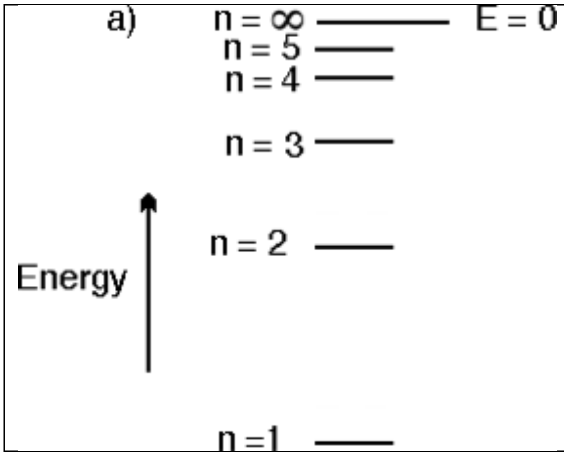
Atomic Structure

1. For a hydrogen atom which electron transition requires the largest amount of energy?

A. n = 4 to n = 10
B. n = 3 to n = 2
C. n = 3 to n = 4
D. n = 1 to n = 3
E. n = 2 to n = 4

	**The separation in energy levels gets smaller and smaller as n increases.**

2. Which of the following do we use to understand quanta of light?

I. The wavelength of a photon is directly proportional to its frequency;
II. The frequency of a photon is directly proportional to its energy;
III. A photon of light is released when an electron undergoes a transition from a low energy level to a level of higher energy.

A. I only
B. II only
C. II and III
D. I and III
E. I, II and III

Important relationships between energy, frequency and wavelength for light (photons) are; **$\lambda \bullet \nu = 3.00 \times 10^{8}$ m s^{-1}** **$E = h\ \nu$** **According to these equations wavelength (λ) and frequency (ν)**	**Also notice in the energy level diagram in Q1, that when an electron moves from a lower energy level to a higher energy level a photon of light must be absorbed, not released.**

are inversely proportional, not directly proportional. Energy is directly proportional to the frequency of a photon.	

3. One of the emission lines in the spectrum of a hydrogen atom has a frequency of $7.40 \times 10^{13}\ s^{-1}$. Calculate the wavelength of this line in nanometers.

A. 4.90×10^{-20} nm
B. 4.90×10^{-11} nm
C. 4.06×10^{-6} nm
D. 4.06 nm
E. 4.06×10^{3} nm

Important relationships between energy, frequency and wavelength for light (photons) are;
$\boldsymbol{\lambda \bullet \nu = 3.00 \times 10^{8}\ m\ s^{-1}}$
$\boldsymbol{\lambda \bullet 7.40 \times 10^{13}\ s^{-1} = 3.00 \times 10^{8}\ m\ s^{-1}}$
$\boldsymbol{\lambda \bullet 7.40 \times 10^{13}\ s^{-1} = \frac{3.00 \times 10^{8}\ m\ s^{-1}}{7.40 \times 10^{13}\ s^{-1}} = 4.06 \times 10^{-6}\ m}$
$\boldsymbol{4.06 \times 10^{-6}\ m \left(\frac{1.00 \times 10^{9}\ nm}{1\ m}\right) = 4.06 \times 10^{3}\ nm}$

4. Which of the following is the best explanation for why the 3rd ionization energy in calcium is significantly greater compared to the first ionization energy in calcium.

A. for a given element each successive ionization energy is significantly larger;
B. the third electron is removed from n = 3 shell where the effective nuclear charge is six times higher compared to the effective nuclear charge on an electron in the n = 4 shell;
C. since the inner core electrons in each shell form a spherical shape, removing an electron will always be significantly greater compared to removing electrons from non-spherical orbitals;
D. electron-electron repulsions decrease for each successive inner shell of electrons, lower electron-electron repulsions means higher ionization energy;
E. in calcium since the 3d orbital's are empty, the first ionization energy is significantly smaller than usual, making the third ionization energy appears significantly larger.

The electron configuration for Calcium is $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$
The first ionization energy is the energy required to remove the electron in the outer most shell, in this case the n = 4 shell. The second ionization energy also removes an electron from the n = 4 shell. The third ionization energy must remove an electron from the n = 3 shell, since the first two ionizations removed electrons from the n = 4 shell.
Z^*_{eff}(for a particular electron) = Z (the # of protons) - #inner core electrons
For an electron (both) in the n = 4 shell
$Z^*_{eff}(4s) = 20 – 18 = +2$
For an electron in the n = 3 shell
Z^*_{eff}(3s or 3p) = 20 – 10 = +10

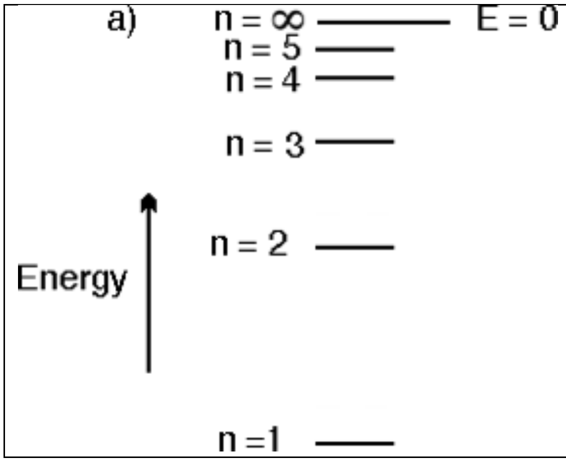
5. The effective nuclear charge on an electron in n = 3 shell in a copper atom is,

A. +1
B. +11
C. +17
D. +19
E. +27

The electron configuration for copper is $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^1 3d^{10}$

Z^*_{eff}(for a particular electron) = Z (the # of protons) - #inner core electrons
For an electron in the n = 3 shell
$Z^*_{eff}(n = 3) = 29 – 10 = +19$

6. When an electron occupies the n = 4 shell in a hydrogen atom
A. the energy required to ionize the atom from the n = 4 level is less than the energy required to ionize the atom when the electron is in the n = 1 level;
B. energy will have to be gained for the electron to get to the n = 1 shell;
C. the electron must go to n = 0 level to be ionized;
D. energy will have to be lost to ionize the electron;
E. the atom must absorb a photon for the electron to move to the n = 3 level or any other level below the n = 4 level.

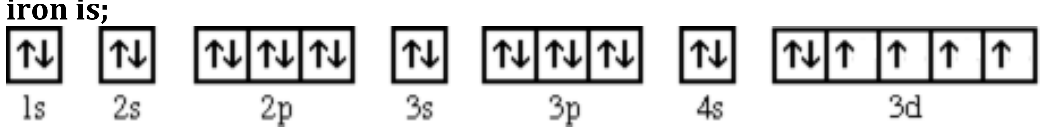


When an electron is in the n = 4 level, it is higher in energy compared to an electron in the n = 1, 2 or 3 levels. To ionize an electron the electron must be moved to the n = infinity level. SO it would require less energy to remove an electron in the n = 4 level compared to the n = 1 level.

7. How many unpaired electron does the element iron, Fe, have?

A) 0
B) 2
C) 4
D) 5
E) 6

The electron configuration for copper is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$
An electron configuration does not tell us about numbers of unpaired electrons, only an orbital diagram does that. The orbital diagram for iron is;

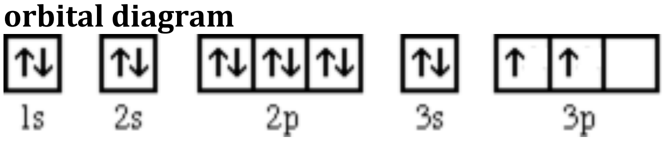


Remember for subshells with degenerate orbitals we must half fill orbitals with electrons before pairing electrons, Hunds rule.

8a. Write the complete electron configuration and indicate the number of unpaired electrons for each of the following species in their ground state,

i. Si __2__ unpaired electrons

electron configuration: **$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$**
orbital diagram

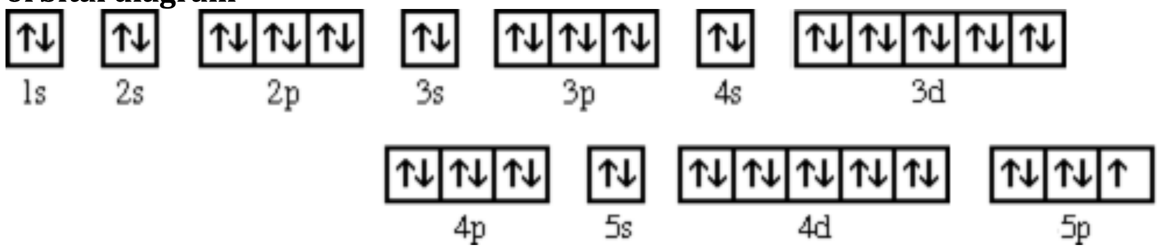


ii. I

___1__ unpaired electrons

electron configuration: **$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^5$**

orbital diagram



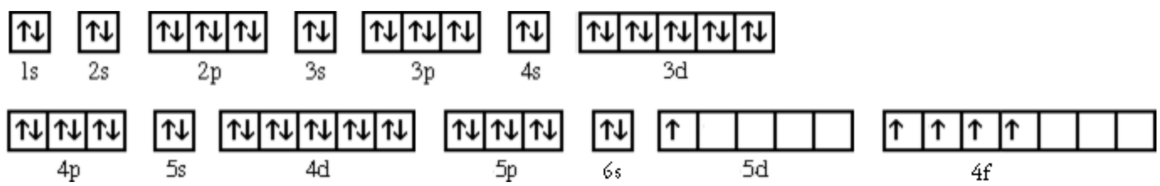
iii. Pm

___5___ unpaired electrons

electron configuration:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 5d^1\ 4f^4$

orbital diagram

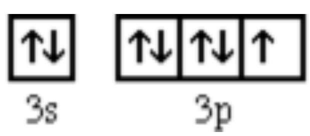


b) Show the orbital diagram (clearly labeled) for the valence electrons for Cl.

electron configuration:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

orbital diagram of the valence electrons ($3s^2\ 3p^5$)



9. Briefly explain each of the following statements.

a) the first ionization energy for sodium is very much smaller than the second ionization for sodium; (8)

electron configuration for sodium is $1s^2\ 2s^2\ 2p^6\ 3s^2$
The first electron is removed from the n = 3 shell. The effective nuclear charge on an electron in the n = 3 shell is +1 (see Q4 above). The second electron ionized comes from the n = 2 shell. The effective nuclear charge on an electron in the n = 2 shell is +9 (see Q4 above). Since the effective nuclear charge experienced by the electron in the n = 2 shell is so much greater compared to the effective nuclear charge experienced by the electron in the n = 3 shell, the energy required to remove an electron from the n = 2 shell is much greater than the energy required to remove an electron from n = 3 shell.

b) the atomic radius for potassium is larger than the atomic radius for Ca; (NOTE: Be sure to include both species in your explanation.) (8)

electron configuration for potassium is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
electron configuration for calcium is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
The valence electrons in both potassium and calcium are in the same shell, the n = 4 shell. The effective nuclear charge experienced by the electrons in the n = 4 shell in potassium is +1, while the effective nuclear charge experienced by the electrons in the n = 4 shell in calcium is +2. The valence electrons in calcium experience a greater attraction to the nucleus compared to the valence electrons in potassium due to their greater effective nuclear charge. The greater attraction results in a smaller atomic radius for calcium compared to potassium.

c) Given the following first ionization energies (in units of kJ mol^{-1}) (9)

Li	Be	B	C	N	O	F
520	899	801	1086	1402	1314	1681
Na	Mg	Al	Si	P	S	Cl
496	?	578	786	1012	1000	1251

Estimate the first ionization energy for magnesium and describe, at least, three different ionization energy trends, from the periodic table, you used to arrive at your estimate.

The first ionization energy for magnesium must be greater than the first ionization energy for sodium. Since both sodium and magnesium have valence electrons in the same shell (the n = 3 shell), the valence electrons in magnesium experience a greater attraction, and therefore require more energy to remove compared to the valence electrons in sodium. The ionization energy for magnesium must be smaller than the first ionization energy of beryllium. The valence electrons in magnesium are in the n = 3 shell, and the valence electrons in beryllium are in the n = 2 shell. The higher the n value the higher the energy of the electron, and the less energy required to remove the electron. Finally the first ionization energy of magnesium must be greater than the first ionization for aluminum. This is because the electron to be removed from aluminum comes from a 3p subshell, which is higher in energy compared to the energy of the 3s subshell and as a result the first ionization energy for aluminum is less than the first ionization energy of magnesium. A reasonable 'guess' for the first ionization for magnesium might be 625 kJ mol^{-1}.

Use these responses for Questions 10 - 12. (A response can be used more than once or not at all.)

A. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
B. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
C. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$
D. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3d^3$
E. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

10. a Group II metal; **E**

11. a transition metal ion. **C**

12. a noble gas. **B**

A Group II metal will have valence electrons in an ns subshell. So choice E, is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$, the electron configuration for Ca, is the best choice for Q10.
A transition metal ion will contain electrons in a d subshell, and if the charge is 2+ or higher, will not have any (n – 1)s electrons. So the best choice for Q11 is C, $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$, the electron configuration for Cu^{2+}.
A noble gas will have a filled shell. So the best choice for Q12 is B, $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$, the electron configuration for Ar.

13. The wavelength of a photo of light emitted by fireworks is 650 nm. Calculate the frequency of this light.

A) $4.31 \times 10^{-31}\ s^{-1}$
B) $4.62 \times 10^{5}\ s^{-1}$
C) $1.95 \times 10^{11}\ s^{-1}$
D) $2.17 \times 10^{-15}\ s^{-1}$
E) $4.62 \times 10^{14}\ s^{-1}$

Important relationships between energy, frequency and wavelength for light (photons) are;

$$\lambda \bullet \nu = 3.00 \times 10^8\ m\ s^{-1}$$

$$650\ nm \left(\frac{1\ m}{1.00 \times 10^9\ nm}\right) = 6.50 \times 10^{-7}\ m$$

$$6.50 \times 10^{-7}\ m \bullet \nu = 3.00 \times 10^8\ m\ s^{-1}$$

$$6.50 \times 10^{-7}\ m \bullet \nu = \frac{3.00 \times 10^8\ m\ s^{-1}}{6.50 \times 10^{-7}\ m} = 4.62 \times 10^{14}\ s^{-1}$$

14. The energy of a photon of light released from a hydrogen atom is 2.04×10^{-18} J. If the electron is originally in the n = 4 level, what level does the electron occupy after emitting the photon?

A) n = 1
B) n = 2
C) n = 3
D) n = 4
E) n = 5

The relationship between the energy absorbed or released by an atom with the electron changes energy levels is given by,

$$\Delta E = -2.18 \times 10^{-18}\ \mathrm{J}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

The energy of the photon is 2.04×10^{-18} J, and since the energy is released a negative sign must precede the energy in the equation;

$$-2.04 \times 10^{-18}\ \mathrm{J} = -2.18 \times 10^{-18}\ \mathrm{J}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$ **substituting n = 4 for the initial energy level and rearranging we have,**

$$\frac{-2.04 \times 10^{-18}\ \mathrm{J}}{-2.18 \times 10^{-18}\ \mathrm{J}} = \left(\frac{1}{n_f^2} - \frac{1}{4_i^2}\right)$$

$$0.936 = \left(\frac{1}{n_f^2} - 0.0625\right)$$

$$0.9985 = \left(\frac{1}{n_f^2}\right)$$

$n^2 = 1.001$ so n = 1

Use the following choices to answer Q15 and Q16. (A response can be used more than once or not at all.)

A. O
B. Mg
C. Si
D. P
E. Cl

15. Has valence electrons that experience an effective nuclear charge of +6.

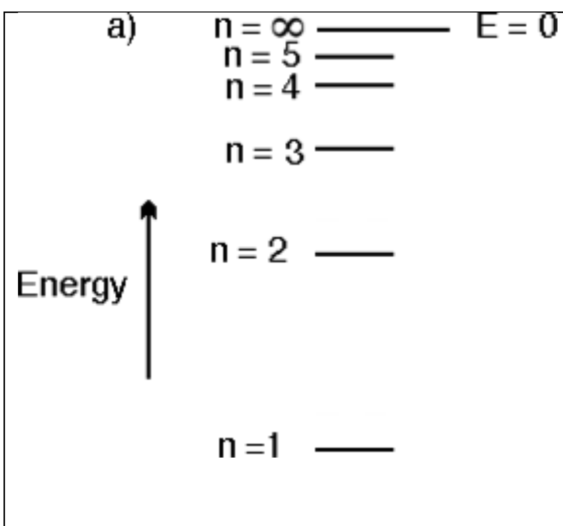
16. Has electrons in the second level that experience an effective nuclear charge of +13.

An atom with a +6 effective nuclear charge is in Group VI. So choice A, oxygen is the correct choice. The electron configuration for O is $1s^2\ 2s^2\ 2p^4$, so two electron in the n = 1 shell shield the nucleus (Z = +8) so the valence electrons experience a +6 effective nuclear charge.

An atom with electrons in the n = 2 shell that experience an effective nuclear charge of +13, then the n = 2 electrons are shielded by two electrons in the n = 1 shell, so the total nuclear charge must 15, which is phosphorus. So the choice is D. The electron configuration for P is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$, so two electron in the n = 1 shell shield the nucleus (Z = +15) so the electrons in the n = 2 level experience a +13 effective nuclear charge.

17. When an electron occupies the n = 6 shell in a hydrogen atom

A. energy will have to be lost to ionize the electron;
B. energy will have to be gained for the electron to get to the n = 1 shell;
C. the electron must go to n = 0 level to be ionized;
D. **the energy required to ionize the atom from the n = 6 level is less than the energy required to ionize the atom when the electron is in the n = 1 level;**
E. the atom must absorb a photon for the electron to move to the n = 5 level or any other level below the n = 6 level.



This question is identical to Q6 above. When an electron is in the n = 6 level, it is higher in energy compared to an electron in the n = 1, 2, 3, 4 or 5 levels. To ionize an electron the electron must be moved to the n = infinity level. So it would require less energy to remove an electron in the n = 6 level compared to the n = 1 level.