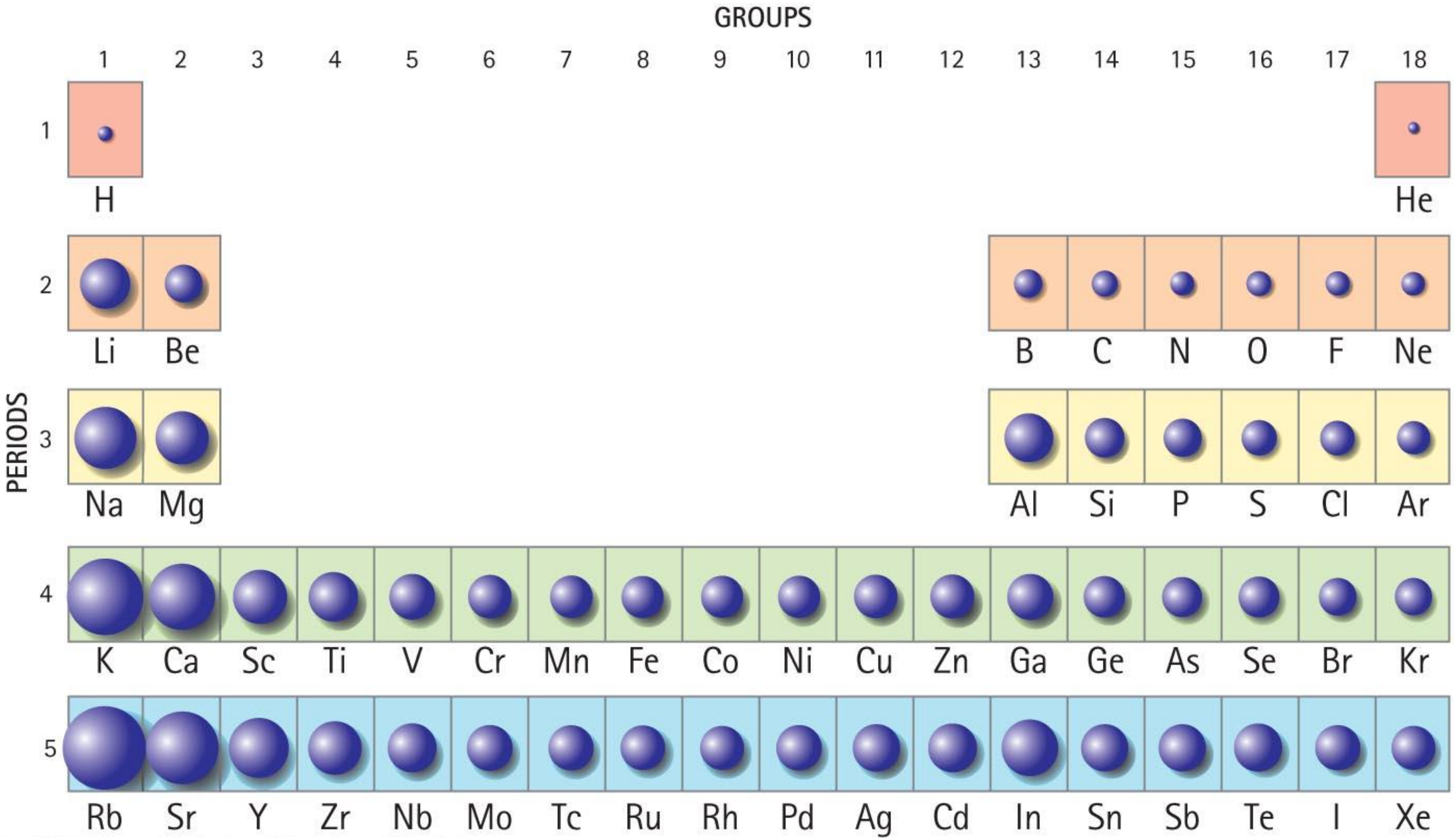
The Periodic Table

GROUPS

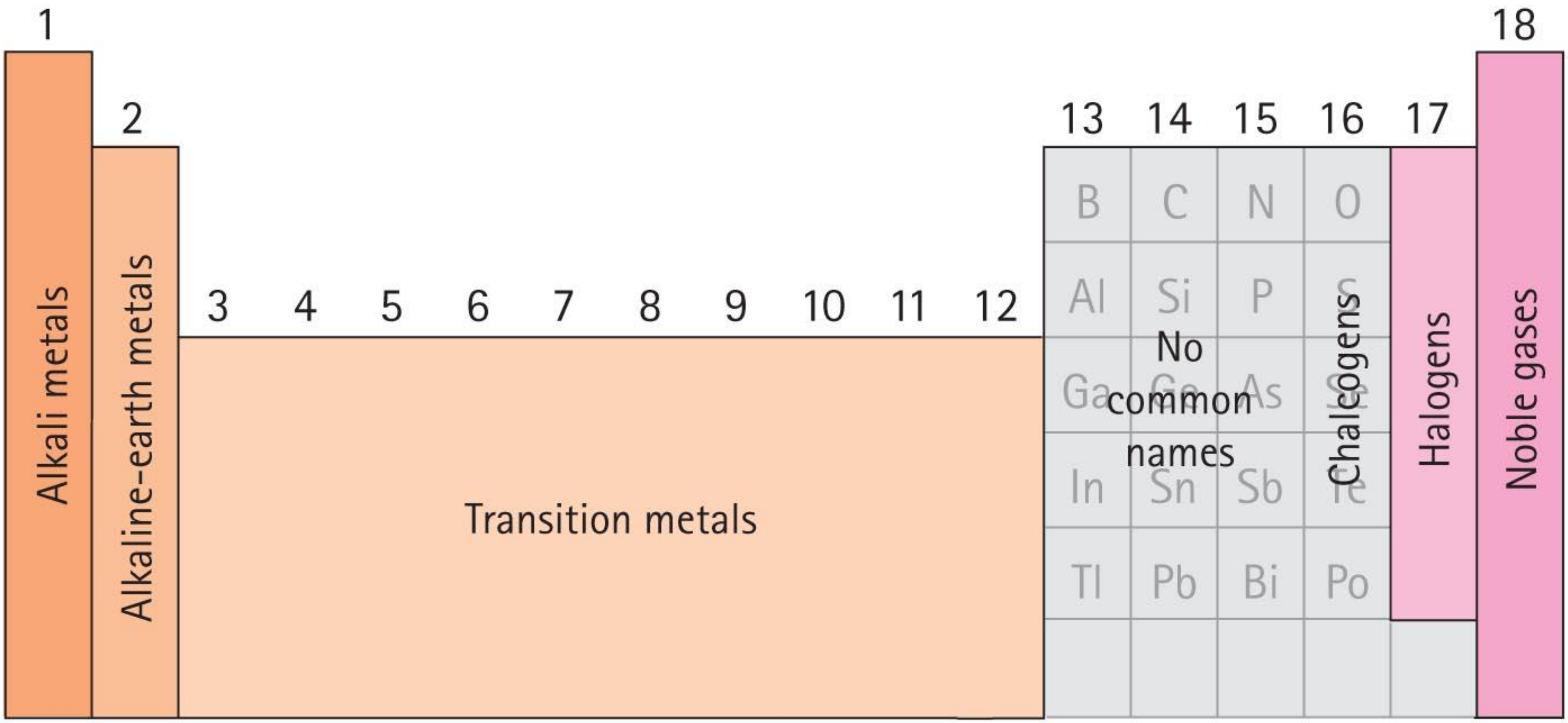
PERIODS	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	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
6	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
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Uun	111 Uuu	112 Uub						

6th-period subset	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
7th-period subset	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

The Periodic Table



The Periodic Table



The Periodic Table

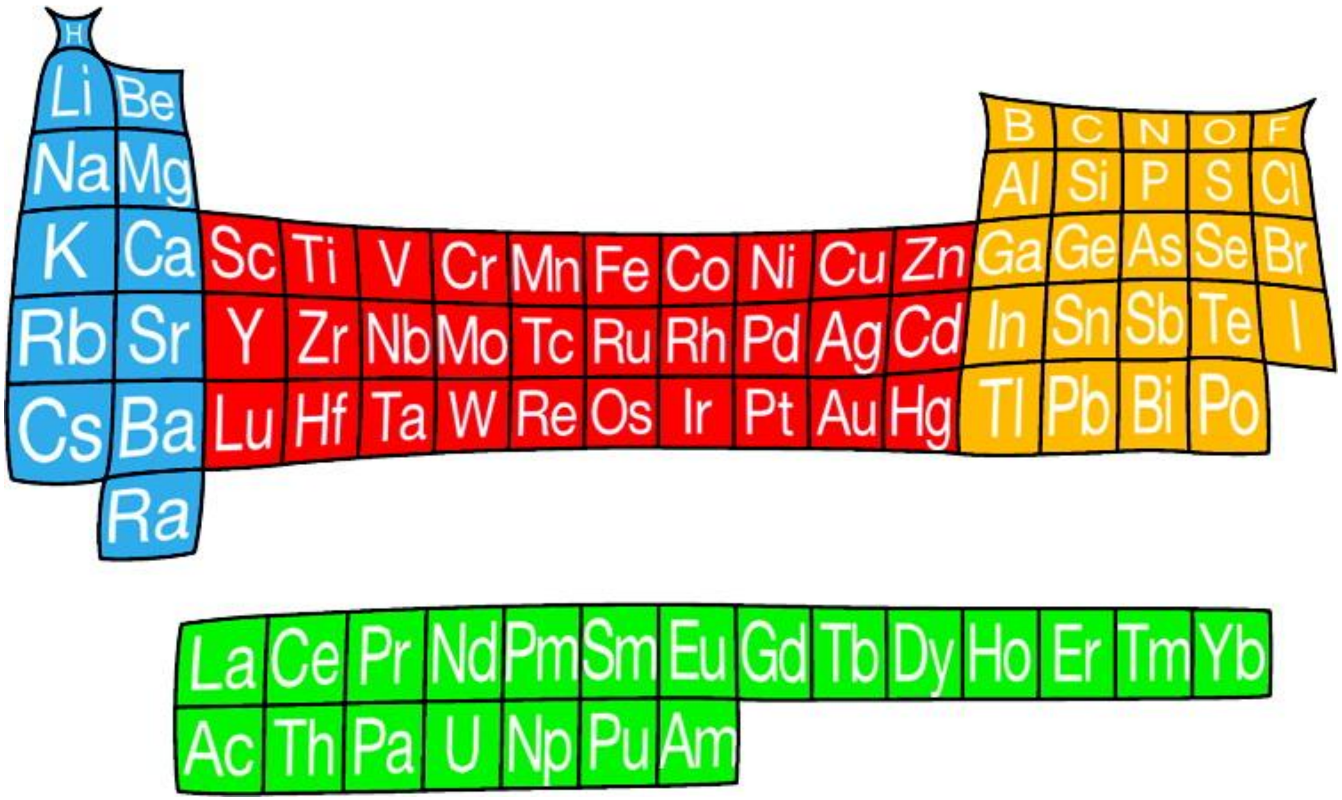
PERIODS																		
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 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
6	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
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Uun	111 Uuu	112 Uub						

Inner transition metals

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

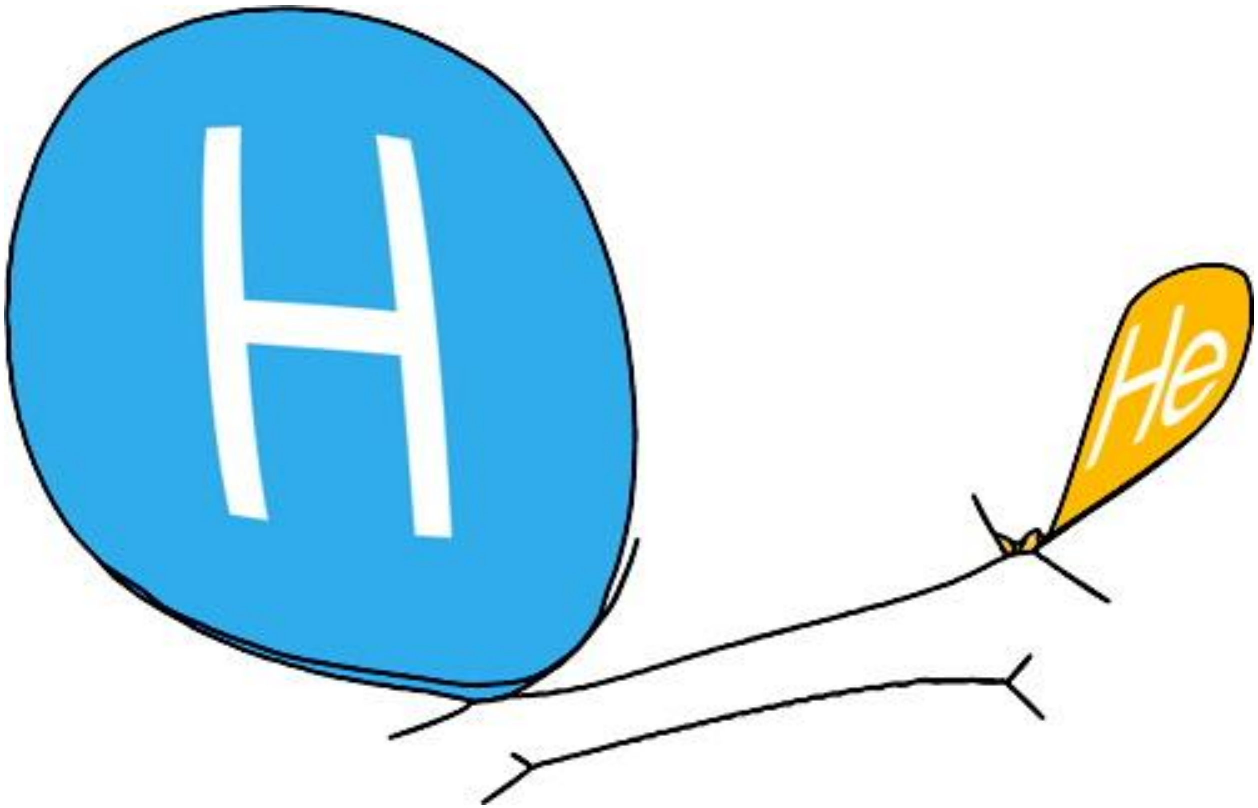
Useful Periodic Table Tools

- http://periodictable.com/
  - Interactive periodic table with pictures of all elements
  - Also available as an app for the iPad ("The Elements a Visual Exploration"- $13.99 on iTunes)
- Nova Elements app (free from iTunes)
  - Also interactive, allows you to build elements, also to build molecules in "Essential Elements"



Diffusion cartogram representing atomic radius.

Published in: Mark J. Winter; *J. Chem. Educ.* Article ASAP



Diffusion cartogram representing abundance in the universe by atom numbers.

Published in: Mark J. Winter; *J. Chem. Educ.* Article ASAP

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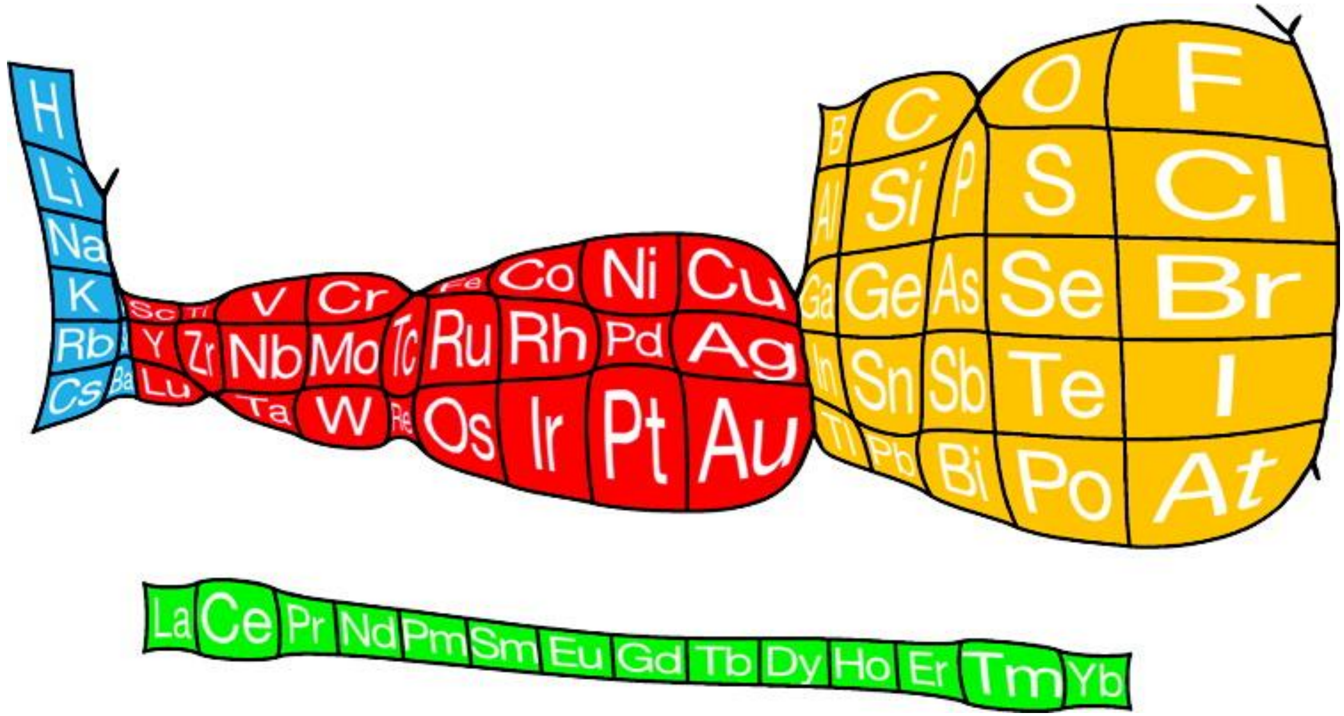


Diffusion cartogram representing electron affinity.

Published in: Mark J. Winter; *J. Chem. Educ.* Article ASAP

The Quantum Hypothesis

Quantum Hypothesis

Max Planck, German physicist, hypothesized—

warm bodies emit radiant energy in discrete bundles called *quanta*. Energy in each energy bundle is proportional to the frequency of radiation.

Einstein stated that light itself is quantized. A beam of light is not a continuous stream of energy but consists of countless small discrete quanta of energy, each quantum called a *photon*.

The Quantum Hypothesis

Is light a wave, or a stream of particles?

Light can be described by both models—it exhibits properties of both a wave or a particle, depending on the experiment.

The amount of energy in a photon is directly proportional to the frequency of light:

$$E \sim f$$

The Quantum Hypothesis

Using the quantum hypothesis:

- Danish physicist Niels Bohr explained the formation of atomic spectra as follows:
  - —The potential energy of an electron depends on its distance from the nucleus.
  - —When an atom absorbs a photon of light, it absorbs energy. Then a low-potential-energy electron is boosted to become a high-potential-energy electron.

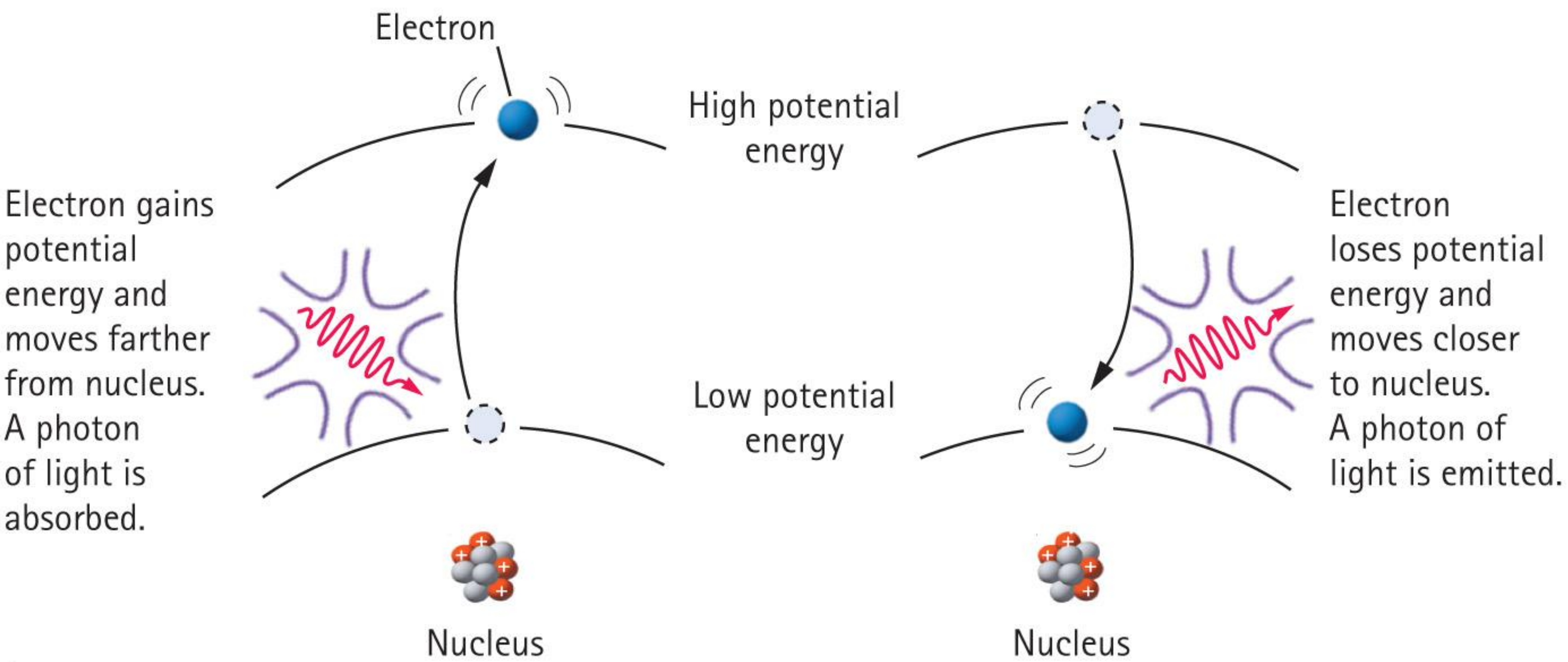
The Quantum Hypothesis

Using quantum hypothesis:

- When an electron in any energy level drops closer to the nucleus, it emits a photon of light.
- Bohr reasoned that there must be a number of distinct energy levels within the atom. Each energy level has a principal quantum number *n,* where *n* is always an integer. The lowest level is $n = 1$ and is closest to the nucleus.

Electrons release energy in discrete amounts that form discrete lines in the atom's spectrum.

The Quantum Hypothesis



The Quantum Hypothesis

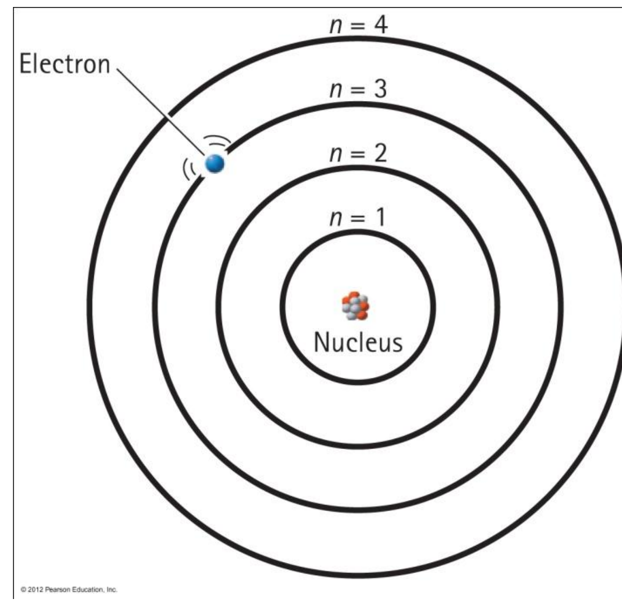
Bohr's model explains why atoms don't collapse:

- Electrons can lose only specific amounts of energy equivalent to transitions between levels.
- An atom reaches the lowest energy level called the *ground state,* where the electron can't lose more energy and can't move closer to the nucleus.

The Quantum Hypothesis

Planetary model of the atom:

Photons are emitted by atoms as electrons move from higher-energy outer levels to lower-energy inner levels. The energy of an emitted photon is equal to the difference in energy between the two levels. Because an electron is restricted to discrete levels, only lights of distinct frequencies are emitted.



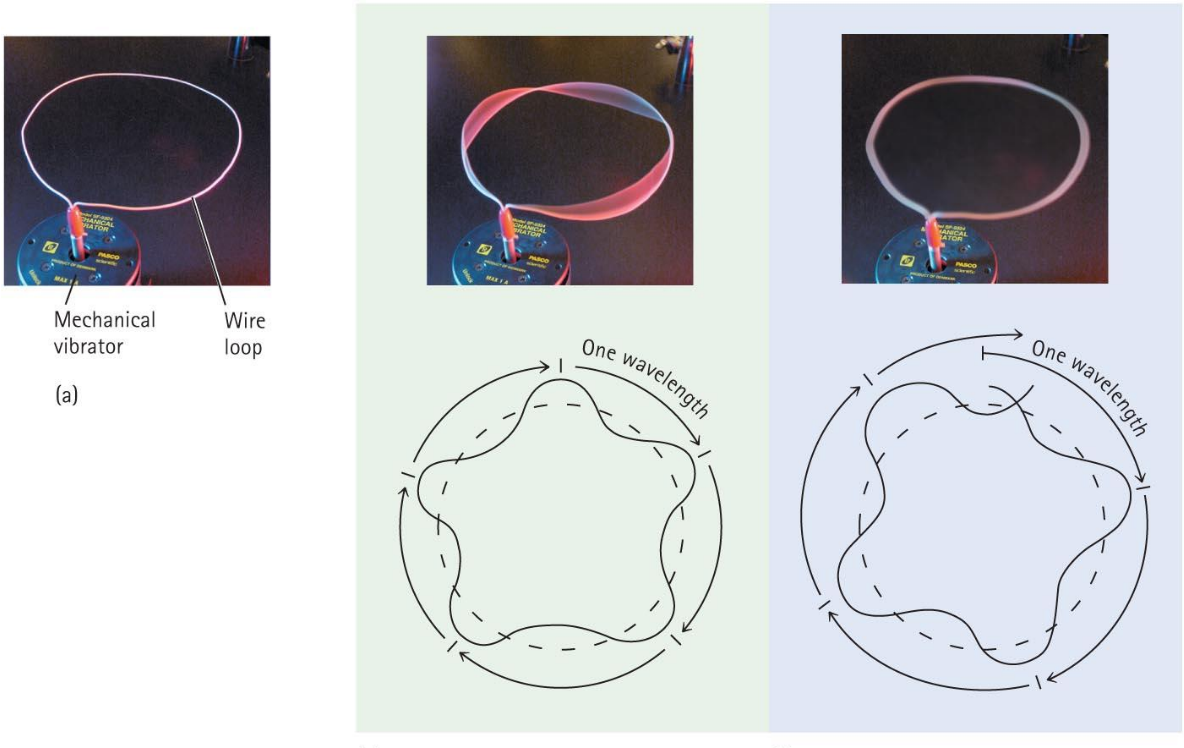
Electron Waves

An electron's wave nature explains why electrons in an atom are restricted to particular energy levels. Permitted energy levels are a natural consequence of standing electron waves closing in on themselves in a synchronized manner.

The orbit for $n = 1$ consists of a single wavelength, $n = 2$ is of two wavelengths, and so on.

Electron Waves

For a fixed circumference, only an integral number of standing waves can occur, and likewise in the paths of electrons about the nucleus.



(b) Wavelength is self-reinforcing. (c) Wavelength produces chaotic motion.

Quantum Numbers

- Each element in its ground state has a unique set of *four quantum numbers*
- Quantum numbers deal with the *probability* of finding an electron within a given region of space outside of the nucleus
- The arrangement of the electrons is called the *electron configuration*

Quantum Numbers

- Principal quantum number – *n*
  - Positive integer values
  - As *n* increases, the size of the orbital increases
  - Increasing *n* means electron is at higher energy
- Angular momentum quantum number – *l*
  - Can have positive integer values from 0 to (*n*-1)
  - Defines the shape of the orbital
  - Value of *l* is designated by the letters *s, p, d, f*

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

Quantum Numbers

- Magnetic quantum number- m_l
  - Can have integral values between –*l* and +*l*, including 0
  - Describes the orientation of the orbital in space
- Spin quantum number – m_s
  - Can have values of +1/2 or -1/2
  - Indicates the two opposite directions an electron can spin

Quantum Numbers

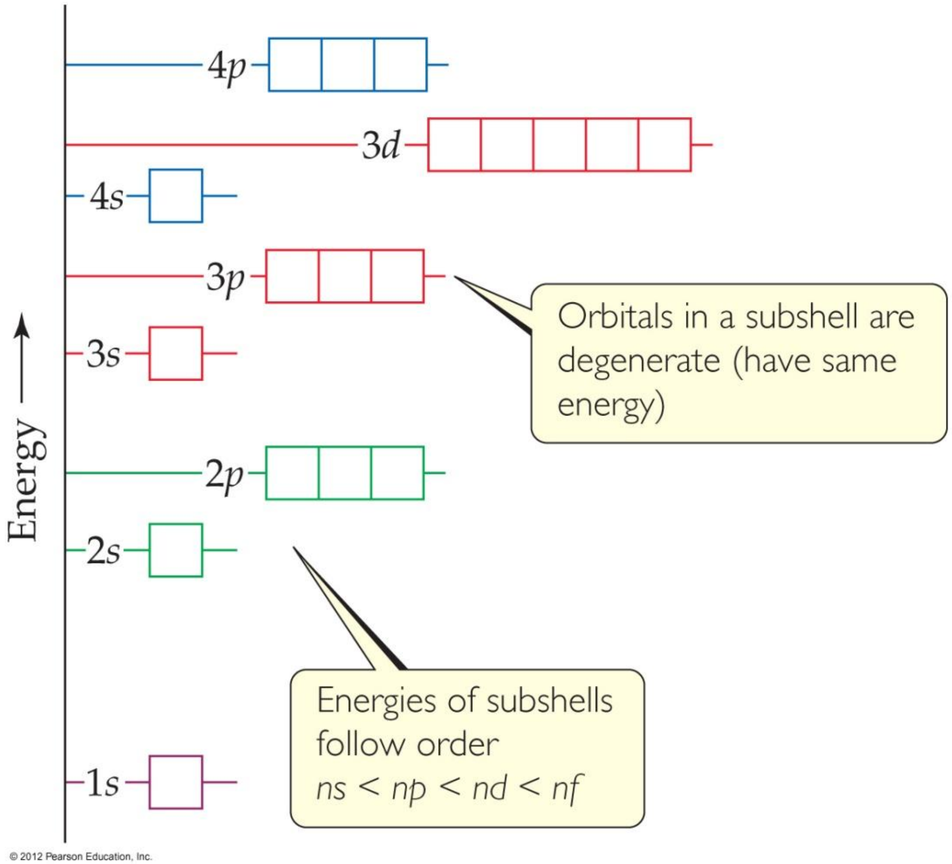
TABLE 6.2 • Relationship among Values of *n*, *l*, and m_l through *n* = 4

n	Possible Values of *l*	Subshell Designation	Possible Values of m_l	Number of Orbitals in Subshell	Total Number of Orbitals in Shell
1	0	1*s*	0	1	1
2	0	2*s*	0	1	
	1	2*p*	1, 0, −1	3	4
3	0	3*s*	0	1	
	1	3*p*	1, 0, −1	3	
	2	3*d*	2, 1, 0, −1,−2	5	9
4	0	4*s*	0	1	
	1	4*p*	1, 0, −1	3	
	2	4*d*	2, 1, 0, −1, −2	5	
	3	4*f*	3, 2, 1, 0, −1, −2, −3	7	16

Orbital Diagrams

- Electron shell- set of electrons with the same value of *n*
- Subshell- electrons that have the same *n* and *l* values
  - A shell with a principal quantum number *n* will consist of exactly *n* subshells
- Orbital- wave functions that describe the spatial distribution of electrons
  - Defined by the values of *n*, *l*, and m_l
- Orbital diagrams represent each orbital with a box and show the electrons as arrows either pointed up (spin up) or down (spin down)

Energies of Orbitals



- As the number of electrons increases, though, so does the repulsion between them.
- Therefore, in many-electron atoms, orbitals on the same energy level are no longer degenerate.

Rules

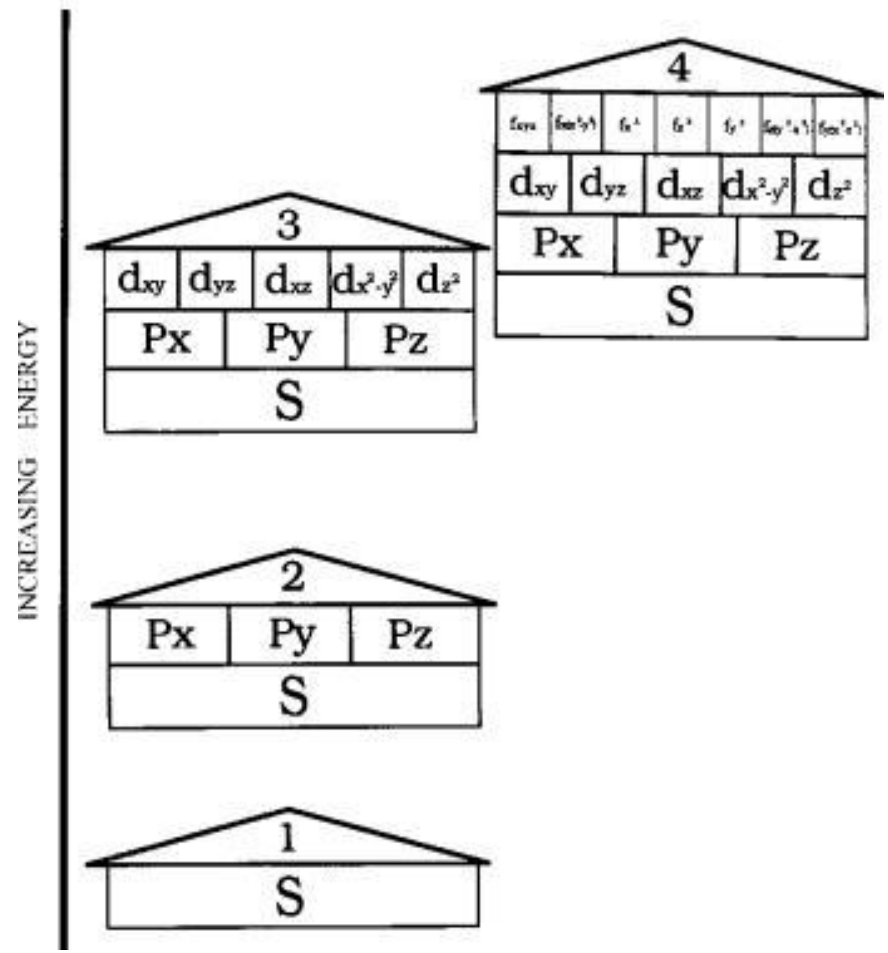
- Pauli exclusion principle states that no two electrons can have the same set of four quantum numbers n, l, m_l, and m_s
- Each orbital can hold a maximum of two electrons
- Hund's Rule states that each orbital in a sublevel must be singly filled first and all electrons in the singly occupied orbitals must have the same spin

Examples

- Draw the orbital diagram for
  - Carbon
  - Phosphorus
  - Lithium
  - Calcium

“Housing Electrons” Example

- Each house represents the principal quantum number, *n*
- Each floor in the house represents the second quantum number, *l*
- Each room represents the third quantum number or orbital, m_l
- Each electron has its own “address”



Garafalo, A. *J. Chem. Educ.* **1997**, *74*, 709-710

Electron Configurations

- A particular arrangement of electrons in an atom
- Most stable electron configuration of an atom is the ground state, where all of the electrons are in the lowest possible energy states
- We represent the electron configuration by writing the energy level (principal quantum number), followed by the symbol for the subshell (s, p, d, f), adding a superscript to indicate the number of electrons in that subshell

Electron Configurations

$$4p^5$$

- This term shows the distribution of all electrons in an atom.
- Each component consists of
  - A number denoting the energy level,
  - A letter denoting the type of orbital,
  - A superscript denoting the number of electrons in those orbitals.

Electron Configurations

TABLE 6.3 • Electron Configurations of Several Lighter Elements

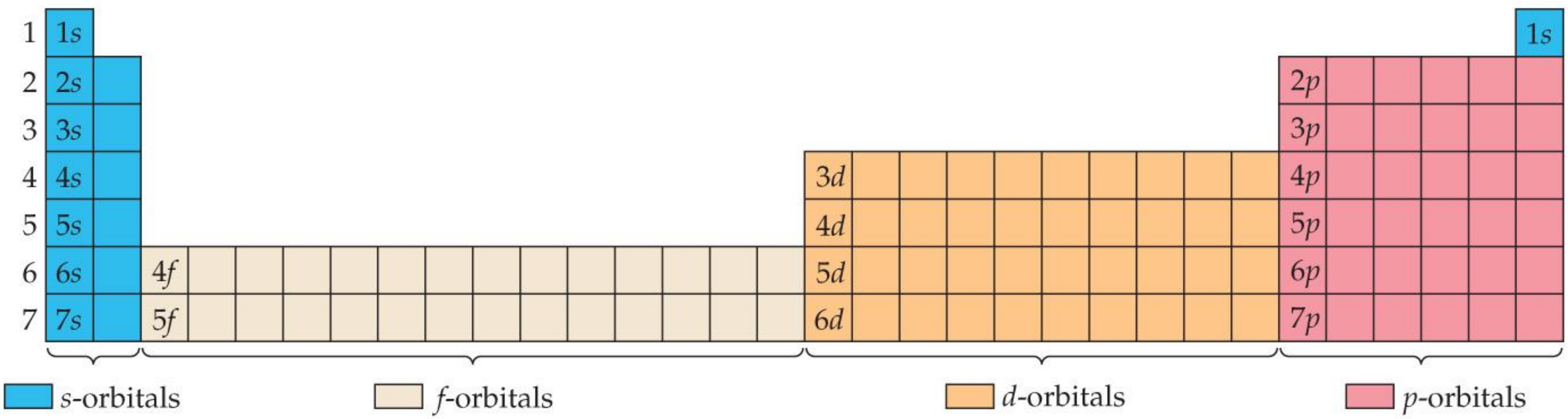
Element	Total Electrons	Orbital Diagram 1s	2s	2p	3s	Electron Configuration
		1*s*	2*s*	2*p*	3*s*	
Li	3	↑↓	↑	[ ] [ ] [ ]	[ ]	$1s^2 2s^1$
Be	4	↑↓	↑↓	[ ] [ ] [ ]	[ ]	$1s^2 2s^2$
B	5	↑↓	↑↓	[↑] [ ] [ ]	[ ]	$1s^2 2s^2 2p^1$
C	6	↑↓	↑↓	[↑] [↑] [ ]	[ ]	$1s^2 2s^2 2p^2$
N	7	↑↓	↑↓	[↑] [↑] [↑]	[ ]	$1s^2 2s^2 2p^3$
Ne	10	↑↓	↑↓	[↑↓] [↑↓] [↑↓]	[ ]	$1s^2 2s^2 2p^6$
Na	11	↑↓	↑↓	[↑↓] [↑↓] [↑↓]	↑	$1s^2 2s^2 2p^6 3s^1$

Examples

- Write the electron configuration for
  - Carbon
  - Phosphorus
  - Lithium
  - Calcium
- Reference periodic table with orbital designations

Periodic Table

- We fill orbitals in increasing order of energy.
- Different blocks on the periodic table (shaded in different colors in this chart) correspond to different types of orbitals.



Identifying Elements Using the Spectroscope

Spectroscope:

- an instrument that separates and spreads light into its component frequencies
- allows analysis of light emitted by elements when they are made to glow—identifies each element by its characteristic pattern

Each element emits a distinctive glow when energized and displays a distinctive spectrum.

Identifying Elements Using the Spectroscope

Atomic spectrum

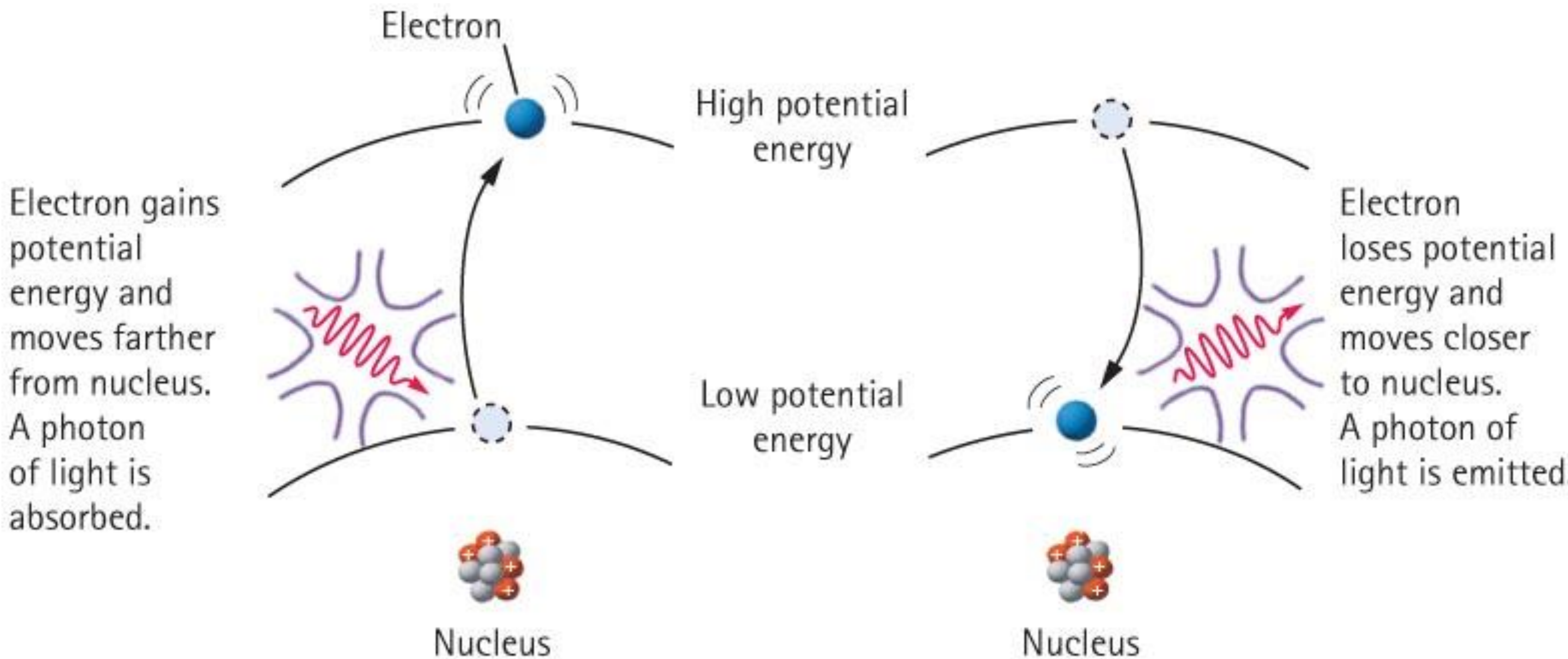
is an element's fingerprint—a pattern of discrete (distinct) frequencies of light.

Discoveries of atomic spectrum of hydrogen:

- A researcher in the 1800s noted that hydrogen has a more orderly atomic spectrum than others.
- Johann Balmer expressed line positions by a mathematical formula.
- Johannes Rydberg noted that the sum of the frequencies of two lines often equals the frequency of a third line.

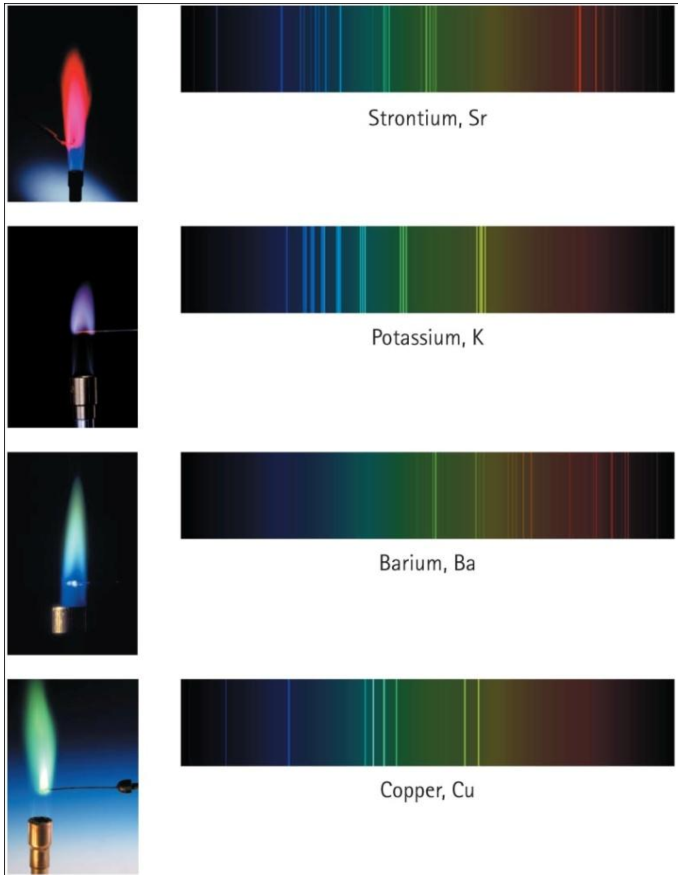
Identifying Atoms Using the Spectroscope

Atomic Excitation

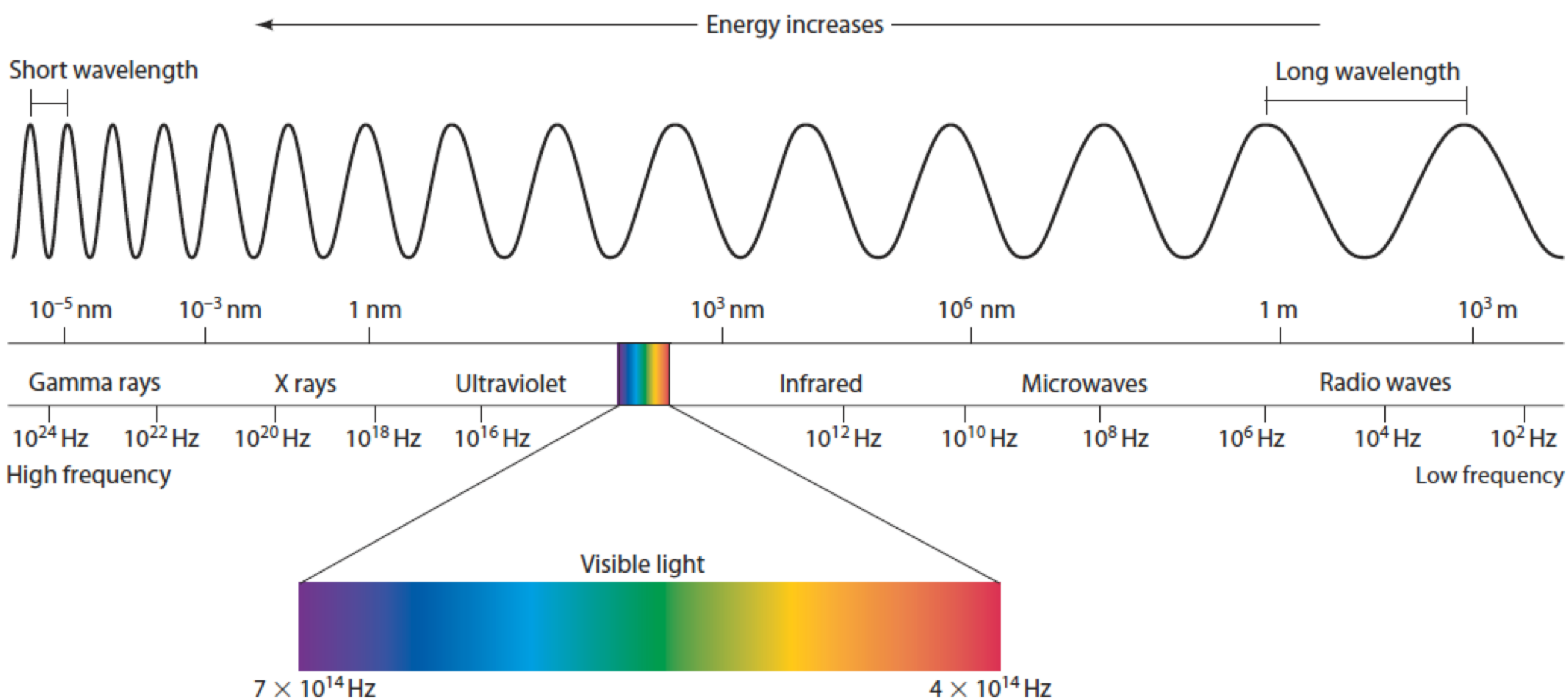


Identifying Atoms Using the Spectroscope

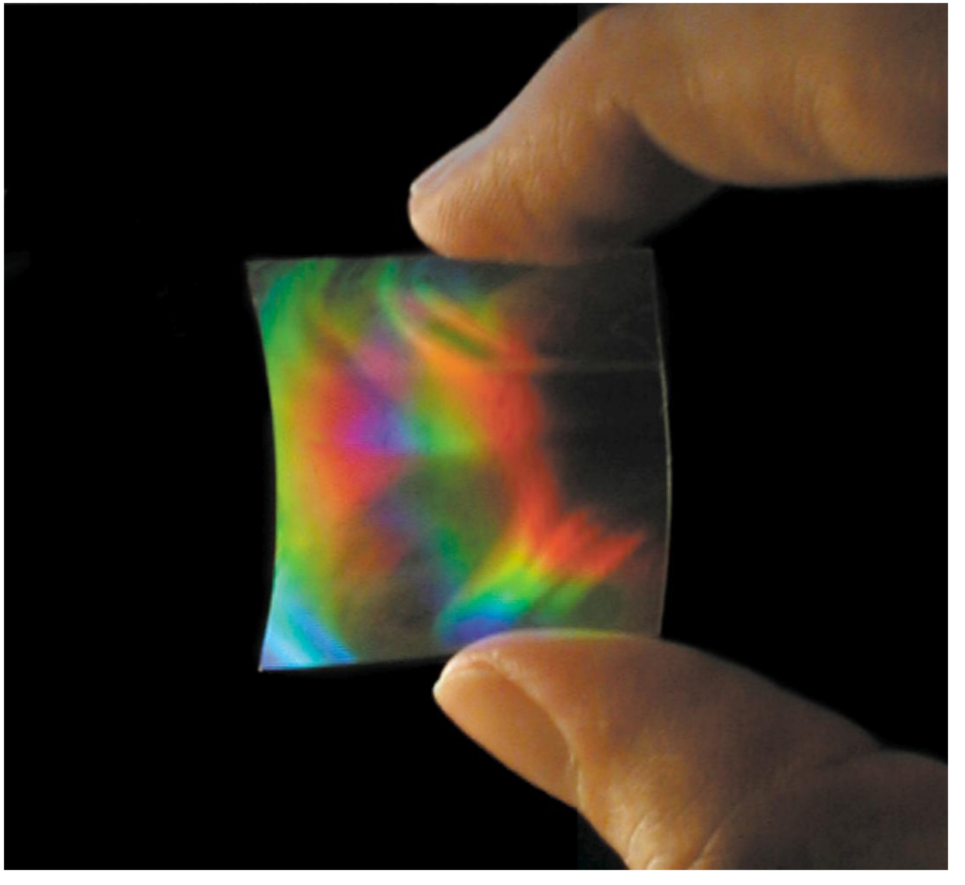
Spectral Lines of Various Elements



The full range of frequencies and wavelengths of electromagnetic radiation is known as the **electromagnetic spectrum**.

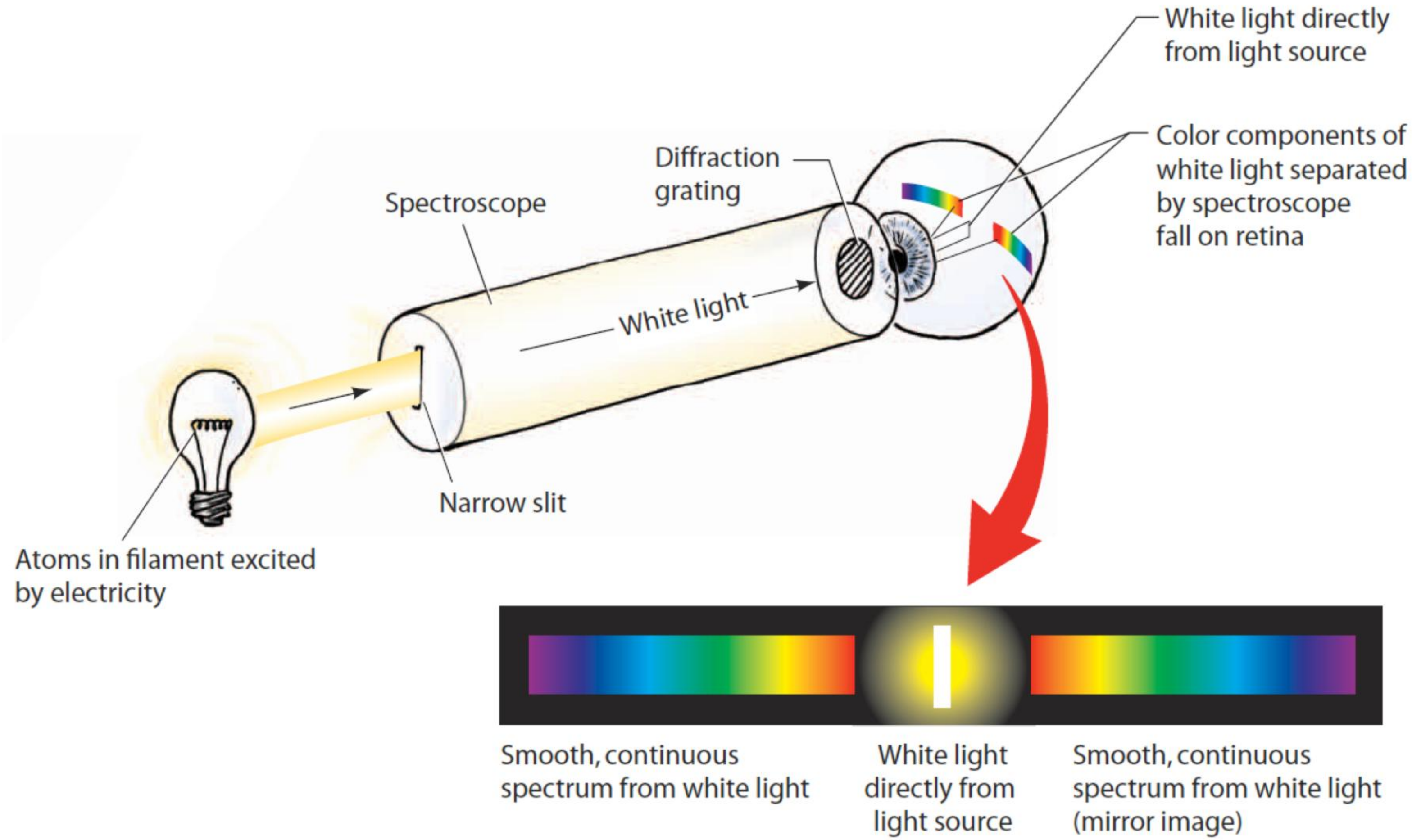


A **spectroscope** is an instrument used to observe the color components of any light source.



Diffraction grating

A **spectroscope** is an instrument used to observe the color components of any light source.



A **spectroscope** is an instrument used to observe the color components of any light source.

The spectroscope allows us to analyze the frequencies of light emitted by elements as they are made to glow.

Bright Lights Demonstration