BAHIR DAR UNIVERSITY

COLLEGE OF SCIENCE

DEPARTMENT OF CHEMISTRY

Pre-University Remedial program for the 2014 E.C.

ESSLCE Examinee

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1.1. The concept of the atom

- The concept of the atom has evolved over time, from the early Greek philosophers belief that matter was made up of indivisible particles to the modern understanding of the atoms substructure and behavior.
- An atom is the smallest unit of an element that is capable of existence, either alone or in chemical combination with other atoms of the same or another element.
- The fundamental particles those every atoms composed are the proton, electron and neutron.

Dalton's atomic theory

John Dalton was the first to put forward a scientific theory about the physical properties of the atom. Dalton was formulated his basic postulates based on his experimental results.

1. All matter is made up of small particles called atoms
2. Atoms are indivisible, neither created nor destroyed
3. All atoms of the same elements are identical in mass and other properties
4. Atoms of different elements have different properties
5. Compounds are formed when atoms are combined in a small whole number ratio.

Modern atomic theory

1. All matter is made up of small particles called atoms

2. Atoms are divisible into smaller particles called proton, electron and neutron

3. All atoms of the same elements may not have the same properties. e.g. mass

4. Atoms of different elements have different properties

5. Compounds are formed when atoms are combined in a small whole number ratio.

The two shortcomings of Dalton's atomic theory are:

1. Atoms are indivisible, neither created nor destroyed

2. All atoms of the same elements are identical in mass and other properties

1.2. The discoveries of subatomic particles

1. The discoveries of subatomic particles began in the late 19th century with the discovery of electron by J.J. Thomson in 1897. This was followed by the discovery of proton be Ernest Rutherford in 1917, and neutron by James Chadwick in 1932.

J.J. Thomson: discovered the electron in 1897 through his famous cathode ray tube experiments.

- ✓ Cathode rays are streams of negatively charged particles that are produced when an electric current is passed through a vacuum tube filled with gas at low pressure.

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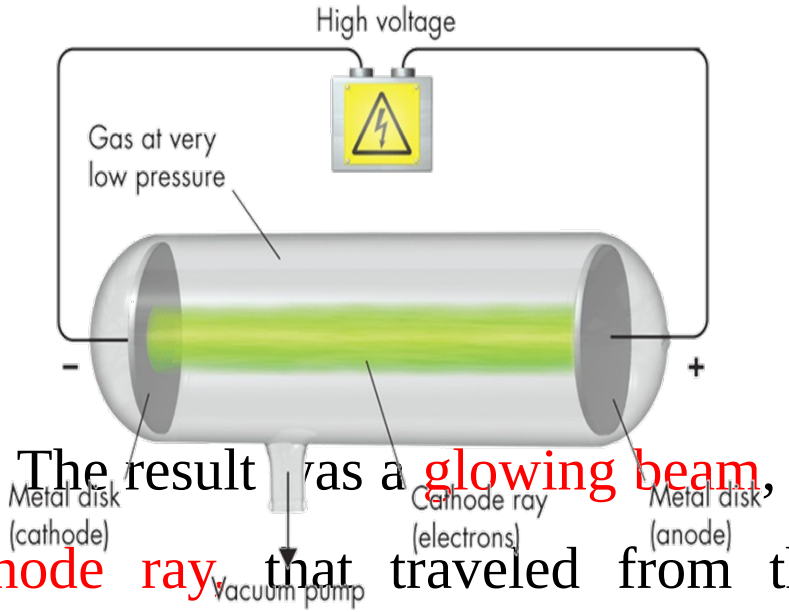
- ✓ Thomson observed that the cathode rays were deflected by electric and magnetic fields in a way that suggested they were made up of negatively charged particle, they are now known as electrons.
- The electrodes were connected to a source of electricity.



The result was a glowing beam, or cathode ray, that traveled from the cathode to the anode.

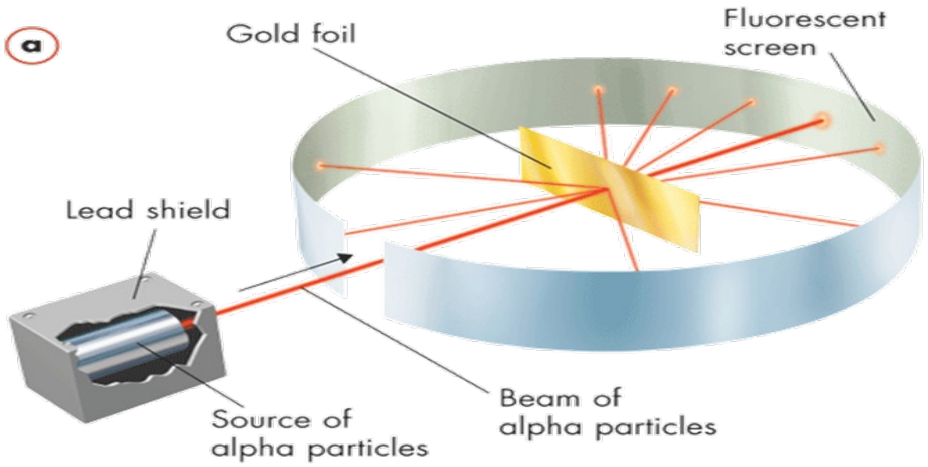
- Cathode rays normally travel in straight lines, but are deflected when a magnet is brought nearby.

- Robert A. Millikan (1868–1953) determined the charge on the electron by observing the behaviour of electrically-charged oil drops in an electric field.
- Based on careful experiments, Millikan established the charge on an electron as $e = -1.602 \times 10^{-19}$ C.
- Using this charge and Thomson's charge-to-mass ratio of an electron, Millikan calculated an electron's mass.
- Millikan's values for electron charge and mass are similar to those accepted today.

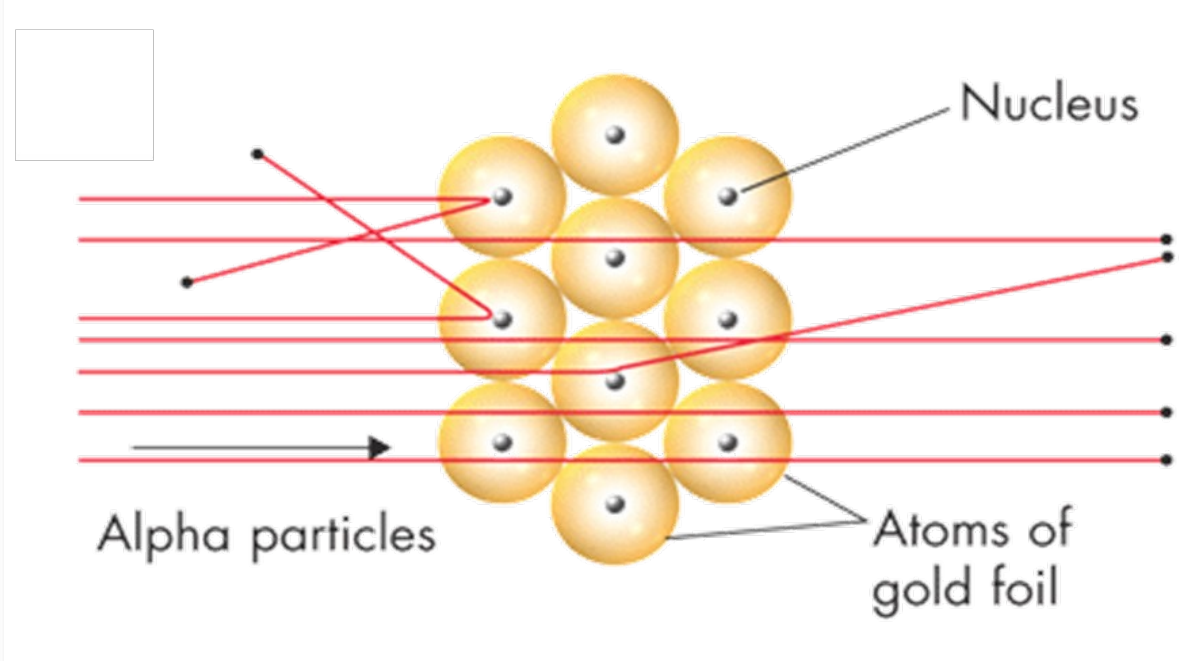
Cont…

<u>Ernest Rutherford</u>: a New Zeeland physicist

- ✓ He bombarded a thin sheet of gold foil with alpha particles (positively charged particle) emitted from a radioactive source.
- ✓ Rutherford expected the alpha particles to pass through the gold foil with the minimal deflection.



- The Rutherford experiment showed that some alpha particles were deflected at large angles, while a few even bounced straight back.
- This result suggested that the positively charged alpha particles were encountering a concentrated positive charge within the atom.



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- ✓ To explain this phenomenon, Rutherford proposed a new model of the atom known as the Rutherford model.
- ✓ In this model, he suggested that the atom had the small, dense, positively charged nucleus at its center, which was surrounded by negatively charged electrons in orbit.
- ✓ He concluded that all the positive charge and almost all of the mass are concentrated in a small region that has enough positive charge to account for the great deflection of some of the alpha particles.

Cont…

James Chadwick: discovered the neutron in 1932

- ✓ He had been experimenting with the atomic nucleus, trying to understanding its structure and properties.
- ✓ At the time, know that the atomic nucleus was made up of positively charged protons, but they couldn't explain why it did not fly apart due to electrostatic repulsion between the protons.

Cont.…

- ✓ Chadwick proposed that there must be a neutral particle, which he called the neutron, that was also present in the nucleus and held to bind the protons together.
- ✓ He conducted a series of experiments in which he bombarded beryllium with alpha particles then the radiation of beam of a neutral particle was emitted.
- ✓ showed that this radiation was best explained as a beam of neutral particles (Neutron).
- ✓ These particles were found to have about the same mass as protons but no electric charge.

- Eugen Goldstein (1850–1930) observed a cathode-ray tube and found rays traveling in the direction opposite to that of the cathode rays.
- He concluded that they were composed of positive particles.
- Such positively charged subatomic particles are called protons.

Properties of Subatomic Particles

Particle	Symbol	Relative charge	Relative mass (mass of proton = 1)	Actual mass (g)
Electron	e^-	1–	1/1840	9.11×10^{-28}
Proton	p^+	1+	1	1.67×10^{-24}
Neutron	n^0	0	1	1.67×10^{-24}

Atomic mass and isotopes

- ✓ A shorthand method of showing the atomic number and mass number of a nuclide along with its symbol, X, is:

  Where, A is mass number

  Z is atomic number. e.g.

- ✓ In a neutral atom the atomic number (Z) = number of protons in the nucleus = number of electrons

  Mass number (A) = number of protons + number of neutrons

  Number of neutrons = A - Z

Cont.…

- An isotope is the variation of an element that has the same numbers of protons but a different numbers of neutrons in the nucleus.
- This means that atoms of the same element have the same atomic number but different atomic mass. E.g. carbon-12 and carbon-13
- The average/ relative atomic mass (Ar) of an element is the sum of the masses of its isotopes, each multiplied by its natural abundances

e.g. Calculate the Ar for naturally occurring chlorine if the distribution of isotopes is 75.77% and 24.23% .

Ar = = 35.45

Exercise :

1. Consider the species ^{72}Zn, ^{75}As and ^{74}Ge. These species have: (Atomic number: Zn = 30, As = 33, and Ge = 32)

A. the same number of electrons.
B. the same number of protons and neutrons.
C. the same number of protons.
D. the same mass number.
E. the same number of neutrons.

2. Lithium has an elemental atomic mass of 6.941 amu and has two naturally occurring isotopes, 6Li and 7Li. Their masses are 6.0151 amu and 7.0160 amu respectively. What are the natural abundances of the isotopes of Lithium?

Electromagnetic radiation

- Maxwell (1873), proposed that visible light consists of electromagnetic waves.

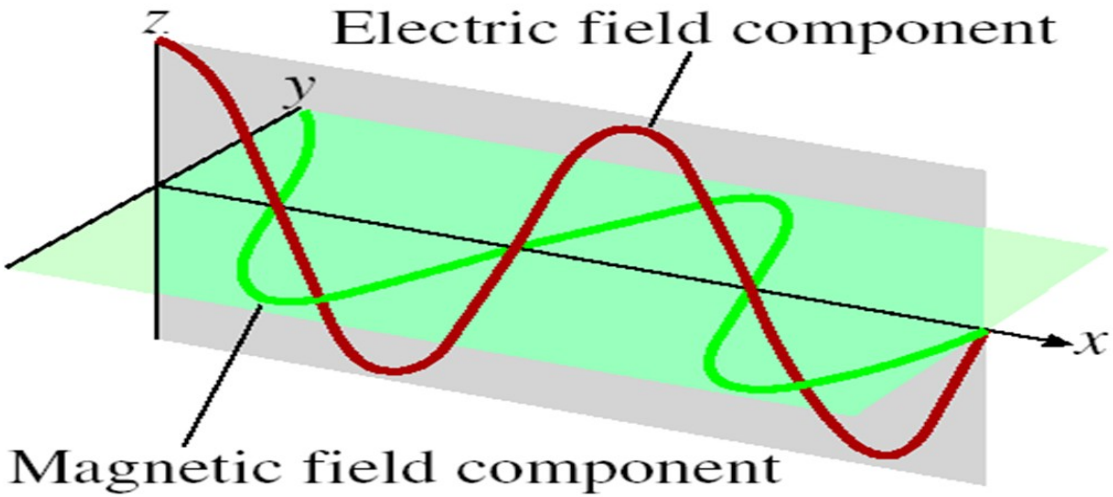


Fig ong the z axis. The electric E and magnetic H fields oscillate in the x-y plane and perpendicular to the direction of propagation.

- Electromagnetic radiation is the emission and transmission of energy in the form of electromagnetic waves.

- The wave properties of electromagnetic radiation are described by two interdependent variables, frequency and wavelength.
- Wavelength: the distance b/n identical points on successive waves.
- Frequency (v) is the number of waves that pass through a particular point in space per second
- Amplitude is the vertical distance from the midline of a wave to the peak or trough.

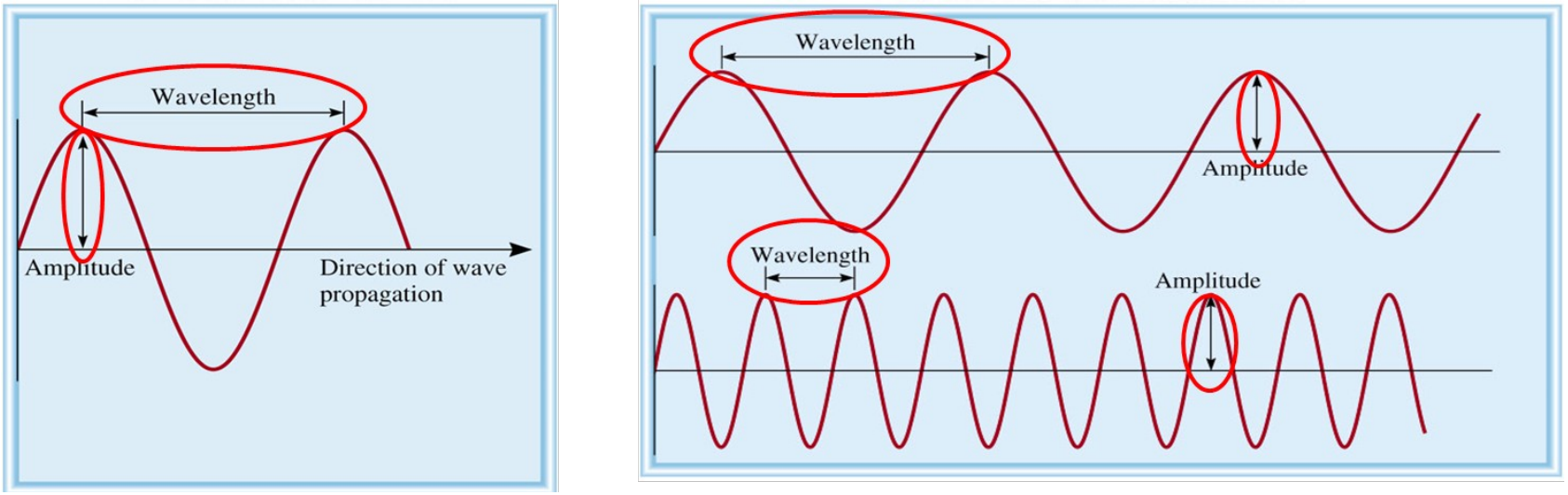
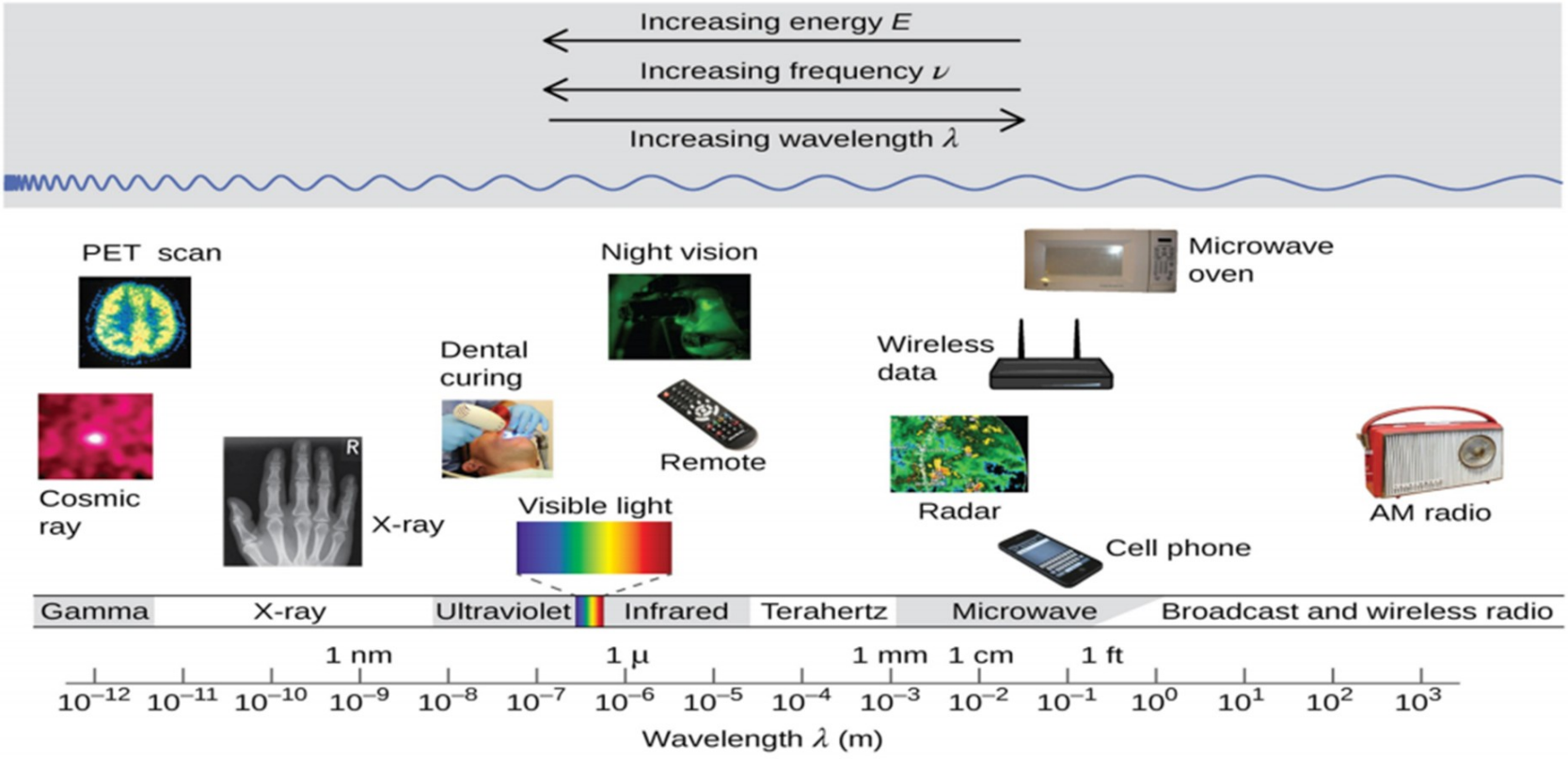


Fig. 2. (a) Wavelength and amplitude. (b) Two waves having different wavelengths and frequencies.

- The speed of the electromagnetic wave (light), c , is the product of its frequency)and its wavelength.
- The speed (u) of the wave = x v



The Quantum Theory and Photon

- The quantum theory is concerned with the rules that govern the gain or loss of energy from an object.
- When solids are heated, they emit electromagnetic radiation over a wide range of wavelengths.
- He proposed that a hot glowing object could emit (or absorb) only certain amounts of energy.
- $E = nh\nu$

where E is the energy of the radiation, v is the frequency, n is a positive integer (n = 1, 2, 3, …) called quantum number and h is proportionality constant now called Planck's constant. h = 6.626 × 10−34 J s).

- Energy (light) is emitted or absorbed in discrete units (quantum).

- **The photoelectric effect:**

-Solved by Einstein in 1905

- a phenomenon in which electrons are ejected from the surface of certain metals exposed to light of at least a certain minimum frequency.

- Photon is a “particle” of light

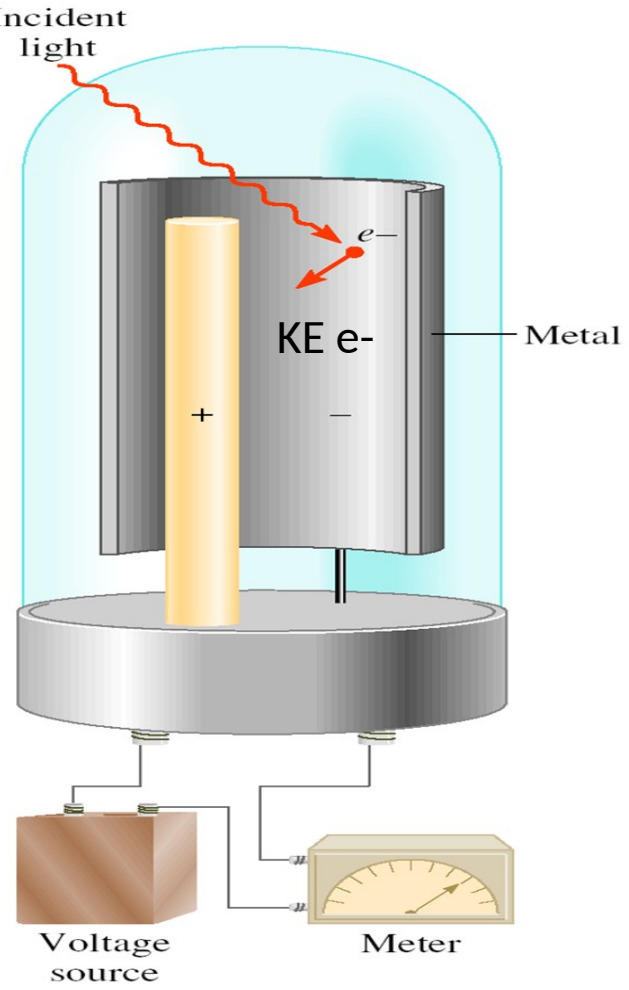
$$h\nu = KE + W$$

- $h\nu = h\nu_o + \frac{1}{2} m_e v^2$

- v_o is the threshold frequency,

  m_e is the mass of the electron

  and v is the velocity of the emitted electron.



Questions

1. Infrared radiation is part of the electromagnetic spectrum. Name two other radiations that are part of the electromagnetic spectrum.

2. Which of these groups has the highest frequency?

A. Radio waves B. Ultraviolet C. Visible light

3. List the following types of electromagnetic radiation in order of decreasing energy per photon.

Microwaves, Visible light, Gamma rays, X-rays, and Radio waves

4. A photon has a frequency of 6.0×10^4 Hz. Convert this frequency into wavelength (nm). (Ans: l = 5.0 x 1012 nm)

5. When copper is bombarded with high-energy electrons, X rays are emitted. Calculate the energy (in joules) associated with the photons if the wavelength of the X rays is 0.154 nm.(Ans: E = 1.29 x 10 -15 J).

6. he work function of a metal surface is 1.55 eV. UV light of wavelength 430 nm is incident on the metal. Calculate the maximum kinetic energy of the photoelectrons. (Ans: KE= 2.15×10^{-19} J)

7. It takes 208.4 kJ of energy to remove 1 mole of electrons from the atoms on the surface of rubidium metal. If rubidium metal is irradiated with 254-nm light, what is the maximum kinetic energy the released electrons can have?

1.3. The Bohr theory of the hydrogen atom

- The Bohr theory of the hydrogen atom was proposed by Niels Bohr in 1913 to explain the spectral lines of the hydrogen atoms.
- It was one of the first attempts to describe the behavior of atoms using quantum mechanics.
- The Bohr model assumes that electrons in hydrogen atoms occupy discrete energy levels. This levels are quantized, meaning that only certain values of energy are allowed.
- When an electron absorbs energy, it jumps to the higher energy level, and when it emits energy, it drops back down to a lower level.

- **Bohr's model of the hydrogen atom is based on four assumptions.**

1.The electron in a hydrogen atom moves in a circular orbit about the nucleus.

2. The hydrogen atom has only certain allowable energy levels (stationary states).

3. Electrons do not give off radiation when they are in an allowed orbit(while in one of its stationary states) . Thus, the orbits are stable.

4. The electron moves to another stationary state (orbit) only by absorbing or emitting a photon whose energy equals the difference in the energy between the two states.

$E = E_f - E_i = h\nu = hc/\lambda$ where E_f & E_i = final & initial orbital energies

Total energy of hydrogen atom: $E_n = -R_H \left(\frac{1}{n^2}\right)$

n (principal quantum number) = 1,2,3,…

R_H (Rydberg constant) = 2.18×10^{-18}J

The radius of the Bohr orbit can be calculated using:

r= (

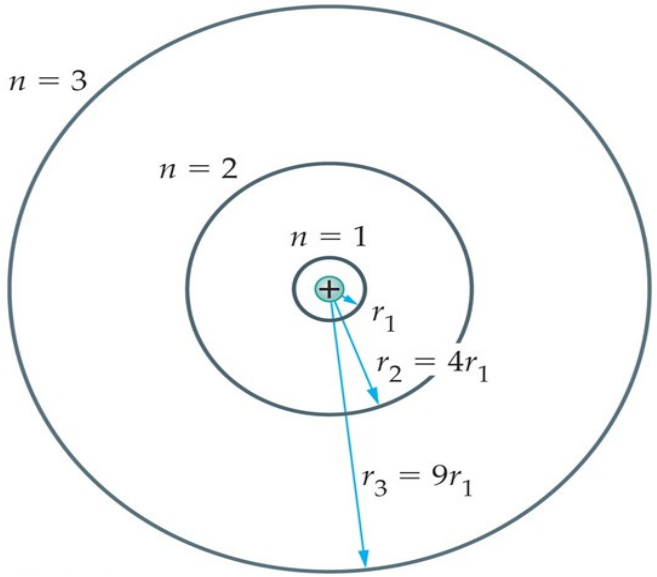
where a_0 = Bohr radius

= 5.292 E^{-11} m

n is a positive integer which is called quantum number.

r is the radius of the orbit

- The higher orbits in the Bohr model increase in radius as n2, as indicated in the figure below.



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- The difference in energy between the levels determines the frequency and wavelength of the absorbed or emitted radiation, according to the equation: E = hv. Where E is energy, h is planks constant and v is frequency.
- Energy is released when an electron in the hydrogen atom drops from the higher to the lower stationary states, the amount of energy and the region (wavelength) can be determined as:

$$E = R_H()$$

Where, Rh is the Rydberg's constant, n_i & n_f are the initial and final states respectively. $R_H = -2.18^{-18} J$

Cont.…

❖ When electron is drops:

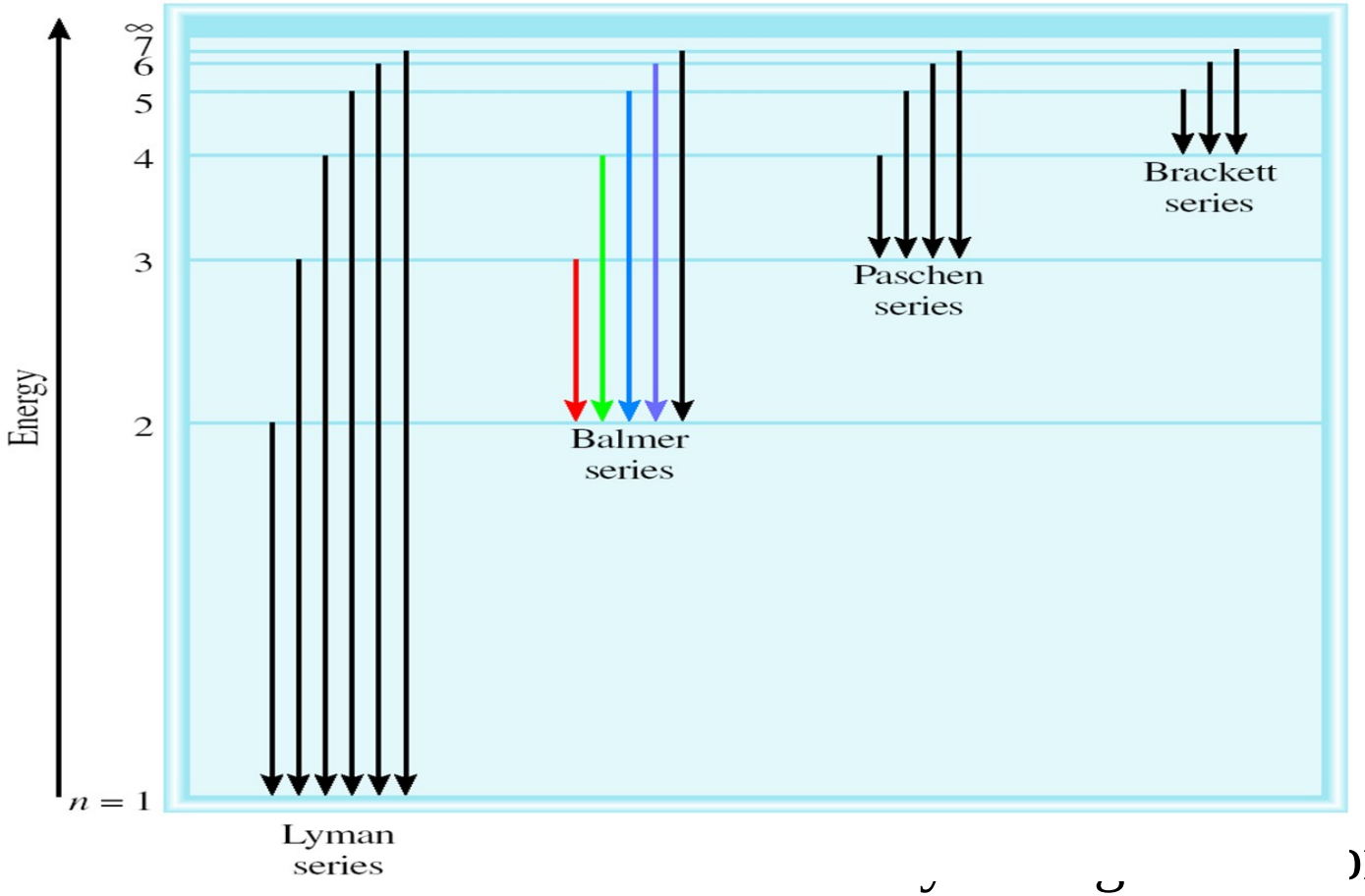
From n = 2, 3, 4… to n = 1 it is called Lyman series (Uv region)

From n = 3, 4,5, … to n = 2 it is called Balmer series (visible region)

From n = 4, 5, 6, … to n = 3 it is called Paschen series (IR region)

From n = 5, 6, 7, … to n = 4 it is called Bracket series (IR region)

From n = 6, 7, 8, … to n = 5 it is called Pfund series (IR region)



$$E_{photon} = \Delta E = E_f - E_i$$

$$E_f = -R_H \left(\frac{1}{n_f^2}\right)$$

$$E_i = -R_H \left(\frac{1}{n_i^2}\right)$$

$$\Delta E = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

…on, it could also be applied to the single electron ions He^{+}, Li^{2+}, Be^{3+}, and so forth, which differ from hydrogen only in their nuclear charges, and so one-electron atoms and ions are collectively referred to as hydrogen-like atoms.

Limitations of Bohr model

- Cannot be used to predict structure and measure relative intensities of spectral lines
- Does not apply to multi-electron atoms
- Provides wrong information about the value of angular momentum of an electron in atom
- Does not follow the Heisenberg uncertainty principle.

Questions

1. What is the energy and the wavelength (λ) of the line in the spectrum of H that represents the movement of an e- from a Bohr orbit of n = 4 to the orbit of n = 6? (Ans: E= 7.561 x10^{-20} J, λ = 2.627 E^{-6} m = 2627 nm infrared)

2. What is the radius, in angstroms, of the orbital of an electron with n = 4 in a hydrogen atom? (Z= +1 for hydrogen (Ans r= 8.46 Å)

3. Which transition results in the longest wavelength of light?

a. n = 5 to n = 2

b. n = 4 to n = 2

c. n = 3 to n = 2

4. The electron of a hydrogen atom is in the n = 3 level. What is its energy?

- De Broglie (1924) proposed that a particle with a mass, m, moving at a speed, v, has a wave nature consistent with a wavelength given by

$$\lambda = h/mu$$

u = velocity of e-

m = mass of e-

Calculate the de Broglie wavelength of an electron that has a velocity of 1.00 × 106 m/s. (electron mass, me = 9.11 × 10–31 kg; h = 6.626 × 10–34 J s)

1.4. The quantum mechanical model of the atom

- The quantum mechanical model of the atom is a mathematical framework that explains the behaviors of electrons in an atom.
- It is also known as the wave mechanical model or Schrodinger model.
- Erwin Schrödinger developed a model of the hydrogen atom based on the wave nature of the electron in the late 1920s.
- The wave equations that are acceptable solutions to the Schrödinger equation are called wave functions.

- Wave function (y) describes:

1. energy of e- with a given y
2. probability of finding e- near a particular point in space.

Quantum Numbers

- Quantum numbers are a set of four parameters that describe the properties of an electron in an atom.

1. Principal quantum number (n): describes the size and the energy of the orbital. It is also referred to as the shell number, and it takes on positive integer values (1, 2, 3, etc.)

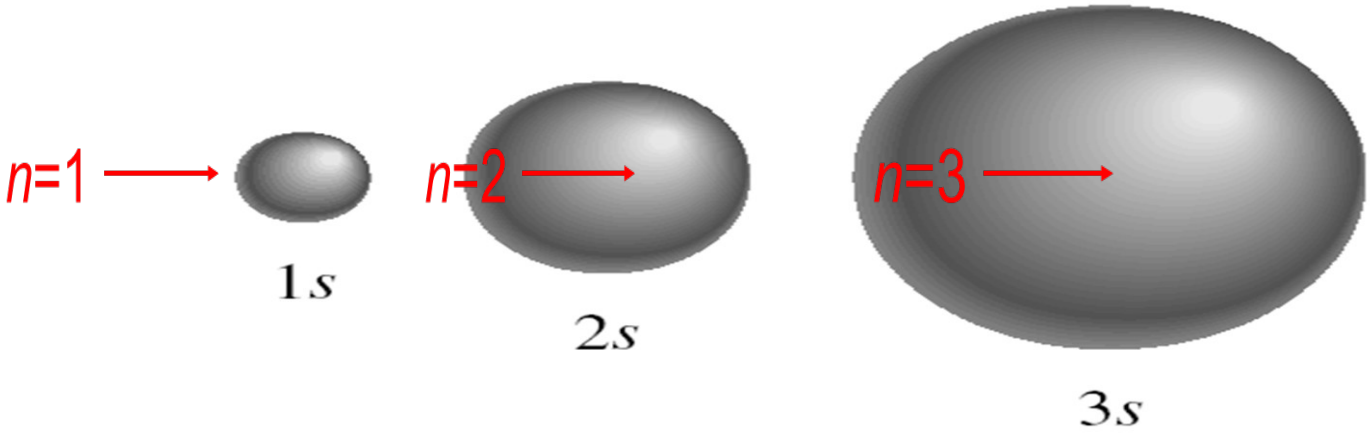
- Size of orbital increases with the increase of principal quantum number n.
- Higher the n value, greater is the energy

Cont….

- For a given principal quantum number, n, the total number of orbitals is determined as: Number of orbitals = n^2 in a shell

For example, the total numbers of orbitals for n = 2 is 4, because $2^2=4$

Exercise: what is the total numbers of orbitals for n = 4?



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2. Azimuthal quantum number (l): describes the shape of an orbital (sometimes called a subshell).

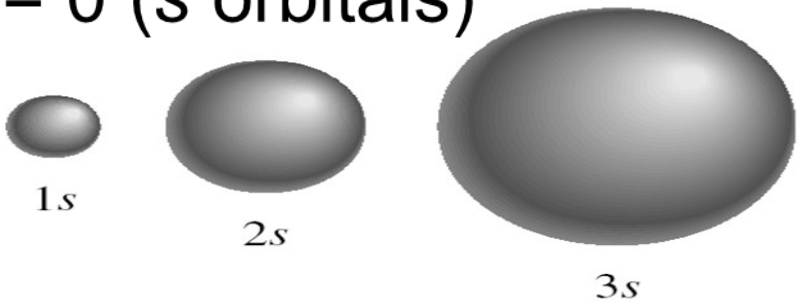
- It takes on integer values from 0 to n-1. each value of l corresponds to a specific subshell (s, p, d and f).

❖ Example what are the possible angular momentum quantum numbers for the sublevel notation of 3p.

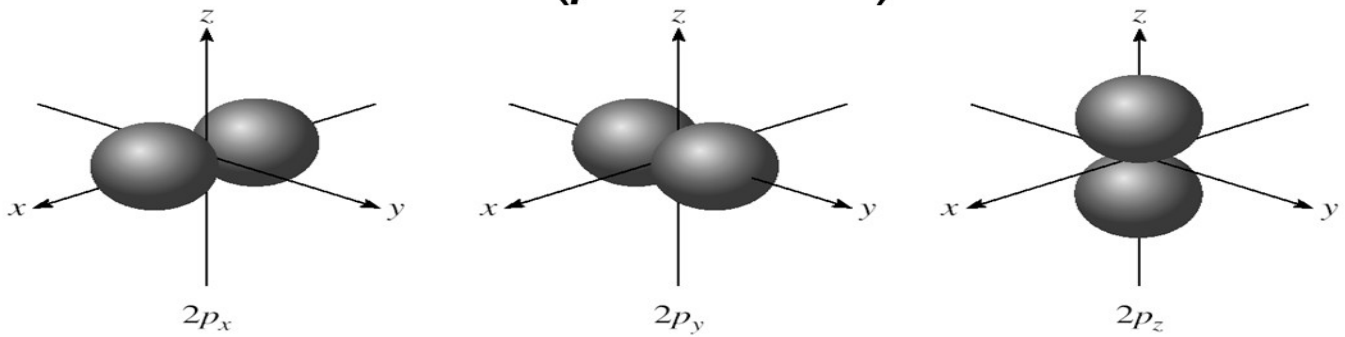
Solution n=3, l is from 0 up to n-1, n-1=2

thus, l= 0, 1 & 2

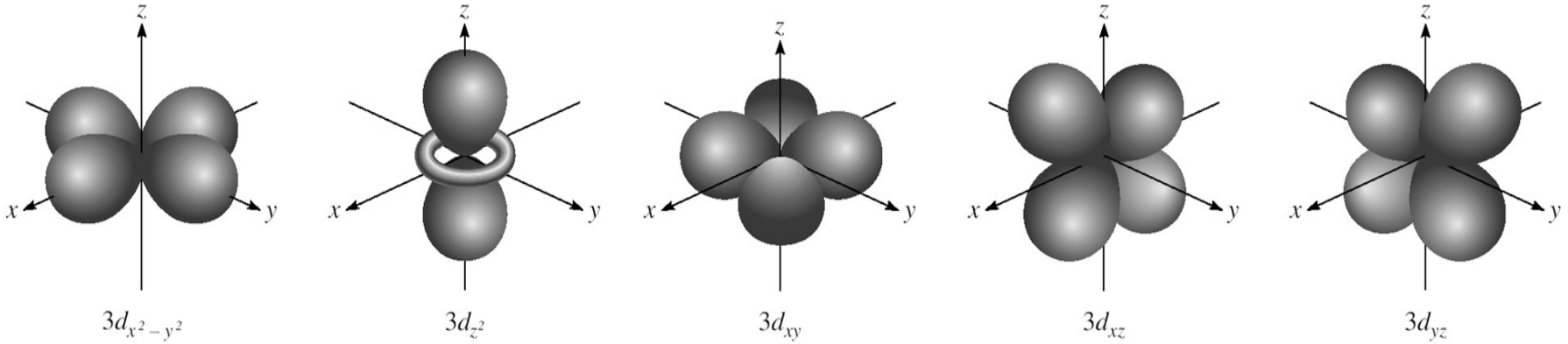
l = 0 (*s* orbitals)



l = 1 (*p* orbitals)



l = 2 (*d* orbitals)



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3. Magnetic quantum number (m_l): describes the orientation of an orbital in space. It takes on integer values from –l to +l, and each values of m_l corresponds to a specific orbital.

Example determine the m_l values for n=3

Solution l=0, 1, 2

m_l value for l=0 is 0, m_l values for l=1 are -1, 0, +1 and m_l values for l=2 are -2, -1, 0, +1, +2

4. Spin quantum number (m_s): describes the angular momentum or spin of an electron. It takes on one of two values, + or -

- Pauli exclusion principle - in a given atom no two electrons can have the same set of four quantum numbers.
- An orbital can hold only two electrons, and they must have opposite spins.

Questions

1. For principal quantum level n = 3, determine the number of allowed subshells, and give the designation of each.
2. For l = 2, determine the magnetic quantum numbers (ml) and the number of orbitals.
3. How many 2p orbitals are there in an atom?
4. How many electrons can be placed in the 3d subshell?
5. How many electrons in an atom could have a set of quantum numbers n=7, l=3 and ml= -1?
6. Which type of orbital has the following combination of quantum numbers?

7. How many 1s orbitals are there in an atom? 4p orbitals? 4d orbitals? 1s: 1 4p: 3 4d: 5

8. What is the maximum number of orbitals with:

n = 4 l = 1 3 (the 4p orbitals)

n = 2 l = 2 none (l must be < n)

n = 3 l = 2 5 (the 3d orbitals)

n = 5 l = 1 ml = -1 1 (3 q.n. define a unique orbital)

9. Describe the electrons defined by the following quantum numbers:

n l ml	
3 0 0	3s electron or orbital
2 1 1	2p electron or orbital
4 2 -1	4d electron or orbital
3 3 2	not allowed (l must be < n)
3 1 2	not allowed (ml must be between -l and l)

11. The orientation in space of an atomic orbital is associated with.

A) The principal quantum number.

B) The angular momentum quantum number.

C) The magnetic quantum number.

D) The Spin Quantum Number.

E) None of the above.

12. Which of the following set of quantum numbers (ordered n, l, ml, ms) are possible for an electron in an atom? Check all that apply.

(a) 3, 1, -2, -1/2 (b) 3, 2, 1, -1 (c) 2, 1, 3, 1/2

(d) 4, 2, -1, -1/2 (e) 3, 2, 0, -1/2 (f) 5, 2, 1, -1/2

(g) 3, 3, 1, -1/2 (h) -3, 2, 2, -1/2

13. How many electrons in an atom can have each of the following quantum numbers or sublevel designations?

(a) n = 2, l = 1, ml = 0 (b) 5s

(c) n = 4, l = 3

1.4.2. Electron configuration and periodic table

- The electronic configuration of an atom describes the distribution of the electrons among atomic orbitals in the atom.
- The notation uses numbers to designate the energy levels or principal quantum number and the letters s, p, d and f to identify the sublevels or subshells.
- A superscript number following a letter indicates the number of electrons in the designated subshell (e)

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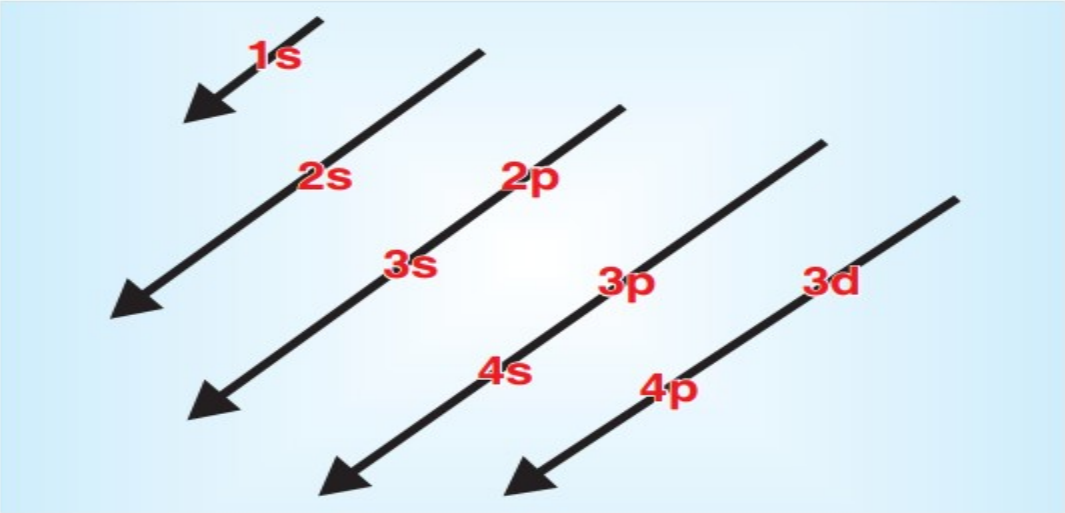
The electron configurationfor any atom follows the following three principles:

1. Aufbau principle: according to this principle electrons occupy the lowest-energy orbital before entering to the higher energy orbital.



For example, the correct electron configuration of Cl is $1s^2 2s^2 2p^6 3s^2 3p^5$,

but not $1s^2 2s^2 2p^6 3s^1 3p^6$.

2. Hund's Principle: This principle states that the degenerated orbitals should be singly occupied before spin paired.

For example, $2p^3$ correct

not correct

p_x p_y p_z

3. Pauli's exclusion principle: No two electrons can have the same four quantum numbers. i.e. they must differ in at least one of the four quantum numbers.

- **Orbital Energies**

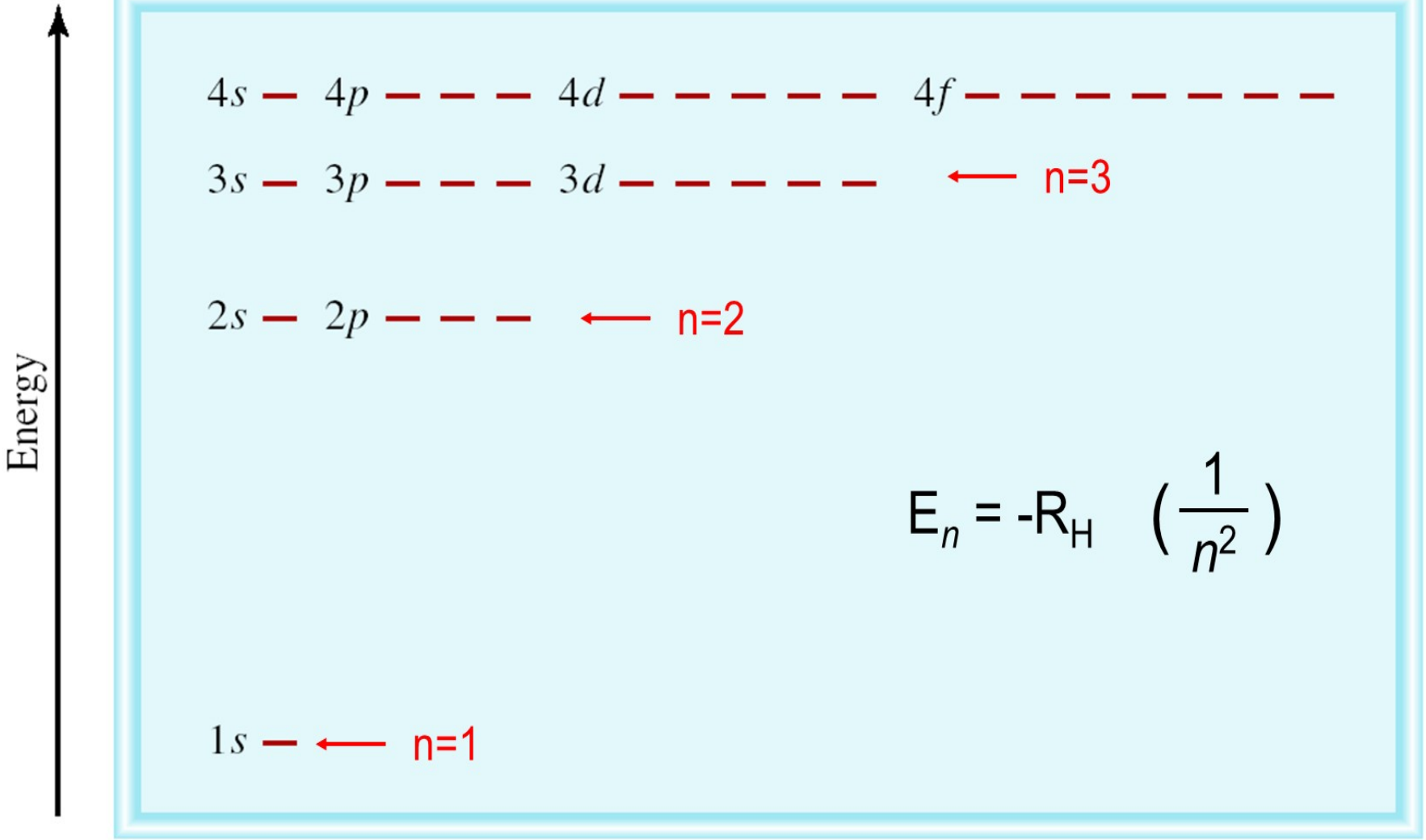
✓ The energy of atomic orbitals increases as the principal quantum number, n, increases.

✓ The energy increases as we move up from 1s to the 2s and then 2p, 3s, and 3p orbitals, showing that the increasing n value has more influence on energy than the increasing l value for small atoms.

✓ However, this pattern does not hold for larger atoms.

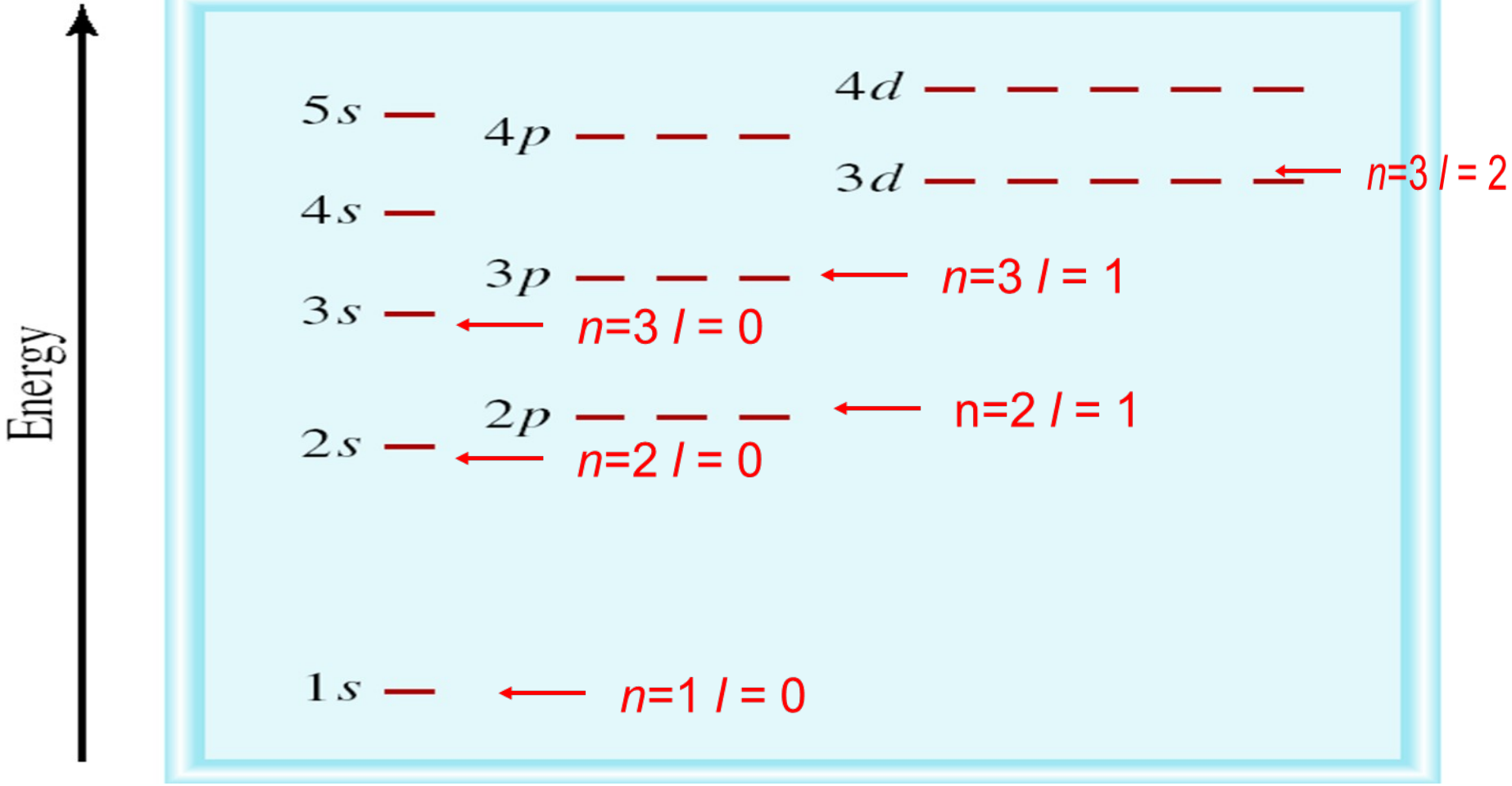
Energy of orbitals in a ***single*** electron atom

Energy only depends on principal quantum number ***n***

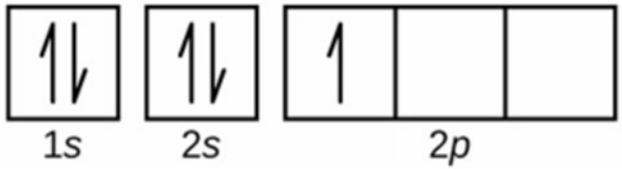


Energy of orbitals in a ***multi***-electron atom

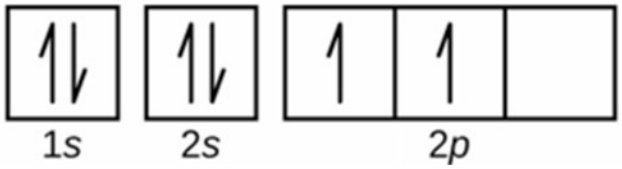
Energy depends on ***n*** and ***l***



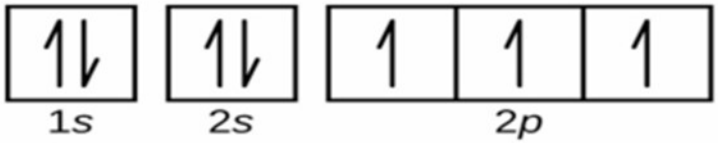
B $1s^2 2s^2 2p^1$



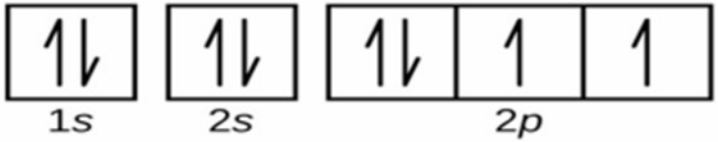
C $1s^2 2s^2 2p^2$



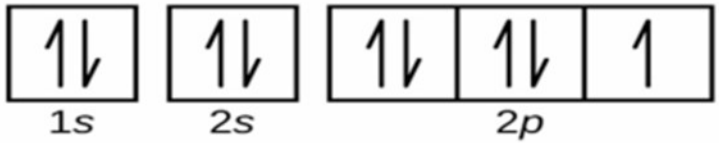
N $1s^2 2s^2 2p^3$



O $1s^2 2s^2 2p^4$



F $1s^2 2s^2 2p^5$



Ne $1s^2 2s^2 2p^6$

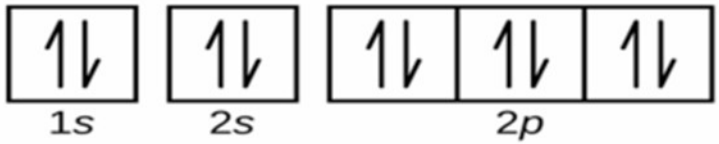


Figure 5.15: A core-abbreviated electron configuration (right) replaces the core electrons with the noble gas symbol whose configuration matches the core electron configuration of the other element.

- What are the possible quantum numbers for the last (outermost) electron in Cl?
- Write the electron configurations of Fe^{2+} , Al^{3+} , P^{3-}

- There are exceptional elements with half-filled or completely filled subshells apparently represent conditions of preferred stability.
- Example: Cr and Cu with 24 and 29 atomic numbers, respectively.

- For a stability issue an electron shifts from the 4s into the 3d orbital to gain the extra stability of a half-filled 3d subshell (in Cr) or a filled 3d subshell (in Cu).
- The electron–electron repulsions experienced by pairing the electrons in the 4s orbital are larger than the gap in energy between the 4s and 3d orbitals.

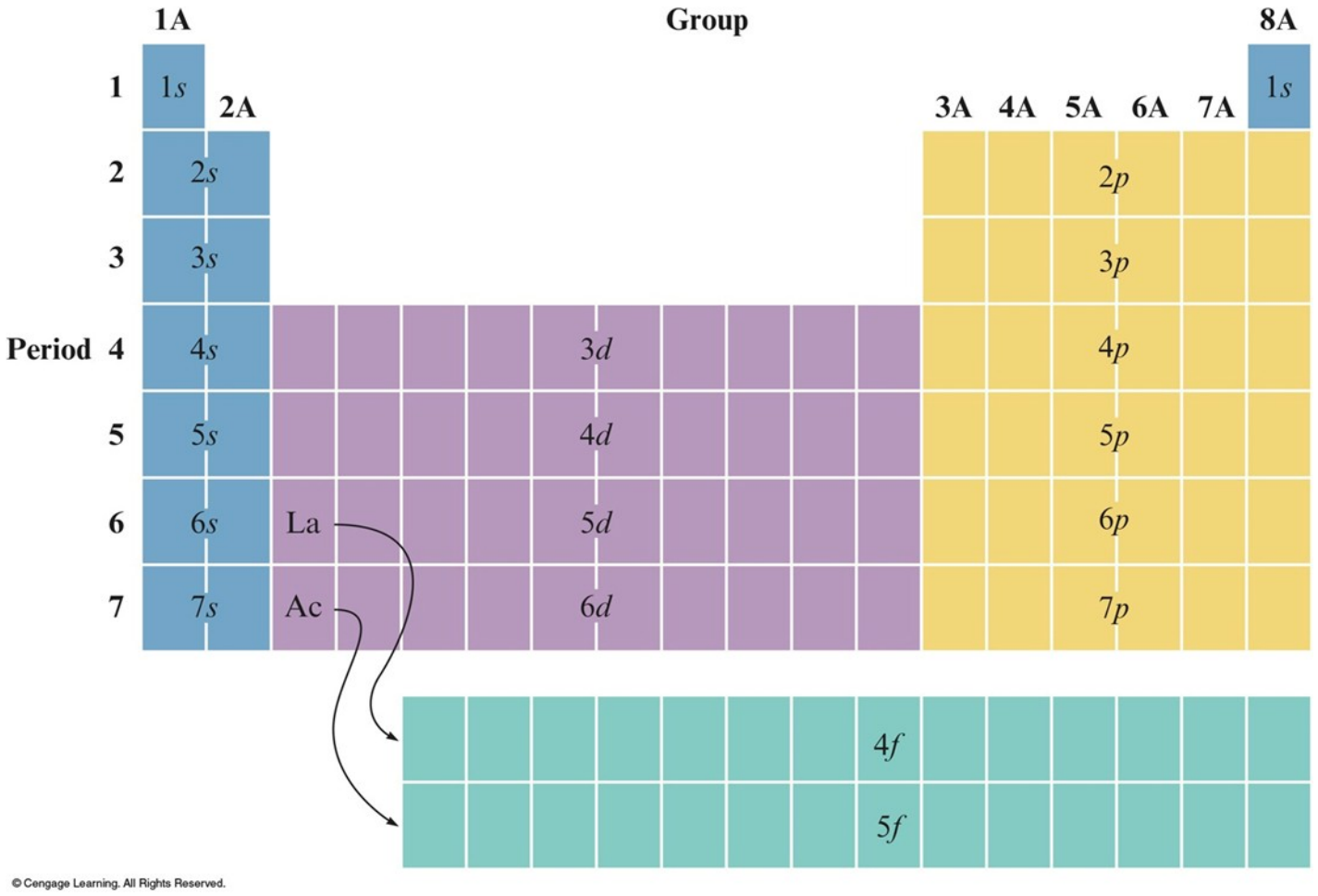
The periodic table

- ✓ The periodic table was created by Dmitri Mendeleev and Lothar Meyer in 1869 and 1870, respectively, to outline their theories.
- ✓ They claimed that the properties of the elements may be modeled as periodic functions of their atomic weights, which is now changed as new elements have been identified.
- ✓ The modern periodic table, where the elements are listed in numerical order based on how many protons (and electrons) they have.
- ✓ Elements with the same valence electrons are grouped into vertical columns inside the table. The group numbers range from 1 to 18. (Arabic numbers).

Cont…

- ✓ The vertical groups of three d-block elements are called triads.
- ✓ Rows in the periodic table are called periods.
- ✓ The lanthanoids include the 14 elements Ce–Lu, and the actinoids include Th–Lr; however, common usage places La with the lanthanoids, and Ac with the actinoids

The Orbitals Being Filled for Elements in Various Parts of the Periodic Table



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Classification of the Elements

- Representative or main group elements: These consist of all s- and p-block elements.

✓ Are those in which the last electron added enters an s or a p orbital in the outermost shell.

✓ The valence electrons for main group elements are those with the highest n level.

✓ E.g Gallium (Ga, 31) has the electron configuration [Ar]$\underline{4s}^2 3d^{10} \underline{4p}^1$, which contains 3 valence electrons (underlined).

- The completely filled d orbitals count as core, not valence, electrons

Transition elements: These are d-block elements.

- There are four series of transitional elements, 3d, 4d, 5d and 6d depending on the energy levels of d-orbitals
- the last electron added enters a d orbital.
- The valence electrons (those added after the last noble gas configuration) in these elements include the ns and (n – 1) d electrons.
- E.g Fe : $1s^2$, $2s^2$,$2p^6$, $3s^2$,$3p^6$, $\underline{4s^2}$, $\underline{3d^6}$

Inner transition elements: These are the f-block elements.

- ✓ the last electron added occupies an f orbital.
- ✓ The valence shells of the inner transition elements consist of the (n – 2)f, the (n – 1)d, and the ns subshells.
- ✓ There are two inner transition series:
  a. The lanthanide series,4f : lanthanide (La) through lutetium (Lu)
  b. The actinide series,5f: actinide (Ac) through lawrencium (Lr).

- Because of their similarities to the other members of the series, Lanthanum and actinium are included and used to name the series, even though they are transition metals with no f electrons.

Periodic Trends

Atomic Radius

- In general as we go across a period from left to right, the atomic radius decreases.
  - Effective nuclear charge increases, therefore the valence electrons are drawn closer to the nucleus, decreasing the size of the atom.
- In general atomic radius increases in going down a group.
- Orbital sizes increase in successive principal quantum levels.

- Which should be the larger atom? Why?

  Na Cl

- Which should be the larger atom? Why?

  Li Cs

Ionization Energy

- Energy required to remove an electron from a gaseous atom or ion.

$X(g) \rightarrow X^+(g) + e^-$

$Mg \rightarrow Mg^+ + e^-$ $I_1 = 735$ kJ/mol (1st IE)

$Mg^+ \rightarrow Mg^{2+} + e^-$ $I_2 = 1445$ kJ/mol (2nd IE)

$Mg^{2+} \rightarrow Mg^{3+} + e^-$ $I_3 = 7730$ kJ/mol *(3rd IE)

*Core electrons are bound much more tightly than valence electrons.

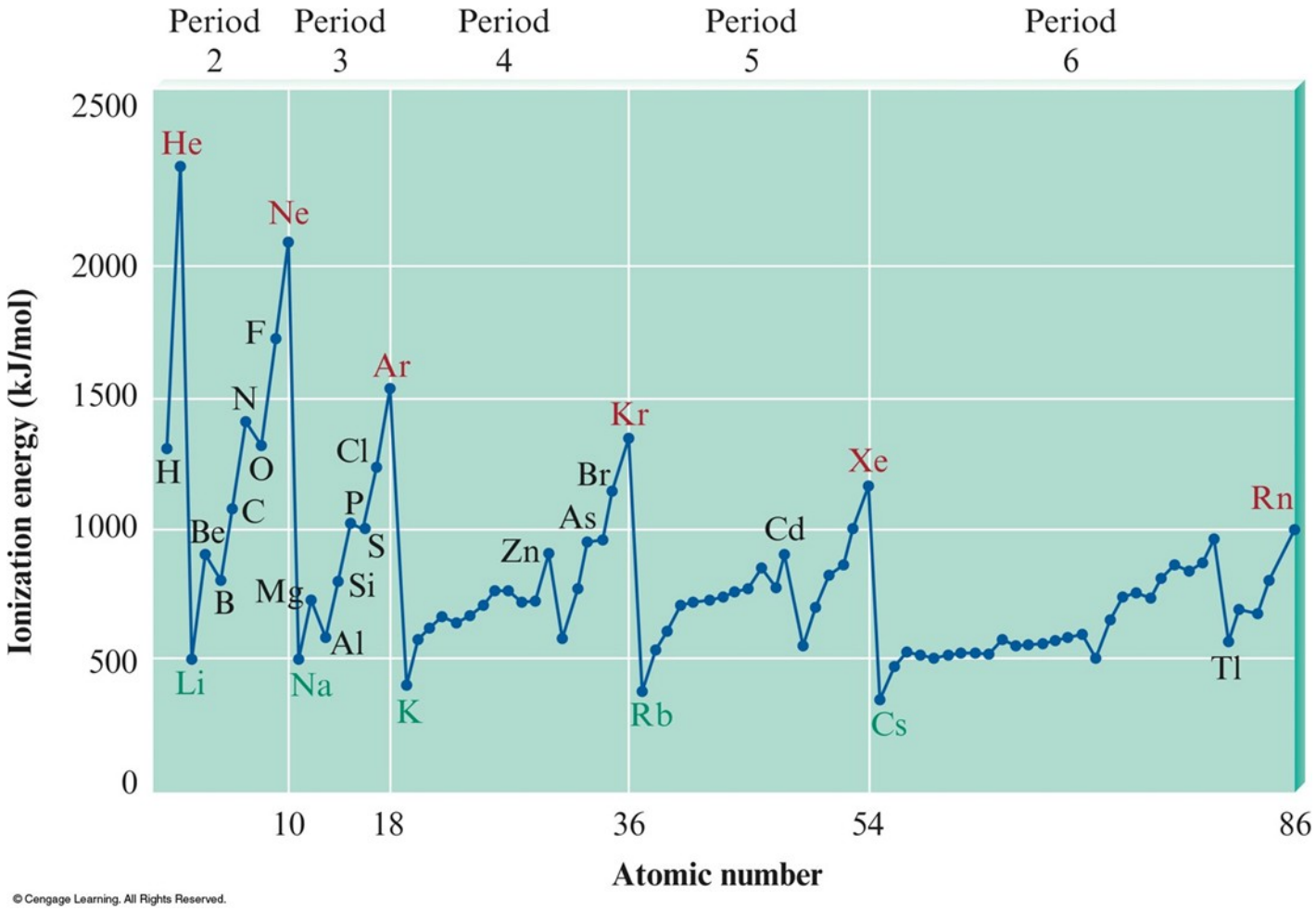
- In general, as we go across a period from left to right, the first ionization energy increases.
- Why?
- Electrons added in the same principal quantum level do not completely shield the increasing nuclear charge caused by the added protons.
- Electrons in the same principal quantum level are generally more strongly bound from left to right on the periodic table.
- In general, as we go down a group from top to bottom, the first ionization energy decreases.
- Why?
- The electrons being removed are, on average, farther from the nucleus.

- As we move across a period, Z_{eff} generally increases so atomic radii become smaller.
- As a result, the attraction between the nucleus and the outer electrons increases, so an electron becomes more difficult to remove.
- In general, ionization energy increases across a period; it is easier to remove an electron from an alkali metal than from a noble gas.
- Explain why the graph of ionization energy versus atomic number (across a row) is not linear.

  electron repulsions
- Where are the exceptions?

  some include from Be to B and N to O

The Values of First Ionization Energy for the Elements in the First Six Periods



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Electron Affinity

- Energy change associated with the addition of an electron to a gaseous atom.

$X(g) + e^- \rightarrow X^-(g)$

- In general as we go across a period from left to right, the electron affinities become more negative.
- In general electron affinity becomes more positive in going down a group.

- Which atom would require more energy to remove an electron? Why?

  Na Cl

- Which atom would require more energy to remove an electron? Why?

  Li Cs

Electronegativity

- Electronegativity is the ability of an atom to attract electrons when the atom is part of a compound.
- The electronegativity scale was devised by Linus Pauling.
- Atoms of the elements in the upper right of the periodic table (small, non-metal atoms) attract bonding electrons most strongly.

- The most non-metallic and hence most electronegative element, fluorine, is assigned a value of 4.0.
- all other electronegativities are assigned relative to Fluorine.
- Fluorine attracts electrons better than any other element.
- Electronegativities generally increase from left to right across a period. This is due to an increase in nuclear charge.
- Within a group, electronegativity decreases from top to bottom.

1. How does electronegativity differ from electron affinity?
2. Why are the electronegativity values of metals generally low?