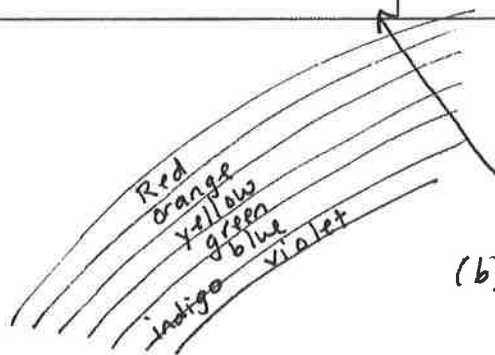


5



(a) as we proceed outward from the inner band of rainbow (from violet to red), wavelength increases.
(from $\approx 400 \text{ nm}$ to $\approx 700 \text{ nm}$)

(b) as we proceed outward from violet to red, frequency decreases
(from $\approx 7.5 \times 10^{14} \text{ s}^{-1}$ to $\approx 4.3 \times 10^{14} \text{ s}^{-1}$)
(so, energy decreases as well)

13 True / False:

(a) visible light is a form of electromagnetic radiation: True

(b) UV light has longer wavelengths than visible light: False UV is higher energy, so higher frequency, and shorter wavelength, than visible light.

(c) X rays travel faster than microwaves: False
Both are EM radiation, so both travel at a speed of "c" (assuming both are in a vacuum).

(d) EM radiation and sound waves travel at same speed. False EM radiation is about 10^6 times faster than sound.
 $c \approx 3 \times 10^8 \text{ m/s}$
Sound speed $\approx 3 \times 10^2 \text{ m/s}$ (343 m/s at sealevel, 20°C, dry air or $\approx 768 \text{ mph}$)

15 Arrange from in order of increasing wavelength:

X rays	UV	green	red	IR	radio
shortest λ					longest λ
highest ν					lowest ν
highest E					lowest E

23 (a) Calculate E_{photon} if $\nu = 6.75 \times 10^{12} \text{ s}^{-1}$

$$E = h\nu = (6.63 \times 10^{-34} \text{ J}\cdot\text{s}) (6.75 \times 10^{12} \text{ s}^{-1}) = \boxed{4.48 \times 10^{-21} \text{ J}}$$

(b) calculate E_{photon} if $\lambda = 322 \text{ nm}$

$$E = h\nu \quad \nu = \frac{c}{\lambda} \quad \Rightarrow \quad E = \frac{hc}{\lambda}$$

$$E = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s}) (3.00 \times 10^8 \text{ m/s})}{(322 \times 10^{-9} \text{ m})} = \boxed{6.18 \times 10^{-19} \text{ J}}$$

(c) Find λ if $E_{\text{photon}} = 2.87 \times 10^{-18} \text{ J}$

$$E = \frac{hc}{\lambda} \quad \lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s}) (3.00 \times 10^8 \text{ m/s})}{2.87 \times 10^{-18} \text{ J}}$$

$$= \boxed{6.93 \times 10^{-8} \text{ m}, \text{ or } 69.3 \text{ nm}}$$

27 Sunburn from UV: $\lambda = 325 \text{ nm}$

$$(a) E_{\text{photon}} = h\nu = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{(325 \times 10^{-9} \text{ m})} = \boxed{6.12 \times 10^{-19} \text{ J}}$$

w/ no extra SF

(b) convert to energy per mole of photons

$$(6.12 \times 10^{-19} \text{ J/photon}) \left(\frac{6.02 \times 10^{23} \text{ photons}}{1 \text{ mole of photons}} \right) = 368424 \text{ J/mole}$$

(c) # photons in a 1.00 mJ burst

$$(1.00 \times 10^{-3} \text{ J}) \left(\frac{1 \text{ photon}}{6.12 \times 10^{-19} \text{ J}} \right) = \boxed{1.63 \times 10^{15} \text{ photons}}$$

$$\boxed{3.68 \times 10^5 \text{ J/mole}}$$

$$\text{or } \boxed{368 \text{ kJ/mole}}$$

(d) The avg $\frac{\text{kJ}}{\text{mole}}$ was already calculated: $\boxed{368 \text{ kJ/mole}}$

(the book says that this is the average bond energy in skin molecules...)

31. Molybdenum (Mo) has min frequency of $\boxed{1.09 \times 10^{15} \text{ s}^{-1}}$ ← given ...

(a)

$$(a) E_{\text{min}} = h\nu_{\text{min}} = (6.63 \times 10^{-34} \text{ J}\cdot\text{s})(1.09 \times 10^{15} \text{ s}^{-1}) = 7.2267 \rightarrow \boxed{7.23 \times 10^{-19} \text{ J}}$$

$$(b) \lambda_{\text{max}} = \frac{c}{\nu}$$

$$= \frac{3.00 \times 10^8 \text{ m/s}}{1.09 \times 10^{15} \text{ s}^{-1}} = 2.75 \times 10^{-7} \text{ m, or } \boxed{275 \text{ nm}}$$

(c) Mo irradiated by EM radiation w/ $\lambda = 120 \text{ nm}$
find KE_{max} of emitted electrons.

$$KE = E_{\text{photon}} - \text{work function}$$

$$= \frac{hc}{\lambda} - \phi$$

$$= \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{120 \times 10^{-9} \text{ m}} - 7.2267 \times 10^{-19} \text{ J}$$

$$= 1.6575 \times 10^{-18} \text{ J} - 7.2267 \times 10^{-19} \text{ J}$$

$$= 9.3483 \times 10^{-19} \text{ J} \rightarrow \boxed{9.3 \times 10^{-19} \text{ J}} \quad (\text{actually 1SF..})$$

$$(d) KE = \frac{1}{2} m v^2$$

$$v = \sqrt{\frac{2KE}{m}} = \left(\frac{2(9.3483 \times 10^{-19} \text{ J})}{9.11 \times 10^{-31} \text{ kg}} \right)^{1/2} = \boxed{1.4 \times 10^6 \text{ m/s}}$$

(really 1SF..)

(e) Na: $\phi = 4.41 \times 10^{-19} \text{ J}$ ($\lambda_{\text{max}} = 451 \text{ nm}$: blue)

Mo: $\phi = 7.23 \times 10^{-19} \text{ J}$ ($\lambda_{\text{max}} = 275 \text{ nm}$: UV)

Sodium requires less energy to lose an electron, so sodium would be expected to be higher on the activity series of metals; (more reactive than Mo to form a cation)

(Though photoelectric effect is a different process from a single replacement rxn...)

- 33 Bohr suggested that electrons in the atom can only have certain/specific energies. This is consistent with line spectra: if electrons can only have certain energies, they should only be able to absorb or emit specific energies as they change energy levels. Line spectra show that elements ~~do~~ (and compounds) do only emit certain energies: instead of emitting/absorbing the full spectrum (rainbow of colors, only certain colors are seen (or absorbed)).
Line spectra provide evidence that electron energies are quantized.

- 35 If electrons make transition to higher n -level, they will need to absorb energy, and as e^- transition to lower n -levels, they will need to emit energy.
(As n -level increases, the e^- distance from the nucleus increases. e^- are attracted to nucleus due to positive charge on nucleus, so energy is needed to separate e^- from nucleus.)
- So.. (a) $n=4$ to $n=2$ e^- will release / emit energy.
(b) 2.12 \AA "orbit" to 8.46 \AA e^- will absorb energy since moving further from nucleus.
(c) an e^- adds to an H^+ ion and ends up in $n=3$ thus e^- came from outside the atom (essentially, $n=\infty$) it will release energy as it gets closer to nucleus from far away/ outside of atom.

- 37 $E = -2.18 \times 10^{-18} \text{ J} \left(\frac{1}{n^2} \right)$ or $\Delta E = -2.18 \times 10^{-18} \text{ J} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$
- (a) $E @ n=2$
 $E = (-2.18 \times 10^{-18} \text{ J}) \left(\frac{1}{2^2} \right) = -2.18 \times 10^{-18} \text{ J} \left(\frac{1}{4} \right) = \underline{-5.45 \times 10^{-19} \text{ J}}$ at $n=2$
- $E @ n=6$ $E = -2.18 \times 10^{-18} \text{ J} \left(\frac{1}{6^2} \right) = \frac{-2.18 \times 10^{-18} \text{ J}}{36} = -6.05(5) \times 10^{-20} \text{ J}$
 $\underline{-6.06 \times 10^{-20} \text{ J}}$ at $n=6$
- energy released if e^- "falls" from $n=6$ to $n=2$:
- $-6.05(55) \times 10^{-20} \text{ J} - -5.45 \times 10^{-19} \text{ J} = 4.8444 \times 10^{-19} \text{ J} \rightarrow \underline{4.84 \times 10^{-19} \text{ J}}$
- So, the wavelength of light emitted will be:
- $E = \frac{hc}{\lambda}$ $\lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{4.84(44) \times 10^{-19} \text{ J}} = 4.106 \times 10^{-7} \text{ m}$
 $\underline{\underline{411 \text{ nm}}}$
- This will be violet light, so yes: visible.

$$E = -2.18 \times 10^{-18} \text{ J } \left(\frac{1}{n^2} \right)$$

$$(a) \text{ energy for } n=1 = -2.18 \times 10^{-18} \text{ J } \left(\frac{1}{1^2} \right) = -2.18 \times 10^{-18} \text{ J}$$

$$\text{energy for } n=\infty = -2.18 \times 10^{-18} \text{ J } \left(\frac{1}{\infty^2} \right) = 0$$

so, to "move" the e- from $n=1$ to $n=\infty$ (so it leaves the atom) would require $2.18 \times 10^{-18} \text{ J}$

$$(2.18 \times 10^{-18} \text{ J}) \left(\frac{6.02 \times 10^{23} \text{ photons}}{\text{mole photons}} \right) \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right) = \boxed{1310 \text{ kJ/mole}}$$

(1312 before rounding)

(b) The experimental value for ionization of hydrogen is 1310 kJ/mole.. it agrees perfectly to the 3 sig figs we have.

$$47. \quad \Delta p \Delta x \geq \frac{h}{2} = \frac{h}{2 \cdot 2\pi} = \frac{h}{4\pi}$$

(a) 1.50 mg mosquito moving @ $v = 1.40 \frac{\text{m}}{\text{s}} \pm .01 \frac{\text{m}}{\text{s}}$
calculate uncertainty in position.

$$\Delta p = \Delta m v = m \Delta v = (1.50 \times 10^{-6} \text{ kg}) (.01 \frac{\text{m}}{\text{s}}) = 1.50 \times 10^{-8} \text{ kg m/s}$$

$$\Delta x \geq \frac{h}{4\pi \Delta p} = \frac{6.63 \times 10^{-34} \text{ J}\cdot\text{s}}{4\pi (1.50 \times 10^{-8} \text{ kg m/s})} = 3.52 \times 10^{-27} \frac{(\text{kg m}^2/\text{s}^2)(\text{s})}{\text{kg m/s}}$$

$$\boxed{\Delta x \geq 4 \times 10^{-27} \text{ m}}$$

(b) proton w/ speed of $(5.00 \pm .01) \times 10^4 \frac{\text{m}}{\text{s}}$
mass of a proton = $1.6726 \times 10^{-27} \text{ kg}$
calculate uncertainty in position

$$\Delta p = m \Delta v = (1.6726 \times 10^{-27} \text{ kg}) (.01 \times 10^4 \frac{\text{m}}{\text{s}}) = 1.6726 \times 10^{-25} \text{ kg m/s}$$

$$\Delta x \geq \frac{h}{4\pi \Delta p} = \frac{6.63 \times 10^{-34} \text{ J}\cdot\text{s}}{4\pi (1.6726 \times 10^{-25} \text{ kg m/s})} = 3.15 \times 10^{-10} \text{ m}$$

$$\boxed{\Delta x \geq 3 \times 10^{-10} \text{ m}} \\ \text{or } 3 \text{ \AA}$$

49a) Why does Bohr model violate the uncertainty principle?

The uncertainty Principle says that it is impossible to know both the position and momentum of something with great precision; the more precisely you determine position, the less precisely you can determine the momentum, and vice versa.

Specifically, the momentum and position of an electron in an atom can not be determined simultaneously, at least not precisely.

In the Bohr model, the electron was thought to orbit in a well defined (circular) orbit, with specific energies. The uncertainty principle says we can't say what the electron's exact position (or pathway) and energy are. Instead the e- location can only be "pinpointed" to a space about the size of an orbital!

#46
on next page

50.

- (a) Bohr model says e^- orbits nucleus at $R = 0.53 \text{ \AA}$
 Quantum mechanical model says "most probable distance" from nucleus is 0.53 \AA . Explain difference...

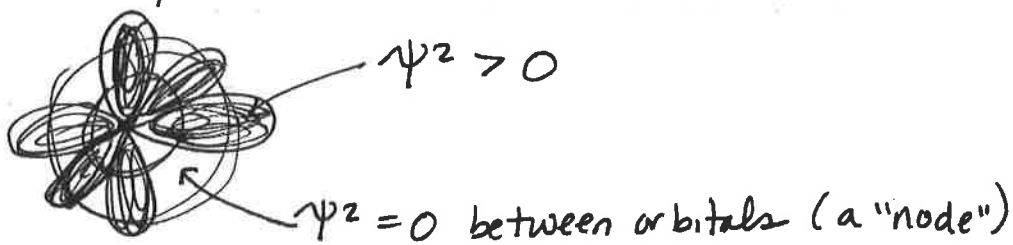
Bohr's model violates ^{the} uncertainty principle, since it says that the e^- is exactly 0.53 \AA away from the nucleus; in a defined location, on a specific path (with a specific momentum as well)

The Quantum model says that the e^- is typically about 0.53 \AA from nucleus, but could be closer or further from nucleus. It does not have a precise location (or a well defined path). It is consistent w/ uncertainty principle since the e^- location is not pinpointed. 0.53 \AA from nucleus would be in the middle of the orbital the e^- is in.

- (b) Classical physics treated the electron as a particle, with specific momentum and location. Schrodinger's wave equation takes the electron's wave properties into account. This fits with the uncertainty principle; an electron's (or anything's) position can only be determined to about the size of its wavelength; you can't determine position more precisely than $\pm \approx$ a wavelength.

- (c) ψ^2 is the probability of "finding" the electron in a certain location *

ψ^2 is essentially zero in the parts of an atom between orbitals, and it is greater than zero in places where there are orbitals.



* if you integrate ψ^2 over a certain region/volume of the atom, you get the probability of finding the electron in that region.

46.

Find the wavelength of an electron that has been accelerated to a speed of $8.95 \times 10^6 \text{ m/s}$
 Mass of an $e^- = 9.1094 \times 10^{-31} \text{ kg}$

$$\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34} \text{ J s}}{(8.95 \times 10^6 \frac{\text{m}}{\text{s}})(9.1094 \times 10^{-31} \text{ kg})} = \underline{\underline{8.13 \times 10^{-11} \text{ m}}}$$


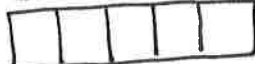

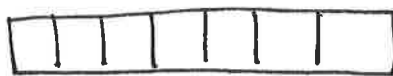
or $0.813 \times 10^{-10} \text{ m}$
 0.813 \AA

This is fairly comparable to the size of a small atom or ion. A typical atom diameter would be a few angstroms, for example,
 potassium: $D \approx 4 \text{ \AA}$, Fluorine: $D \approx 1.4 \text{ \AA}$

- 62 In a chlorine atom, the average distance of a 3s electron is less than the avg distance of a 3p electron (distance from the nucleus). therefore (a) the 3p orbital would be expected to be higher in energy. (b) The 3s electron would require more energy than a 3p electron, and less energy than a 2p electron, to leave the atom.

The 3s electron is further from the nucleus, on avg, than a 2p electron. so the 3s electron has more energy than the 2p electron, so the 3s electron will need less additional energy to entirely leave the atom.

- 65 Maximum # of electrons that can occupy each subshell:

(a) 3p :	6		(3 orbitals x 2 e ⁻ per orbital) = 6 e ⁻
(b) 5d :	10		(5 orbitals) (2 e ⁻ / orbital) = 10 e ⁻
(c) 2s :	2		1 orbital holds 2 e ⁻
(d) 4f :	14		(7 orbitals) (2 e ⁻ / orbital) = 14 e ⁻

↑
answers.

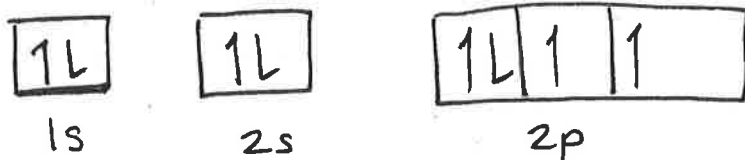
- 67 (a) valence electrons are the electrons in the outermost shell (highest n-level) of an atom
 (b) core electrons are the electrons not in the outermost shell: they are at an n-level below the valence level

for example, Lead: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^2$

$[1s^2 \text{ through } 5p^6]$ ← these are the core electrons.
 $4f^{14}, 5d^{10}$

↑
The highest n-level is n=6, so these are the valence electrons.

- (c) each box in an orbital diagram represents an orbital (!).
 (d) So, each box (orbital) can hold up to 2 electrons (represented by arrows)
 ↑
or "half arrows"



for example, oxygen's orbital diagram is shown above

- ↑ shows one electron with a spin of +1/2
 ↓ shows one electron with a spin of -1/2

68 Report # of valence electrons, # of core electrons, # of unpaired e⁻ and classify as paramagnetic or diamagnetic.

(a) Carbon $1s^2 2s^2 2p^2$

Orbital diagram: $1s$ (1L), $2s$ (1L), $2p$ (1L, 1L, 0)

Annotations: 2 core e⁻ (under 1s and 2s), 4 valence e⁻ (under 2s and 2p), 2 unpaired e⁻ (paramagnetic) (under 2p orbitals)

(b) Phosphorus $1s^2 2s^2 2p^6 3s^2 3p^3$

Orbital diagram: $3s$ (1L), $3p$ (1L, 1L, 1L)

Annotations: 10 core electrons (under 1s, 2s, 2p), 5 valence electrons (under 3s and 3p), 3 unpaired electrons (paramagnetic) (under 3p orbitals)

(c) Neon $1s^2 2s^2 2p^6$

Orbital diagram: $2s$ (1L), $2p$ (1L, 1L, 1L)

Annotations: 2 core electrons (under 1s), 8 valence electrons (under 2s and 2p), zero unpaired electrons (diamagnetic) (under 2p orbitals)

69 Write e⁻ config and orbital diagrams; and classify as para/dia magnetic

(a) Cs (# 55) $[Xe] 6s^1$ $6s$ (1L) paramagnetic

(b) Ni (# 28) $[Ar] 4s^2 3d^8$ $4s$ (1L), $3d$ (1L, 1L, 1L, 1L, 1L) paramagnetic

(c) Se (# 34) $[Ar] 4s^2 3d^{10} 4p^4$ $4s$ (1L), $3d$ (1L, 1L, 1L, 1L, 1L), $4p$ (1L, 1L, 1L) paramagnetic

(d) Cd (# 48) $[Kr] 5s^2 4d^{10}$ $5s$ (1L), $4d$ (1L, 1L, 1L, 1L, 1L) diamagnetic

(e) U (# 92) ~~$[Xe] 5f^4$~~ * paramagnetic: see below

(f) Pb (# 82) $[Xe] 6s^2 4f^{14} 5d^{10} 6p^2$ $6s$ (1L), $4f$ (1L, 1L, 1L, 1L, 1L, 1L, 1L, 1L), $5d$ (1L, 1L, 1L, 1L, 1L), $6p$ (1L, 1L) } para magnetic

* Uranium is an exception to the rules...

It is actually $[Xe] 7s^2 5f^3 6d^1$
 You wouldn't have to predict or explain why on a test.

Orbital diagram for Uranium: $7s$ (1L), $5f$ (1L, 1L, 1L, 0, 0, 0, 0), $6d$ (1L, 0, 0, 0)

72 (a) [Noble gas] $ns^2 np^5$ group VII A (halogens), 1 unpaired e⁻

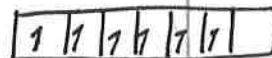


(b) [] $ns^2 (n-1)d^2$ group IV B (Ti, Zr, etc) 2 unpaired e⁻



(c) [] $ns^2 (n-1)d^{10} np^1$
group III A (B, Al ...)
1 unpaired electron

(d) [] $ns^2 (n-2)f^6$
Sm and Pu
6 unpaired e⁻



77

6.77 Certain elements emit light of a specific wavelength when they are burned. Historically, chemists used such emission wavelengths to determine whether specific elements were present in a sample. Characteristic wavelengths for some of the elements are given in the following table:

Ag	328.1 nm	Fe	372.0 nm
Au	267.6 nm	K	404.7 nm
Ba	455.4 nm	Mg	285.2 nm
Ca	422.7 nm	Na	589.6 nm
Cu	324.8 nm	Ni	341.5 nm

(a) which emit in visible spectrum
visible is from 400 nm to 700 nm,
so Ba, Ca, K, and Na
emit visible light

(b) The highest-energy photons
are emitted by gold (Au)
since it has the shortest λ

The lowest energy photons are
emitted by Na, since it has
the longest λ .

$$(c) v = 6.59 \times 10^{14} s^{-1}$$

$$\lambda = \frac{c}{v} = \frac{3.00 \times 10^8 m/s}{6.59 \times 10^{14} s^{-1}} = 4.55 \times 10^{-7} m$$

$$\text{or } \boxed{455 \text{ nm}}$$

emitted by Barium.

81



(a) If a plant is orange, it probably absorbs blue light strongly,
since orange and blue are complementary colors.

(b) If a plant absorbs strongly @ 455 nm, it is absorbing blue / blue-violet,
(I looked on page 209)

So the carotenoid will be orange / orange-yellow.

Upon actually reading the question,

I have realized that they asked

as to calculate the photon energy of the 455 nm light.

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} J \cdot s)(3.00 \times 10^8 m/s)}{(455 \times 10^{-9} m)} = \boxed{4.37 \times 10^{-19} J}$$

Quantum part 2 (Chapter 6)

79

UV-A : 320 to 380 nm (and UV-C is 100 to 280 nm)
 UV-B : 290 to 320 nm or 290?

(a) frequency if $\lambda = 320 \text{ nm}$ ("extreme" UV ≈ 10 to 100 nm or 120 nm)
 (EUV)

$$c = \lambda \nu \quad \nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{320 \times 10^{-9} \text{ m}} = 9.375 \times 10^{14} \text{ s}^{-1}$$

$$\boxed{9.38 \times 10^{14} \text{ Hz}}$$

(b) calculate the energy of a mole of these photons.

$$E = h\nu \quad \text{so } E_{\text{photon}} = (6.63 \times 10^{-34} \text{ J}\cdot\text{s}) (9.375 \times 10^{14} \text{ s}^{-1})$$

$$= 6.2156 \times 10^{-19} \text{ J}$$

$$\left(\frac{6.2156 \times 10^{-19} \text{ J}}{\text{photon}} \right) \left(\frac{6.02 \times 10^{23} \text{ photons}}{\text{mole}} \right) = 374181 \text{ J/mole}$$

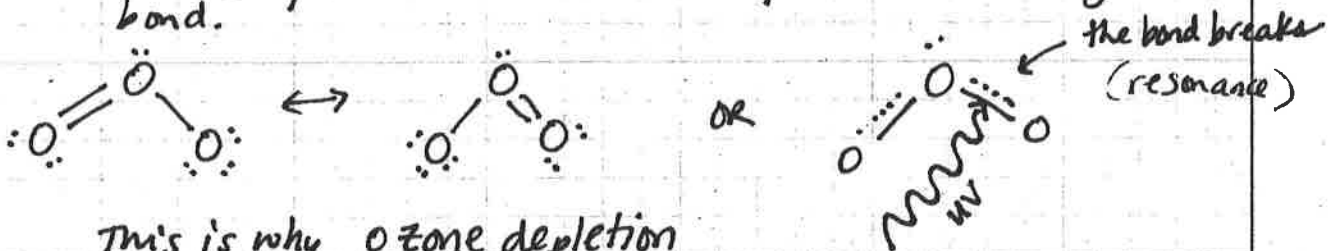
(c) UV-B photons are more energetic than UV-A photons

$$\boxed{374 \text{ kJ/mole}}$$

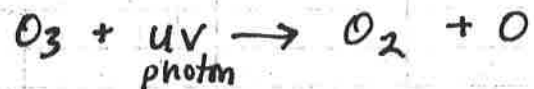
UV-B photons have a shorter wavelength than UV-A
 Energy is inversely prop to wavelength ($E = \frac{hc}{\lambda}$)

(d) yes, since UVB is more energetic than UV-A, it would make sense that UVB would be a greater cause of sunburn than UVA

(e) Ozone (O_3) absorbs UVB; when it is "hit" by a UVB photon, it absorbs the photon, breaking its bond.



This is why ozone depletion is a concern; as O_3 levels decrease, more UV will reach the earth's surface from outer space.



Increased UV reaching earth would cause increase skin cancer, cataracts, etc.

28

$\text{:N} \equiv \text{N:}$ Bond energy = 941 kJ/mole

$$\lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{(941000 \text{ J/mole}) \left(\frac{1 \text{ mole}}{6.02 \times 10^{23}} \right)} = 1.27 \times 10^{-7} \text{ m} \text{ or } \boxed{127 \text{ nm}}$$

The bond requires a wavelength of 127 nm or shorter to break. This corresponds to ultraviolet light, or, specifically, UV-C.

(b) N_2 requires a higher photon energy, and shorter wavelength, than nearly all other bonds. It is a triple bond, so very strong.