SOLUTIONS - CHAPTER 3 Problems

1) The recommended intake of food calories per day is 1940 food calories/day for adult women and 2550 food calories/day for adult men. Express these recommendations in units of kJ/day. Note that 1 food calorie = 1000. calorie (exact) and 1 calorie = 4.184 J (exact).

For women, \[ \text{kJ} = \frac{1940 \text{ food cal}}{1000 \text{ cal}} \times \frac{4.184 \text{ J}}{1 \text{ kJ}} = 8120 \text{ kJ/day} \]

For men, \[ \text{kJ} = \frac{2550 \text{ food cal}}{1000 \text{ cal}} \times \frac{4.184 \text{ J}}{1 \text{ kJ}} = 10700 \text{ kJ/day} \]

2) An electron has a kinetic energy \( E_K = 4.55 \times 10^{-22} \text{ J} \). What is the velocity of the electron?

\[ E_K = \frac{mv^2}{2} \quad \text{and so} \quad v = \left(\frac{2E_K}{m}\right)^{1/2} \]

and so \( v = \left[2(4.55 \times 10^{-22} \text{ J})/(9.109 \times 10^{-31} \text{ kg})\right]^{1/2} = 3.16 \times 10^4 \text{ m/s} \quad (31.6 \text{ km/s}) \)

We found the value for the mass of the electron from the back cover of the textbook. It would be given to you on an exam. Note that if we put everything in the formula for \( v \) in MKS units, we know the answer will also have MKS units.

3) (Burdge, 3.11) What is a wave? Using a diagram, define the following terms associated with waves: \textbf{wavelength}, \textbf{frequency}, \textbf{amplitude}.

Wave – An oscillation that transfers energy (such as water waves or light).

\textbf{wavelength} - The distance between successive peaks (or troughs) in a wave. MKS unit is meters.

\textbf{frequency} - The number of wavelengths passing a fixed point per unit time. MKS unit is \text{s}^{-1}, but the term Hertz is also used for frequency (1 Hz = 1 cycle/s = 1 s\(^{-1}\)).

\textbf{amplitude} - The height of a wave, which is related to the intensity.
4) A low pressure mercury lamp emits light with wavelength $\lambda = 253.7 \text{ nm}$.
   a) What is the frequency of the light emitted by the lamp?

   \[ c = \lambda \nu, \text{ and so } \nu = \frac{c}{\lambda}. \]

   So \[ \nu = \frac{2.998 \times 10^8 \text{ m/s}}{253.7 \times 10^{-9} \text{ m}} = 1.182 \times 10^{15} \text{ s}^{-1} \]

   b) What is the energy of one photon of light from the lamp?

   \[ E = h\nu = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{253.7 \times 10^{-9} \text{ nm}} = 7.830 \times 10^{-19} \text{ J} \]

   c) What is the energy of one mole of photons emitted by the lamp? Give your answer in units of kJ/mol.

   \[ E \text{ (per mol)} = \frac{7.830 \times 10^{-19} \text{ J}}{6.022 \times 10^{23} \text{ photon}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 471.5 \text{ kJ/mol} \]

5) (Burdge, 3.19) How many minutes would it take a radio wave to travel from the planet Venus to Earth, when they are separated by their average distance of 28. million miles? How long would it take an infrared wave to travel the same distance?

   \[ d \text{ (in km)} = 28. \times 10^6 \text{ mile} \times \frac{1.602 \text{ km}}{1 \text{ mile}} = 44.9 \times 10^6 \text{ km} = 44.9 \times 10^9 \text{ m} \]

   \[ t = \frac{d}{v} = \frac{44.9 \times 10^9 \text{ m}}{2.998 \times 10^8 \text{ m/s}} = 149.6 \text{ s} = 2.49 \text{ minutes} \]

Since the speed of light in a vacuum is independent of wavelength, it would take the same time for an infrared wave to travel the same distance.

6) (Burdge, 3.26) The blue color of the sky results from the scattering of sunlight from molecules of air. The blue light has a frequency of about $7.5 \times 10^{14} \text{ Hz}$.

   a) Calculate the wavelength (in nm) associated with this radiation.

   \[ c = \lambda \nu, \text{ and so } \lambda = \frac{c}{\nu} \]

   So \[ \lambda = \frac{2.998 \times 10^8 \text{ m/s}}{7.5 \times 10^{14} \text{ s}^{-1}} = 4.0 \times 10^{-7} \text{ m} = 400 \text{ nm} \]

   b) Calculate the energy (in Joule) of a single photon associated with this frequency of light.

   \[ E = h\nu = (6.626 \times 10^{-34} \text{ J}\cdot\text{s})(7.5 \times 10^{14} \text{ s}^{-1}) = 5.0 \times 10^{-19} \text{ J} \]
7) A laser emits photons with an energy 248. kJ/mol. What is the wavelength of light (in nm) emitted by the laser? In what region of the electromagnetic spectrum is this radiation found?

\[ E_{\text{per photon}} = \frac{248 \times 10^3 \text{ J}}{\text{mol}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ photon}} = 4.12 \times 10^{-19} \text{ J/photon} \]

\[ E = \frac{hc}{\lambda}, \text{ and so } \lambda = \frac{hc}{E} \]

So \[ \lambda = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{4.12 \times 10^{-19} \text{ J}} = 4.82 \times 10^{-7} \text{ m} \times \frac{1 \text{ nm}}{10^{-9} \text{ m}} = 482. \text{ nm} \]

This is in the visible region of the spectrum. (visible region = 400 - 700 nm)

8) Calculate the wavelength (in nm) of a photon emitted by a hydrogen atom when its electron drops from the \( n = 7 \) state to the \( n = 2 \) state.

\[ \frac{1}{\lambda} = R_H \{ (1/n_f^2) - (1/n_i^2) \} = (0.01097 \text{ nm}^{-1}) \{ (1/2^2) - (1/7^2) \} = 0.002509 \text{ nm}^{-1} \]

So \[ \lambda = \frac{1}{(0.002509 \text{ nm}^{-1})} = 397.0 \text{ nm} \]

9) (Burdge, 3.44) Consider the following energy levels of a hypothetical atom:

\[ E_4 = - 1.0 \times 10^{-19} \text{ J} \]
\[ E_3 = - 5.0 \times 10^{-19} \text{ J} \]
\[ E_2 = - 10. \times 10^{-19} \text{ J} \]
\[ E_1 = - 15. \times 10^{-19} \text{ J} \]

An energy level diagram is given below to help picture the processes taking place

```
   E_4
   |   E_3
   |   E_2
   |   E_1
   energy
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a) What is the wavelength of the photon needed to excite an electron from \( E_1 \) to \( E_4 \)?

\[ \Delta E = \frac{hc}{\lambda} = E_4 - E_1 = ( - 1.0 \times 10^{-19} \text{ J}) - ( - 15. \times 10^{-19} \text{ J}) = 14. \times 10^{-19} \text{ J} \]

\[ \lambda = \frac{hc}{\Delta E} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.998 \times 10^8 \text{ m/s})}{14. \times 10^{-19} \text{ J}} = 1.4 \times 10^{-7} \text{ m} = 140 \text{ nm} \]
b) What is the energy (in Joule) a photon must have to excite an electron from $E_2$ to $E_3$?

$$\Delta E = E_3 - E_2 = (-5.0 \times 10^{-19} \text{ J}) - (-10. \times 10^{-19} \text{ J}) = 5 \times 10^{-19} \text{ J}$$

c) When an electron drops from the $E_3$ level to the $E_1$ level, the atom is said to undergo emission of light. Calculate the wavelength (in nm) of the photon emitted in this process.

$$\Delta E = \frac{hc}{\lambda} = E_3 - E_1 = (-5.0 \times 10^{-19} \text{ J}) - (-15. \times 10^{-19} \text{ J}) = 10. \times 10^{-19} \text{ J}$$

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})(2.998 \times 10^8 \text{ m/s})}{2.0 \times 10^{-7} \text{ m}} = 200 \text{ nm}$$

10) (Burdge, 3.57) Thermal neutrons are neutrons that move at speeds comparable to those of air molecules at room temperature. These neutrons are most effective in initiating a nuclear chain reaction in $^{235}\text{U}$. Calculate the de Broglie wavelength (in nm) associated with a beam of neutrons moving at $7.00 \times 10^2 \text{ m/s}$. The mass of a neutron is $1.675 \times 10^{-27} \text{ kg}$.

$$\lambda_{DB} = \frac{h}{mv} = \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})}{(1.675 \times 10^{-27} \text{ kg})(7.00 \times 10^2 \text{ m/s})} = 5.65 \times 10^{-10} \text{ m} = 0.565 \text{ nm}$$

11) (Burdge, 3.66) How is the concept of electron density used to describe the position of an electron in the quantum mechanical treatment of the atom?

According to quantum mechanics we cannot assign a specific location for an electron in an atom (or ion or molecule). Instead we use the concept of electron density, which represents the probability of finding the electron in a particular location in space.

12) The speed of an electron is known to within $60.0 \text{ m/s}$. What is the minimum uncertainty in the position of the electron? Give your answer in units of nm.

$$(\Delta x)(\Delta p) \geq \frac{\hbar}{4\pi}$$, and so $$(\Delta x) \geq \frac{\hbar}{4\pi(\Delta p)}$$

$$\Delta p = m(\Delta v) = (9.109 \times 10^{-31} \text{ kg})(60.0 \text{ m/s}) = 5.465 \times 10^{-29} \text{ kg} \cdot \text{m/s}$$

$$\Delta x \geq \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})}{4\pi(5.465 \times 10^{-29} \text{ kg} \cdot \text{m/s})} = 9.65 \times 10^{-7} \text{ m} = 965. \text{ nm}$$
13) (Burdge, 3.72) Describe the four quantum numbers used to characterize an electron in an atom.

The four quantum numbers (and their restrictions) are as follows:

n (principal quantum number).  \( n = 1, 2, 3, \ldots \)  For the hydrogen atom this determines the energy of the electron.  In addition, the larger the value for \( n \) the greater the average distance between the electron and the nucleus, and the larger the orbital.

\( \ell \) (angular momentum quantum number).  \( \ell = 0, 1, 2, \ldots, (n-1) \).  Determines the shape of the orbitals.  For multielectron atoms the energy of the electron is determined by both \( n \) and \( \ell \).

\( m_\ell \) (magnetic quantum number).  \( m_\ell = 0, \pm1, \pm2, \ldots, \pm\ell \).  Determines the orientation of the orbital.  Note that for a particular value of \( \ell \) there are \( 2\ell + 1 \) distinct orbitals.

\( m_s \) (spin quantum number).  \( m_s = \pm \frac{1}{2} \).  Determines the orientation of the electron spin, either “spin up” (\( m_s = \frac{1}{2} \)) or “spin down” (\( m_s = -\frac{1}{2} \)).

14) (Burdge, 3.76) An electron in an atom is in the \( n = 3 \) quantum level.  List the possible values of \( \ell \) and \( m_\ell \) that it can have.

If \( n = 3 \), the possible values for \( \ell \) and \( m_\ell \) are:

\( \ell = 2; m_\ell = -2, -1, 0, 1, 2 \)
\( \ell = 1; m_\ell = -1, 0, 1 \)
\( \ell = 0; m_\ell = 0 \)

15) An electron in an atom has \( n = 4, \ell = 1 \).  What are the possible values for \( m_\ell \) and \( m_s \) for the electron?

Possible values for \( m_\ell \) are \( m_\ell = 1, 0, -1 \)

Possible values for \( m_s \) are \( m_s = \frac{1}{2} \) or \( -\frac{1}{2} \)

16) (Burdge, 3.84) Give the values for the four quantum numbers of an electron in the following orbitals:

a) 3s  \( n = 3, \ell = 0, m_\ell = 0, m_s = +\frac{1}{2} \) or \(-\frac{1}{2} \)

b) 4p  \( n = 4, \ell = 1, m_\ell = 1, 0, -1, m_s = +\frac{1}{2} \) or \(-\frac{1}{2} \)

c) 3d  \( n = 3, \ell = 2, m_\ell = 2, 1, 0, -1, -2, m_s = +\frac{1}{2} \) or \(-\frac{1}{2} \)

17) (Burdge, 3.86) What is the difference between a 2p\(_x\) and a 2p\(_y\) orbital?

Both orbitals have the same shape.  A 2p\(_x\) orbital is aligned with the x axis of the atom, and a 2p\(_y\) orbital is aligned with the y axis of the atom.
18) (Burdge, 3.87) Why do the 3s, 3p, and 3d orbitals have the same energy in a hydrogen atom but different energies in a many-electron atom?

In a hydrogen atom there is only one electron, and so no interaction of electrons with other electrons in the atom. In that case the energy of the electron only depends on the value for \( n \).

In a multielectron atom different electrons in different orbitals push each other away, raising the energy. For a particular value for \( n \), the smaller the value for \( \ell \) the lower the energy for the orbital.

19) Which of the following pairs of orbitals is lower in energy in a multielectron atom?
   a) 3s or 4s  3s is lower in energy
   b) 3d or 4s  4s is lower in energy
   c) 4p or 5d  4p is lower in energy

20) State the Aufbau principle and the Pauli principle. Explain the role these play in determining the organization of the periodic table.

The Aufbau principle applies to the ground (lowest energy) state of an atom. It says that electrons will add to an atom by going into the lowest energy orbital that has space to accommodate them.

The Pauli principle states that two electrons in the same atom cannot have the same set of quantum numbers.

The combination of these principles determines the order in which electrons add to an atom. Elements in the periodic table are arranged in groups that have the same configuration of valence electrons, as it is the valence electrons that are most important in determining the chemical properties of an element.

21) What is the difference between diamagnetic and paramagnetic? How are these terms related to Hund’s rules.

An atom is diamagnetic if it has no unpaired electron spins. Such an atom is weakly repelled by a magnetic field. An atom is paramagnetic if it has one or more unpaired electron spin. Paramagnetic atoms are strongly attracted by a magnetic field, with the strength of the attraction depending on the number of unpaired electron spins.

Hund’s rules require that we maximize the number of unpaired electron spins (spin up electrons) in an atom, as long as the Aufbau principle and Pauli principle are also satisfied. Because of this, the majority of atoms are paramagnetic. Only atoms that only have completely filled atomic orbitals (such as He, Be, Ne) will be diamagnetic.
22) (Burdge, 3.106) Portions of the orbital filling diagrams representing ground state electron configurations of certain elements are shown here. Which of them violate the Pauli exclusion principle? Which violate Hund’s rule?

![Diagram](image)

(a), (b), and (c)违反了Pauli原理（两个电子在同一个轨道中有相同的自旋）。

(b)违反了Hund’s规则。我们可以将其中的三个电子“向上自旋”并得到一个较大的总自旋值。

(c)良好，如是。

(d)违反了Hund’s规则。我们可以将一个电子从轨道向下自旋并使其自旋向上，然后将所有五个电子自旋向上。

(e)违反了Hund’s规则。我们可以将电子从第4个轨道中移入空轨道，并使其自旋向上，然后将5个中的6个电子自旋向上。

(f)违反了Pauli原理。两个电子在第三个轨道中自旋向下。

23) (Burdge, 3.116) Write the ground state electron configurations for the following elements (write the configurations using both the long and the shorthand method).

a) B
   1s² 2s² 2p⁴ = [He] 2s² 2p⁴
   Long: 1s² 2s² 2p⁴ 2p⁴
   Shorthand: [He] 2s² 2p⁴

b) V
   1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d³ = [Ar] 4s² 3d³
   Long: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d³ 4s² 3d³
   Shorthand: [Ar] 4s² 3d³

c) C
   1s² 2s² 2p² = [He] 2s² 2p²
   Long: 1s² 2s² 2p² 2p²
   Shorthand: [He] 2s² 2p²

d) As
   1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p³ = [Ar] 4s² 3d¹⁰ 4p³
   Long: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p³ 4s² 3d¹⁰ 4p³
   Shorthand: [Ar] 4s² 3d¹⁰ 4p³

e) I
   1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p⁶ 5s² 4d¹⁰ 5p⁵ = [Kr] 5s² 4d¹⁰ 5p⁵
   Long: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p⁶ 5s² 4d¹⁰ 5p⁵ 5s² 4d¹⁰ 5p⁵
   Shorthand: [Kr] 5s² 4d¹⁰ 5p⁵

f) Au
   1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p⁶ 5s² 4d¹⁰ 5p⁵ 6s¹ 4f¹⁴ 5d¹⁰
   Long: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p⁶ 5s² 4d¹⁰ 5p⁵ 6s¹ 4f¹⁴ 5d¹⁰
   Shorthand: [Xe] 6s¹ 4f¹⁴ 5d¹⁰

The advantage of the shorthand notation for atoms with a large number of electrons is clear.

24) (Burdge, 3.111) Explain why the ground-state electron configuration for Cr and Cu are different than what we might expect.

<table>
<thead>
<tr>
<th>Element</th>
<th>Predicted</th>
<th>Actual</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cr</td>
<td>[Ar] 4s² 3d⁴</td>
<td>[Ar] 4s¹ 3d⁵</td>
</tr>
<tr>
<td>Cu</td>
<td>[Ar] 4s² 3d⁹</td>
<td>[Ar] 4s¹ 3d¹⁰</td>
</tr>
</tbody>
</table>

In both cases the difference is because there is a small lowering of energy if a single electron can be transferred from an s orbital to a d orbital to give a d orbital that is either half-filled (5 electrons) or completely filled (10 electrons).
25) For each of the following atoms give the electron configuration. Also give the number of core electrons, the number of valence electrons, and the total number of unpaired electron spins. Identify each atom as diamagnetic or paramagnetic.

a) N \(1s^2 2s^2 2p^3\) = [He] 2s\(^2\) 2p\(^3\) 2 core electrons, 5 valence electrons 3 unpaired electron spins, paramagnetic

b) Ca \(1s^2 2s^2 2p^6 3s^2 3p^6 4s^2\) = [Ar] 4s\(^2\) 18 core electrons, 2 valence electrons 0 unpaired electron spins, diamagnetic

c) Ti \(1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2\) = [Ar] 4s\(^2\) 3d\(^2\) 20 core electrons, 2 valence electrons 2 unpaired electron spins, paramagnetic

Note the concept of valence electrons is most important for the main group elements.

26) What is the difference between a core electron and a valence electron? Which electrons are most important in determining the chemical properties of an element?

The valence electrons are all of the electrons corresponding to those with the highest number of n (principle quantum number) in an atom. The remaining electrons are classified as core electrons. For example, for a sulfur atom, with electron configuration

\(S\) \(1s^2 2s^2 2p^6 3s^2 3p^4\) = [Ne] 3s\(^2\) 3p\(^4\)

the largest value for the quantum number n is n = 3, and so the 3s and 3p electrons are classified as valence electrons (6 total). The remaining electrons, with n = 1 or 2, are core electrons (10 total).

The valence electrons are the more important electrons in determining the chemical properties of an element.

27) Write the ground state electron configurations for the first four alkali earth metals. Write the configurations using both the long and the shorthand method. Based on your answer, explain why the alkaline earth metals are expected to have similar chemical properties.

The alkaline earth metals are the elements in Group 2A of the periodic table. The electron configuration for the first four alkaline earth metals are as follows:

Be \(1s^2 2s^2 = [\text{He}] 2s^2\)
Mg \(1s^2 2s^2 2p^6 3s^2 = [\text{Ne}] 3s^2\)
Ca \(1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 = [\text{Ar}] 4s^2\)
Sr \(1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 = [\text{Kr}] 5s^2\)

All of the alkaline earth metals have a noble gas + ns\(^2\) electron configuration, and so have two valence electrons. Elements with the same number of valence electrons are expected to have similar chemical properties.
28) (Burdge, 3.142) Ionization energy is the minimum energy required to remove an electron from an atom. It is usually expressed in units of kJ/mol, that is, the energy in kilojoules required to remove one mole of electrons from one mole of atoms.

a) Calculate the ionization energy for a hydrogen atom.

When an electron and proton are bound together in a hydrogen atom the energy of the system is negative. A positive energy corresponds to a free electron that is no longer bound to the proton. Therefore the energy required to remove one electron in the n = 1 state from one hydrogen atom is just the energy that must be added to an electron in the n = 1 state to make its final energy greater than zero. The minimum energy to do this is

\[ \Delta E = 0 - E_{n=1} = - ( -2.18 \times 10^{-18} \text{ J}) = 2.18 \times 10^{-18} \text{ J} \]

The ionization energy, \( E_I \), is the energy required to ionize one mole (Avogadro's number) of hydrogen atoms with electrons initially in the n = 1 state.

\[ E_I = N_A \Delta E = (6.022 \times 10^{23} \text{ mol}^{-1})(2.18 \times 10^{-18} \text{ J}) = 1.31 \times 10^6 \text{ J/mol} = 1310 \text{ kJ/mol} \]

b) Repeat the calculation, assuming that the electrons are being removed from the n = 2 state instead of from the ground (n = 1) state.

If the hydrogen atoms are initially in the n = 2 state, then for one atom

\[ \Delta E = 0 - E_{n=2} = - ( -2.18 \times 10^{-18} \text{ J}) = 5.45 \times 10^{-19} \text{ J} \]

and to ionize one mole of atoms requires an energy

\[ E_I = N_A \Delta E = (6.022 \times 10^{23} \text{ mol}^{-1})(5.45 \times 10^{-19} \text{ J}) = 3.28 \times 10^5 \text{ J/mol} = 328. \text{ kJ/mol} \]