Student Name	Room Assignment
A-Byington	Stewart Biology S1/4
Byland-Coates McCall	Otto Maass 10
Cochin-Egan	Stewart Biology N2/2
El Abbas-Inoue	F.D.A. Auditorium
Isbester-Muddiman	Leacock 132
Mulherin-Panagiotou	Otto Maass 217
Panaite-Rozenberg	Leacock 26
Rudd-Stephenson	Rutherford Physics 112
Stetson-Z	Otto Maass 112

If there is a topic in the text that was *not* covered in class, it will *not* be on the exam. Also, please note that you are also *not* responsible for the "side stories". Do, however, make sure to *memorize* the important equations that were posted earlier in the Selected Slides section on webCT.

The midterm consists of 40 multiple-choice questions, which will be machine-scored. Bring a scientific calculator (with its memory, if any, cleared; you will need to prove this to the invigilator), pencils for the computer card and, most importantly, your McGill ID card with your student number.

Answers to Chapter 8 Problems and Other Study Points

Scanned answers to the problems in Chapter 8 and those for problems 9-36 in Chapter 9 have now been posted.

Exercises #24-30 in Chapter 9

These exercises involve calculations of ionization energies. You are welcome to do these, but there will be no ionization calculations on the midterm.
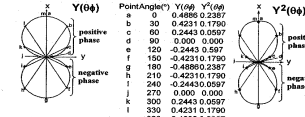

Photoelectric Effect

The photons of a light beam have a characteristic energy determined by the frequency of the light ($E = hv$). In the photoelectric effect, if an electron absorbs the energy of one photon and has more energy than the work function ($E = hv_0$), it is ejected from the material. If the photon energy is too low, the electron is unable to escape the surface of the material. Increasing the intensity of the light beam when $E > hv_0$ increases the number of photons in the light beam, and thus increases the number of electrons emitted without increasing the energy that each electron possesses. The energy of the emitted electrons does not depend on the intensity of the incoming light, but only on the energy of the individual photons.

Electrons can absorb energy from photons when irradiated, but they follow an "all or nothing" principle. All of the energy from one photon must be absorbed and used to liberate one electron from atomic binding, or the energy is re-emitted. If the photon energy is absorbed, some of the energy liberates the electron from the atom, and the rest contributes to the electron's kinetic energy as a free particle.

Chapter 8: Electrons in Atoms

Note value of wavefunction can be positive or negative

65./67. The $2p_z$ orbital $Y(\theta\phi) = \sqrt{\frac{3}{4\pi}}\sin\theta\cos\phi$; Note: in the xy plane $\theta = 90°$ and $\sin\theta = 1$

Plotting in the xy plane requires that we vary only ϕ

PointAngle(°)	$Y(\theta\phi)$	$Y^2(\theta\phi)$
a	0	0.4886	0.2387
b	30	0.4231	0.1790
c	60	0.2443	0.0597
d	90	0.000	0.000
e	120	-0.2443	0.597
f	150	-0.42310	1790
g	180	-0.48860	2387
h	210	-0.42310	1790
i	240	-0.24430	0597
j	270	0.000	0.000
k	300	0.2443	0.0597
l	330	0.4231	0.1790
m	360	0.4886	0.2387

66./68. The $2p_y$ orbital $Y(\theta\phi) = \sqrt{\frac{3}{4\pi}}\sin\theta\sin\phi$, however, in the xy plane $\theta = 90°$ and $\sin\theta = 1$

Plotting in the xy plane requires that we vary only ϕ

PointAngle(°)	$Y(\theta\phi)$	$Y^2(\theta\phi)$
a	0	0.0000	0.0000
b	30	0.2443	0.0597
c	60	0.4231	0.1790
d	90	0.4886	0.2387
e	120	0.4231	0.1790
f	150	0.2443	0.0597
g	180	0.0000	0.0000
h	210	-0.24430	0597
i	240	-0.42310	1790
j	270	-0.48860	2387
k	300	-0.42310	1790
l	330	-0.24430	0597
m	360	0.000	0.000

220
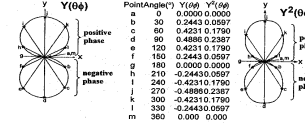
Which of the following particles has the longest wavelength associated with it?

1. An electron (mass 9.109×10^{-31} kg) moving at one-fifth the speed of light

2. A proton (mass 1.673×10^{-27} kg) moving at one-fifth the speed of light

3. An electron (mass 9.109×10^{-31} kg) moving at one-tenth the speed of light.

4. A proton (mass 1.673×10^{-27} kg) moving at one-tenth the speed of light

---

Answer

This question depends on the de Broglie relationship: $\lambda = h/mu$

1. $\lambda_1 = 5h/m_e c$

2. $\lambda_2 = 5h/m_p c$

3. $\lambda_3 = 10h/m_e c$ > Ans. 1 ($5h/m_e c$)

4. $\lambda_4 = 10h/m_p c$ > Ans. 2 ($5h/m_p c$)

m_e (9.109×10^{-31} kg) << m_p (1.673×10^{-27} kg)

Therefore, $\lambda_3 > \lambda_4$ **Ans. 3 is correct**.

---

37. The emission spectrum below for a one-electron (hydrogen-like) species in the gas phase shows all the lines, before they merge together, resulting from transitions to the first excited state from higher energy states. Line A has a wavelength of 27.1 nm.

Increasing wavelength, λ →

(a) What are the upper and lower principal quantum numbers corresponding to the lines labeled A and B?
(b) Identify the one-electron species that exhibits the spectrum.

38. The emission spectrum below for a one-electron (hydrogen-like) species in the gas phase shows all the lines, before they merge together, resulting from transitions to the ground state from higher energy states. Line A has a wavelength of 10.8 nm.

Increasing wavelength, λ →

(a) What are the upper and lower principal quantum numbers corresponding to the lines labeled A and B?
(b) Identify the one-electron species that exhibits the spectrum.

---

37. The emission spectrum below for a one-electron (hydrogen-like) species in the gas phase shows all the lines, before they merge together, resulting from transitions to the first excited state from higher energy states. Line A has a wavelength of 27.1 nm.

$n = 3 \rightarrow n = 2$

Increasing wavelength, λ →

(a) What are the upper and lower principal quantum numbers corresponding to the lines labeled A and B?
(b) Identify the one-electron species that exhibits the spectrum.

38. The emission spectrum below for a one-electron (hydrogen-like) species in the gas phase shows all the lines, before they merge together, resulting from transitions to the ground state from higher energy states. Line A has a wavelength of 10.8 nm.

Increasing wavelength, λ →

(a) What are the upper and lower principal quantum numbers corresponding to the lines labeled A and B?
(b) Identify the one-electron species that exhibits the spectrum.

Slide 1 (top-left):

37. The emission spectrum below for a one-electron (hydrogen-like) species in the gas phase shows all the lines, before they merge together, resulting from transitions to the first excited state from higher energy states. Line A has a wavelength of 27.1 nm.

B A $n = 4 \rightarrow n = 2$ $n = 3 \rightarrow n = 2$

6 5

Increasing wavelength, $\lambda \longrightarrow$

(a) What are the upper and lower principal quantum numbers corresponding to the lines labeled A and B?
(b) Identify the one-electron species that exhibits the spectrum.

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B A

Increasing wavelength, $\lambda \longrightarrow$

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(b) Identify the one-electron species that exhibits the spectrum.

Slide 2 (top-right):

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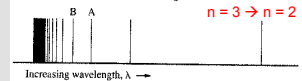
B A $n = 4 \rightarrow n = 2$ $n = 3 \rightarrow n = 2$

Increasing wavelength, $\lambda \longrightarrow$

(a) What are the upper and lower principal quantum numbers corresponding to the lines labeled A and B?
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38. The emission spectrum below for a one-electron (hydrogen-like) species in the gas phase shows all the lines, before they merge together, resulting from transitions to the ground state from higher energy states. Line A has a wavelength of 10.8 nm.
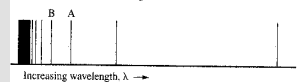
B A $n = 2 \rightarrow n = 1$

Increasing wavelength, $\lambda \longrightarrow$

(a) What are the upper and lower principal quantum numbers corresponding to the lines labeled A and B?
(b) Identify the one-electron species that exhibits the spectrum.

Slide 3 (middle-left):

37. The emission spectrum below for a one-electron (hydrogen-like) species in the gas phase shows all the lines, before they merge together, resulting from transitions to the first excited state from higher energy states. Line A has a wavelength of 27.1 nm.

B A $n = 4 \rightarrow n = 2$ $n = 3 \rightarrow n = 2$

Increasing wavelength, $\lambda \longrightarrow$

(a) What are the upper and lower principal quantum numbers corresponding to the lines labeled A and B?
(b) Identify the one-electron species that exhibits the spectrum.

38. The emission spectrum below for a one-electron (hydrogen-like) species in the gas phase shows all the lines, before they merge together, resulting from transitions to the ground state from higher energy states. Line A has a wavelength of 10.8 nm.

B A $n = 3 \rightarrow n = 1$ $n = 2 \rightarrow n = 1$

5 4

Increasing wavelength, $\lambda \longrightarrow$

(a) What are the upper and lower principal quantum numbers corresponding to the lines labeled A and B?
(b) Identify the one-electron species that exhibits the spectrum.

Slide 4 (middle-right):

37. (a) Line A is for the transition $n = 5 \rightarrow n = 2$, while Line B is for the transition $n = 6 \rightarrow n = 2$

(b) This transition corresponds to the $n = 5$ to $n = 2$ transition. Hence, $\Delta E = hc/\lambda$
$\Delta E = (6.626 \times 10^{-34} \text{ J s}) \times (2.998 \times 10^8 \text{ m s}^{-1}) \div (27.1 \times 10^{-9} \text{ m}) = 7.33 \times 10^{-18} \text{ J}$
$\Delta E = -Z^2 R_H/n_1^2 - -Z^2 R_H/n_2^2$
$4.577 \times 10^{-19} \text{ J} = -Z^2(2.179 \times 10^{-18})/(5)^2 + Z^2(2.179 \times 10^{-18})/(2)^2$
$Z^2 = 16.02$ and $Z = 4.00$ Thus, this is the spectrum for the Be^{3+} cation.

38. (a) Line A is for the transition $n = 4 \rightarrow n = 1$, while Line B is for the transition $n = 5 \rightarrow n = 1$

(b) This transition corresponds to the $n = 4$ to $n = 1$ transition. Hence, $\Delta E = hc/\lambda$
$\Delta E = (6.626 \times 10^{-34} \text{ J s}) \times (2.998 \times 10^8 \text{ m s}^{-1}) \div (10.8 \times 10^{-9} \text{ m}) = 1.84 \times 10^{-17} \text{ J}$
$\Delta E = -Z^2 R_H/n_1^2 - -Z^2 R_H/n_2^2$
$4.577 \times 10^{-19} \text{ J} = -Z^2(2.179 \times 10^{-18})/(5)^2 + Z^2(2.179 \times 10^{-18})/(2)^2$
$Z^2 = 9.004$ and $Z = 3.00$ Thus, this is the spectrum for the Li^{2+} cation.

Slide 5 (bottom-left):

Chemistry 110 Midterm 1 October 11, 2005

1. What are the possible values for the orbital angular momentum quantum number l given the following values for the other three quantum numbers: $n = 4$, $l = ?$, $ml = 0$, $ms = +1/2$

A. 0
B. 1, 2, 3, or 4
C. 3
D. 0, 1, 2, or 3
E. none of these answers

2. If the wavelength of an electron is equal to the radius of the n=2 orbit in He^+, what must the velocity of the electron be?

A. 6.88×10^6 m/s
B. 1.38×10^7 m/s
C. 3.44×10^6 m/s
D. 5.47×10^5 m/s
E. The velocity cannot have an exact value

3. The element seaborgium (Sg) has Z=106. What would you predict the electron configuration to of Sg^{2+} to be?

A. The same configuration as rutherfordium (Rf) Z=104
B. [Rn] $6d4$
C. [Rn] $5f14\ 6d4$
D. [Rn] $5f14\ 6d2\ 7s2$
E. [Rn] $5f12\ 6d4\ 7s2$

4. Praseodymium (Pr Z=59) is one of the lanthanide (rare earth) elements. Which of the following statements is a TRUE statement about the orbital into which the last electron was filled for the element Pr?

A. The number of radial nodes is greater than the number of angular nodes
B. The orbital is a 5f orbital
C. The orbital contains valence electrons
D. The orbital is one member of a set of seven degenerate orbitals
E. none of the other four statements is true

5. Davisson and Germer (as well as G. P. Thompson) conducted a famous and important experiment. Which of the following statements is TRUE concerning their experiment?

A. They observed diffraction patterns from photons confirming that light had wave properties
B. They observed diffraction patterns of electrons confirming that matter had wave properties
C. Their experiment contradicted DeBrogle's equation
D. They observed standing wave patterns from electrons confirming the existence of orbitals
E. None of the other four statements is true

Slide 6 (bottom-right):

Chemistry 110 Midterm 1 October 11, 2005

1. What are the possible values for the orbital angular momentum quantum number l given the following values for the other three quantum numbers: $n = 4$, $l = ?$, $ml = 0$, $ms = +1/2$

A. 0
B. 1, 2, 3, or 4
C. 3
D. 0, 1, 2, or 3
E. none of these answers

Answer D

2. If the wavelength of an electron is equal to the radius of the n=2 orbit in He^+, what must the velocity of the electron be?

A. 6.88×10^6 m/s
B. 1.38×10^7 m/s
C. 3.44×10^6 m/s
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Chemistry 110 Midterm 1 October 11, 2005

1. What are the possible values for the orbital angular momentum quantum number l given the following values for the other three quantum numbers: $n = 4$, $l = ?$, $ml = 0$, $ms = +1/2$

   A. 0
   B. 1, 2, 3, or 4
   C. 3
   D. 0, 1, 2, or 3
   E. none of these answers

2. If the wavelength of an electron is equal to the radius of the $n=2$ orbit in He^+, what must the velocity of the electron be?

   A. 6.88×10^6 m/s
   B. 1.38×10^7 m/s
   C. 3.44×10^6 m/s
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3. The element seaborgium (Sg) has $Z=106$. What would you predict the electron configuration to of Sg^{2+} to be?

   A. The same configuration as rutherfordium (Rh) $Z=104$
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   C. [Rn] $5f^{14} 6d^4$
   D. [Rn] $5f^{14} 6d^2 7s^2$
   E. [Rn] $5f^{12} 6d^4 7s^2$

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   B. They observed diffraction patterns of electrons confirming that matter had wave properties
   C. Their experiment contradicted DeBroglie's equation
   D. They observed standing wave patterns from electrons confirming the existence of orbitals
   E. None of the other four statements is true

Second quadrant (Answer A)

$$\lambda = n^2 a_0 / Z$$
$$u = h/m\lambda$$

Answer A

Fourth quadrant (Answer B)

Answer B

Old Midterm-1, Fall 2005 #24, Version 1

A. $n = 2$, $Z = 4$: $r_n = (2^2)a_0/4 = a_0$ Be^{3+}

B. $n = 1$, $Z = 1$: $r_n = (1^2)a_0/1 = a_0$ H

C. $n = 3$, $Z = 3$: $r_n = (3^2)/3 = 3a_0$ Li^{2+}

D. $n = 4$, $Z = 2$: $r_n = (4^2)/2 = 8a_0$ He^+

E. $n = 3$, $Z = 1$: $r_n = (3^2)/1 = 9a_0$ H

Answer E is correct.

Key Study Points for Chapter 8

1. Force of Coulombic attraction is directly proportional to the electrostatic charges (q_1 and q_2) and inversely proportional to the square of the separation between the charges (r^2).
2. Definition of isotope.
3. Definition of atomic number.
4. Definition of mass number.
5. General formula for an element or ion.
6. Orders of decreasing energy, frequency and wavelength for the different types of electromagnetic radiation.
7. Intensity of a wave is directly proportional to the square of the amplitude.
8. S.I. unit of frequency (Hz, s^{-1}).
9. Meaning of S.I. prefixes: pico, femto, nano, etc.

Key Study Points (2)

10. Range of visible electromagnetic radiation in nanometers: approx. 400 (blue)-800 (red) nm.
11. In-phase constructive interference → wave amplitude augmented.
12. Out-of-phase interference → wave amplitude cancelled.
13. Typical wave properties of visible light: diffraction and refraction.
14. $\lambda\nu = c$ relationship.
14. Planck's quantum theory relationship: $E = h\nu$.
15. Classical wave theory cannot account for atomic line spectra, blackbody radiation (ultraviolet catastrophe) and photoelectric effect.

Key Study Points (3)

16. Blackbody radiation → absorbs and emits all radiation regardless of v. Continuous spectrum, dependent on temperature. Best example → light emitted from a hot metal.
17. Rayleigh-Jeans theory predicted that the energy density for a blackbody would be infinite at low λ and that it was inversely proportional to λ^4 and directly proportional to absolute temperature T.
18. Einstein, photoelectric effect:: $KE_e = \frac{1}{2}m_e u^2 = h(\nu - \nu_0)$, where ν_0 is the threshold frequency and $h\nu_0 = \Phi$ (work function).
19. $E = h\nu$, energy associated with a photon (G.N. Lewis).
20. Bohr theory: electron energy in stationary states is constant

Key Study Points (4)

21. Bohr assumed that the orbital angular momentum is quantized: $nh/2\pi$.
22. Bohr's radius equation:
    $r_n = n^2 a_0$, where a_0 = Bohr radius (52.9 pm).
23. Bohr's energy equation:
    $E_n = -(hcR_H)/n^2 = -2.179 \times 10^{-18}\ J/n^2$.
24. n = 1 (ground state).
25. n = 2, 3, 4, …. (excited states).
26. Energy difference between orbits in Bohr's theory:
    $\Delta E = E_{final} - E_{initial} = hcR_H(1/n_i^2 - 1/n_f^2)$
    $= 2.179 \times 10^{-18}\ hcR_H(1/n_i^2 - 1/n_f^2)$ J
    $= hc/\lambda$
27. Emission: electron in excited state drops down to a lower energy state.

Key Study Points (5)

28. Absorption: electron is a lower energy state is excited to a higher energy one.
29. Two-photon transition: If the energy for a single photon transition is ΔE, then the total energy for a two-photon transition must equal ΔE.
30. For Lyman line spectra in UV: $n_f = 1$
31. For Balmer line spectra in visible: $n_f = 2$
32. For Paschen line spectra in IR: $n_f = 3$
33. Ionization energy is ΔE beween n = 1 and n = infinity states.
34. Bohr model applies to all H-like (one electron) systems, e.g., He^+, Li^{2+}, Be^{3+}, B^{4+}, C^{5+}. etc. Remember Z is involved.

Key Study Points (6)

28. de Broglie: $\lambda = h/mc$ for a photon and $\lambda = h/mu$ for matter (electron, neutron, etc.) where u = velocity of the matter.
29. Wave properties of matter confirmed → electron diffraction work of Davisson-Germer and Thomson. (U.K.).
30. Heisenberg uncertainty principle: $\Delta x \cdot \Delta p \geq h/4\pi$.
31. At the atomic level, the act of measurement changes the system.
32. Wavefunction squared (ψ^2) = electron probability
33. Quantum numbers (n, l, m_l and m_s).
34. s, p, d, f orbitals (l = 0, 1, 2, 3, respectively).
35. m_s = +1/2 or -1/2.
36. Stern-Gerlach experiment with Ag atoms.
37. Degenerate = orbitals of the same energy.
.

Key Study Points (7)

38. For H and H-like species, the orbitals for a given n value are degenerate.
39. No two electrons have exactly the same four quantum numbers.
40. Node = zero probability of finding an electron.
41. One s orbital, three p orbitals, five d orbitals and seven f orbitals.
42. For a s orbital (spherically symmetrical), the total integrated probability at a distance r from the nucleus depends on r^2, i.e., $P(r) = 4\pi r^2 \cdot R^2(r)$. When r = 0, P(r) = 0.
43. Screening strength of electrons in different orbitals in multielectron atoms: s > p > d.
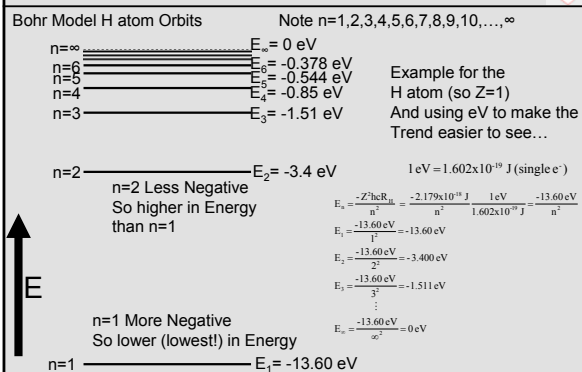44. E is directly proportional to Z^2_{eff}/n^2 and s < p < d for the same value of n in multielectron atoms

Key Study Points (8)

45. Aubau process for electron configurations.
46. Four rules: (1) minimum energy configuration; (2) Pauli exclusion principle; (3) Hund's rule; and (4) once orbitals of same energy are filled singly, additional electrons can be added with opposite spin.
47. Know the class exceptions to the Aufbau process: Cr, Cu.
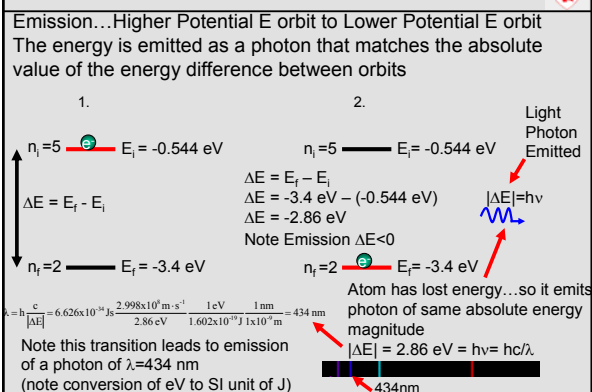48. Know the shapes and labelling of the s, p and d orbirtals

Key Study Points for Chapter 9

1. Distinction between metals and nonmetals.
2. Metals tend to lose electrons, while nonmetals tend to gain electrons.
3. The elements in Groups 1, 2, 13, 14, 15, 16 and 17 with $ns^x np^y$ valence electrons want to achieve the nearest noble gas electronic configuration in order to be more stable.
4. Some transition metals lose electrons to achieve noble gas electronic configurations. Many transition metals do not.
5. Extra stability of half-filled d shells (3d, 4d, 5d)
6. How to calculate atomic and ionic radii.
7. Comparison of atomic and ionic sizes.
8. Isoelectronic species and comparison of their atomic and ionic radii..
9. Periodic and group trends in atomic and ionic radii.
10. Periodic and group trends in ionization energy.
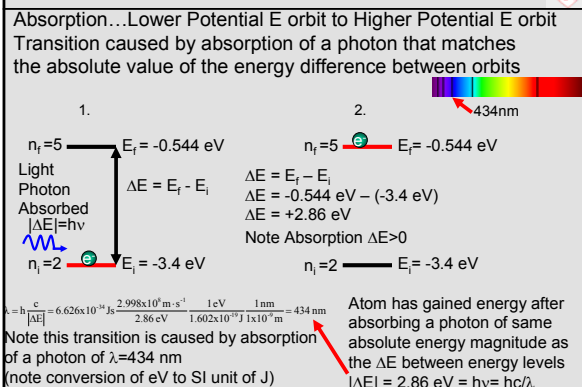11. Periodic and group trends in electron affinity.

Orbit Energies…Potential energies…E<0

Bohr Model H atom Orbits   Note n=1,2,3,4,5,6,7,8,9,10,…,∞

- $n=\infty$ —— $E_\infty = 0$ eV
- n=6 —— E_6= -0.378 eV
- n=5 —— E_5= -0.544 eV
- n=4 —— E_4= -0.85 eV
- n=3 —— E_3= -1.51 eV

Example for the H atom (so Z=1) And using eV to make the Trend easier to see…

- n=2 —— E_2= -3.4 eV

n=2 Less Negative So higher in Energy than n=1

$1 eV = 1.602 \times 10^{-19}$ J (single e$^-$)

$$E_n = \frac{-Z^2 hcR_{H}}{n^2} = \frac{-2.179 \times 10^{-18} J}{n^2} \cdot \frac{1 eV}{1.602 \times 10^{-19} J} = \frac{-13.60 eV}{n^2}$$

$$E_1 = \frac{-13.60 eV}{1^2} = -13.60 eV$$

$$E_2 = \frac{-13.60 eV}{2^2} = -3.400 eV$$

$$E_3 = \frac{-13.60 eV}{3^2} = -1.511 eV$$

$$\vdots$$

n=1 More Negative So lower (lowest!) in Energy

$$E_\infty = \frac{-13.60 eV}{\infty^2} = 0 eV$$

E ↑

- n=1 —— E_1= -13.60 eV

ΔE<0 Emission: Photon of Light Created

Emission…Higher Potential E orbit to Lower Potential E orbit The energy is emitted as a photon that matches the absolute value of the energy difference between orbits

1.
- n_i =5 —— E_i= -0.544 eV
- n_f =2 —— E_f= -3.4 eV

$\Delta E = E_f - E_i$

2.
- n_i =5 —— E_i= -0.544 eV
- n_f =2 —— E_f= -3.4 eV

$\Delta E = E_f - E_i$
$\Delta E = -3.4$ eV – (-0.544 eV)
$\Delta E = -2.86$ eV
Note Emission $\Delta E < 0$

Light Photon Emitted

$|\Delta E| = h\nu$

$$\lambda = h \frac{c}{|\Delta E|} = 6.626 \times 10^{-34} Js \frac{2.998 \times 10^8 m \cdot s^{-1}}{2.86 eV} \frac{1 eV}{1.602 \times 10^{-19} J} \frac{1 nm}{1 \times 10^{-9} m} = 434 nm$$

Note this transition leads to emission of a photon of λ=434 nm (note conversion of eV to SI unit of J)

Atom has lost energy…so it emits photon of same absolute energy magnitude

$|\Delta E| = 2.86$ eV = $h\nu$= hc/λ

434nm

ΔE>0 Absorption: Photon of Light Absorbed

Absorption…Lower Potential E orbit to Higher Potential E orbit Transition caused by absorption of a photon that matches the absolute value of the energy difference between orbits

434nm

1.
- n_f =5 —— E_f = -0.544 eV
- n_i =2 —— E_i = -3.4 eV

Light Photon Absorbed $|\Delta E| = h\nu$

$\Delta E = E_f - E_i$

2.
- n_f =5 —— E_f = -0.544 eV
- n_i =2 —— E_i = -3.4 eV

$\Delta E = E_f - E_i$
$\Delta E = -0.544$ eV – (-3.4 eV)
$\Delta E = +2.86$ eV
Note Absorption $\Delta E > 0$

$$\lambda = h \frac{c}{|\Delta E|} = 6.626 \times 10^{-34} Js \frac{2.998 \times 10^8 m \cdot s^{-1}}{2.86 eV} \frac{1 eV}{1.602 \times 10^{-19} J} \frac{1 nm}{1 \times 10^{-9} m} = 434 nm$$

Note this transition is caused by absorption of a photon of λ=434 nm (note conversion of eV to SI unit of J)

Atom has gained energy after absorbing a photon of same absolute energy magnitude as the ΔE between energy levels

$|\Delta E| = 2.86$ eV = $h\nu$= hc/λ

SI Units

■ Secondary Units written in terms of Primary

TABLE 1.1 ▶ Common SI Secondary Quantities and their Units

Secondary Quantity	Abbreviation	Unit	Equivalent in Other Units
Charge	q	Coulomb (C)	$A \cdot s$
Energy	$E; U; K$	Joule (J)	$kg \cdot m^2 \cdot s^{-2}$
Frequency	ν	Hertz (Hz)	s^{-1}
Force	F	Newton (N)	$kg \cdot m \cdot s^{-2}$
Pressure	P	Pascal (Pa)	$kg \cdot m^{-1} \cdot s^{-2}$ (force per unit area)
Power or intensity	I	Watt (W)	$kg \cdot m^2 \cdot s^{-3}$ (energy per second)
Volume	V	—	m^3

◾Common SI Prefixes

TABLE 1.2 ► Common Prefixes

Prefix	Example	Numerical Factor
femto-	femtosecond (fs)	10^{-15}
pico-	picomole (pmol)	10^{-12}
nano-	nanometer (nm)	10^{-9}
micro-	micromolar (μM)	10^{-6}
milli-	milliliter (mL)	10^{-3}
—		10^{0}
kilo-	kilogram (kg)	10^{3}
mega-	megapascal (MPa)	10^{6}
giga-	gigawatt (GW)	10^{9}
tera-	terahertz (THz)	10^{12}

◾One or Two Letter Abbreviations

◾Carbon: C Oxygen: O Nickel: Ni Gold: Au

$$_Z^A E^{\pm}$$

E = Chemical Symbol
Z = Atomic No. = p
A = Mass Number = p + n
± = Charge = p - e

◾Rayleigh-Jeans Law for Blackbody

$$\rho(\lambda) = \frac{8\pi kT}{\lambda^4}$$

ρ = Energy Density

k = Boltzmann's Constant

T = Absolute Temperature

You do not have to know this equation for calculations but you should know that the classical theory predicts an inverse relationship between Energy Density and Wavelength i.e. ρ proportional to λ^4

Ultraviolet catastrophe…energy density Goes to infinity as wavelength gets Small…A failure of the classical theory

EM Radiation

$$c = \nu\lambda$$

Any wave

$$u = \nu\lambda$$

u = wave velocity
ν= frequency
λ= wavelength

$$E = h\nu$$

$$h = \text{Planck's Constant}$$

$$h = 6.626 \times 10^{-34} \text{ Js}$$

$$c = \text{speed of light}$$

$$KE_{electron} = \tfrac{1}{2} m_e u^2$$

Photoelectric Effect

$$KE_{electron} = h\nu - h\nu_o = h\nu - \Phi$$

$$u = \text{velocity of the electron}$$

$$m_e = \text{mass of the electron}$$

$$\Phi = \text{"work function" for the metal}$$

the energy threshold for removing the e⁻

ν = Frequency of the light that is directed on the metal surface
ν_o = Threshold frequency for the metal

Energy of a Photon

$$E = h\nu$$

$$E = h\frac{c}{\lambda}$$

Radius and Energy of nth Bohr Orbit

$$r_n = n^2 a_o \text{ where } a_o = 52.9 \text{ pm the Bohr Radius}$$

$$E_n = \frac{-hcR_H}{n^2} = \frac{-2.179 \times 10^{-18} \text{ J}}{n^2}$$

Some Important Equations

Energy for change of discrete orbits from n_i to n_f
$\Delta E<0$ for emission $\Delta E>0$ for absorption

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

Energy change
for transition
Between orbits

$$\Delta E_{photon} = \left|\Delta E\right|_{e^- \text{ transition between orbits}}$$

Energy for photon emitted or absorbed for discrete transition between orbits
Note that this energy is simply the absolute value (i.e. positive) of the energy change
Shown above for the transition between orbits

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

Some Important Equations

Note that these equations apply to H as well (Z=1)...so they are general
Memorize these equations as they are general for H, He^+, Li^{2+} etc. (Note the Z)

Orbit energy for H-like species (only one electron)

$$E_n = \frac{-Z^2 hcR_H}{n^2} = \frac{-Z^2 \cdot 2.179 \times 10^{-18}\ J}{n^2}$$

$$Z = \text{Atomic Number} \quad a_o = 52.9\ pm$$

$$\Delta E = Z^2 hcR_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \qquad r_n = \frac{n^2 a_o}{Z}$$

Energy change for e^- jump between orbits from n_i to n_f for H like species
$\Delta E<0$ for emission $\Delta E>0$ for absorption
r_n is radius of nth orbit

Some Important Equations

Note that this equation applies to H as well (Z=1)...

$$\Delta E_{IE} = \frac{Z^2 hcR_H}{1^2}$$

Ionization energy for H-like species (change of discrete orbits from $n_i=1$ to $n_f=\infty$
for H-like species...e^- removed from atom by input of energy...
A photon is absorbed to remove the electron; hence $\Delta E>0$ for absorption
(this is assuming the usual ground state of n=1)

Two New Ideas

◼ de Broglie Equation predicts that matter should have wavelike properties

$$p = mu = \frac{h}{\lambda} \quad or \quad \lambda = \frac{h}{mu}$$

de Broglie's Equation

p=linear momentum
m=mass
u=velocity
h=Planck's constant
λ=wavelength

Two New Ideas

◼ The Uncertainty Principle

$$\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$$

Uncertainty
In Position

Uncertainty
In Momentum

h = Planck's Constant

Uncertainties are standard deviations

Multielectron Atoms

$$\left|\Psi(r,\theta,\varphi)\right|^2 = e^- \text{ Probability at } (r,\theta,\varphi)$$

◼ Ψ^2 is the probability of finding e^- at a single point in space. By mapping Ψ^2 at all points in 3D space (r,θ,ϕ), we can map the 3D electron probability space for each orbital

(r,θ,ϕ) Ψ^2 Probability at one point

Nucleus
At r =0

Each electron will "obey" its own wavefunction Ψ and "explore" a specific volume in space characteristic of the type of orbital the electron is in

Each wave function Ψ has four quantum numbers associated with it.
So Ψ describes the orbital in which an electron is most likely to be found

Each electron is specified by a set of four quantum numbers...They determine where the electron is found in an atom...Know the rules for the quantum numbers

Multielectron Atoms

- P(r)…the radial distribution function
- P(r) is the total integrated probability for a spherical shell at distance r from the nucleus

$$P(r) = 4\pi r^2 \cdot R^2(r)$$

Note you will not have to do calculations or plotting using the radial distribution function
You should understand the concept that it is "total" probability of finding electron density at a distance r from the nucleus…and that it goes to zero at the nucleus (r=0)

Radial Wave Function
Each type of orbital has
Its own functional form
Note P(r) = 0 for r = 0

Multielectron Atoms

Orbital energy depends on n and Z_{eff}

$$E_n = \frac{-Z_{eff}^2 hcR_H}{n^2} \qquad Z_{eff} = Z - S$$

Orbital Energy
For a multielectron atom

You will not have to calculate actual screening and Z_{eff} but you will have to understand the periodic trends and the screening abilities of electrons in various orbitals:
Inner vs. valence
s versus p versus d
S=Screening constant
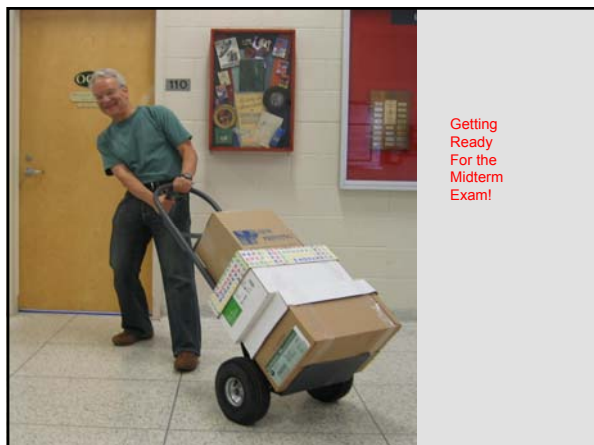
Getting Ready For the Midterm Exam!

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The End!

THAT'S ALL FOLKS!

GOOD LUCK IN THE REST OF THE

CHEM 110 COURSE