Review Packet – Independent Practice

Check your answers at http://shakeribchem.weebly.com

	#e	Atom	e configuration	orbital diagram
1.	8	oxygen	$1s^2 2s^2 2p^4$	↑↓ (1s) ↑↓ (2s) ↑↓ ↑ ↑ (2p)
2.	19	potassium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑↓ (2p) ↑↓ (3s) ↑↓ ↑↓ ↑↓ (3p) ↑ (4s)
3.	1	hydrogen	$1s^1$	↑ (1s)
4.	10	neon	$1s^2 2s^2 2p^6$	↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑↓ (2p)
5.	15	phosphorous	$1s^2 2s^2 2p^6 3s^2 3p^3$	↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑↓ (2p) ↑↓ (3s) ↑ ↑ ↑ (3p)

Abbreviated configurations

6. 30 zinc [Ar] $4s^2\ 3d^{10}$

7. 56 barium [Xe] $6s^2$

8. 35 bromine [Ar] $4s^2 3d^{10} 4p^5$

9. 12 magnesium [Ne] $4s^2$

10. 9 fluorine [He] $2s^2 2p^5$

Electron Configuration Elements (atoms) and Ions

Write the **electron configuration** and **orbital notations** for the following Atoms and ions:

Element / Ions	Atomic number	# of e-	Electron Configuration/Orbital Diagrams
F^{1-}	9	10	$1s^2 2s^2 2p^6$ ↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑↓ (2p)
Na^{1+}	11	10	$1s^2 2s^2 2p^6$ ↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑↓ (2p)

Al^{3+}	13	10	$1s^2 2s^2 2p^6$
			↑↓ ↑↓ ↑↓ ↑↓ ↑↓ 1s 2s 2p
Cl^{1-}	17	18	$1s^2 2s^2 2p^6 3s^2 3p^6$
			↑↓ ↑↓ ↑↓↑↓↑↓ ↑↓ ↑↓↑↓↑↓ 1s 2s 2p 3s 3p
Br^{1-}	35	36	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
			↑↓ ↑↓ ↑↓↑↓↑↓ ↑↓ ↑↓↑↓↑↓ ↑↓ ↑↓↑↓↑↓↑↓↑↓ ↑↓↑↓↑↓ 1s 2s 2p 3s 3p 4s 3d 4p
Mg^{2+}	12	10	$1s^2 2s^2 2p^6$
			↑↓ ↑↓ ↑↓↑↓↑↓ 1s 2s 2p

E. In the space below, write the full (unabbreviated) electron configurations of the following elements:

1) sodium $1s^2 2s^2 2p^6 3s^1$

2) iron $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$

3) bromine $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$

4) barium $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$

5) neptunium 93 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^4$

F. In the space below, write the Noble Gas (abbreviated) electron configurations of the following elements:

6) cobalt [Ar] $3d^7 4s^2$

7) silver [Kr] $5s^1 4d^{10}$ (anomaly)

8) tellurium [Kr] $5s^2 4d^{10} 5p^4$

9) radium [Rn] $7s^2$

10) lawrencium 103 [Rn] $7s^2 5f^{14}$

11) manganese [Ar] $4s^2 3d^5$

~~12) silver~~

13) nitrogen [He] $2s^2 2p^3$

14) sulfur $[Ne]3s^23p^3$

15) argon $[Ne]3s^23p^6$

G. Given an element with atomic number 11, provide the following information:

a. How many electrons will fill each of the following shells and list their subshells.
1st shell: 2
2nd shell: 8
3rd shell: 1

b. Is this element likely to form a cation or anion? cation

c. How do you know?

easier to lose 1 in outer shell

d. What charge will the ion formed by this element have?

+1

H. Explain, based on electron configuration, why the noble gases are so unreactive. Use helium and neon as examples to illustrate your explanation.

full outer shell

Ne - $1s^22s^22p^6$ He - $1s^2$ ↖ s-orbital holds $2e^-$, it is full

full

I. Determine which of the following electron configurations are not valid, and why.

a. $1s^22s^22p^63s^23p^64s^24d^{10}4p^5$

3d next

b. $1s^22s^22p^63s^33d^5$

3p next

c. $[Ra]\ 7s^25f^8$

not noble gas

d. $[Kr]\ 5s^24d^{10}5p^5$

no error

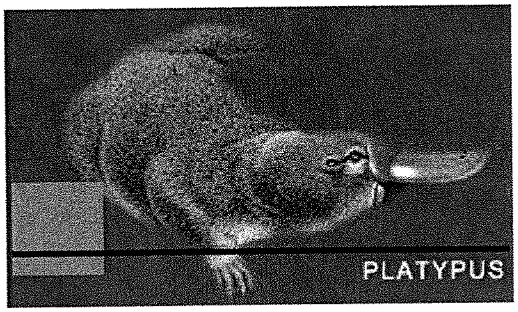
e. [Xe]

can't condense to no sublevels!!

J. Is light a particle or a wave?

Is light composed of waves or of particles? If light is waves, then one can always reduce the amount of light by making the waves weaker, while if light is particles, there is a minimum amount of light you can have - a single ``particle'' of light. In 1905, Einstein found the answer: Light is both! In some situations, it behaves like waves, while in others it behaves like particles.

This may seem odd. How can light act like both a wave and a particle at the same time? Consider a duck-billed platypus. It has some duck-like properties and some beaver-like properties, but it is neither. Similarly, light has some wavelike properties and some particle like properties, but it is neither a pure wave nor a pure particle.



A wave of light has a wavelength, defined as the distance from one crest of the wave to the next, and written using the symbol λ. The wavelengths of visible light are quite small: between 400 mm and 650 nm, where 1 nm = 10^{-9} m is a ``nanometer'' - one billionth of a meter. Red light has long wavelengths, while blue light has short wavelengths.

A particle of light, known as a photon, has an energy *E*. The energy of a single photon of visible light is tiny, barely enough to disturb one atom; we use units of "electron-volts", abbreviated as eV, to measure the energy of photons. Photons of red light have low energies, while photons of blue light have high energies.

+ violet

The energy *E* of a photon is proportional to the wave frequency *v*

$$E = h\,v$$

where the constant of proportionality h is the **Planck's Constant,** h = 6.626×10^{-34} J s.
Also, the relationship between frequency and wavelength can be defined as:

$$v = \frac{c}{\lambda}$$

where c is the speed of light (3×10^{8} metres per second). So, photons still have a wavelength. A famous result of quantum mechanics is that the wavelength relates to the energy of the photon. The longer the wavelength, the smaller the energy. For instance, ultraviolet photons have shorter wavelengths than visible photons, and thus more energy. This is why they can give you sunburn, while ordinary light cannot.

One means by which a continuous spectrum can be produced is by thermal emission from a black body. This is particularly relevant in astronomy. Astronomical spectra can be combination of absorption and emission lines on a continuous background spectrum.

One convenient method of exciting atoms of an element is to pass an electric current through a sample of the element in the vapor phase. This is the principle behind the spectrum tubes. A spectrum tube contains a small sample of an element in the vapor phase. An electric discharge through the tube will cause the vapor to glow brightly. The glow is produced when excited electrons emit visible light energy as they return to their original levels.

When visible light energy from a spectrum tube is passed through a diffraction grating, a bright line spectrum, or line-emission spectrum is produced. Each element has its own unique emission spectrum by which it can be identified, analogous to a fingerprint. Such a spectrum consists of a series of bright lines of definite wavelength. Each wavelength can be mathematically related to a definite quantity of energy produced by the movement of an electron from one discrete energy level to another. Thus, emission spectra are experimental proof that electrons exist in definite, distinctive energy levels in an atom.

Refer to this text to answer the following questions.

1. What is the difference between a line spectrum and a continuous spectrum?

   distinct wavelengths emitted/absorbed

   all colors shown (like by a black body)

2. Each line in the emission spectrum of the hydrogen corresponds to an electromagnetic radiation with a specific wavelength. Match the 4 observed colors with the following wavelengths: 410 nm, 434 nm, 486 nm, and 656 nm.

   410 nm ↓ violet

   434 nm ↓ violet/indigo

   486 nm ↓ blue

   656 nm ↓ red

3. How are electrons "excited"? What happens when the electrons "relax"?

   ↓ E is applied

   ↓ photon emitted, e^- returns to ground state

4. Each element has its own unique line emission spectrum, just like fingerprints. Explain how this technique can be used to determine the elemental composition of stars.

   Can use emission spectra of star's light

5. How can the difference in the brightness of spectral lines be explained?

   ↓ more common e^- jump

6. According to the modern theory of the atom, where may an atom's electrons be found?

   "clouds" or orbitals of probability

7. State the equation used to determine the energy content of a packet of light of specific frequency.

   $E = h\nu$

8. What form of energy emission accompanies the return of excited electrons to the ground state?

   ↓ photon emitted (EM)

9. Explain, in terms of electron transition, how bright-line spectra are produced by atoms.

   - Energy excites e^-, jumps up to higher E-level
   - Returning to ground state, photon of specific λ is emitted
   - If λ (↑ wavelength) of photon is within 400nm - 700nm, seen as visible color band on spectrum
   - Larger energy "jumps" = $\downarrow\lambda = \uparrow\nu$

useful equations

$c = \lambda \times \nu$ $c = 3.00 \times 10^{8}$ m/s

$E = h \times \upsilon$ $h = 6.63 \times 10^{-34}$ J·s

1 m = 1×10^{9} nm 1 kJ = 1000 J

example

Light with a wavelength of 525 nm is green. Calculate the energy in joules for a green light photon.

- find the frequency: $c = \lambda \times \upsilon$ $\nu = \dfrac{c}{\lambda}$ $\nu = \dfrac{3.00\times10^{8}\, m/s}{525\ nm \times \dfrac{1\ m}{1\times10^{9}\, nm}}$ $\nu = 5.71\times10^{14}\, 1/s$

- find the energy: $E = h \times \upsilon$ $E = (6.626\times10^{-34}\, J\cdot s)(5.71\times10^{14}\, 1/s)$ $E = 3.78\times10^{-19}\, J\,/\,photon$

K. Solving for photon energy

* 1. Find the energy, in joules per photon, of microwave radiation with a frequency of 7.91×10^{10} s^{-1}.
2. Find the energy in kJ for an x-ray photon with a frequency of 2.4×10^{18} s^{-1}.
3. A ruby laser produces red light that has a wavelength of 500 nm. Calculate its energy in joules.
4. What is the frequency of UV light that has an energy of 2.39×10^{-18} J?
5. The frequency of violet light is about $7.495\text{x}10^{14}$ Hz. What is the wavelength of this radiation?

$$\lambda = \frac{3.00\times10^{8}\, m/s}{7.495\times10^{14}\, s^{-1}} = 4.00\times10^{-7}\, m \times \frac{10^{9}\, nm}{1\, m} = 400 nm \ (\text{violet})$$

** 6. What is the frequency of a photon that has a wavelength of 1428 nm? What type of radiation is this?
7. A popular radio station broadcasts at 107.9 MHz (M = 10^6). Find the wavelength of this radiation, in meters, and the energy of one of these photons, in J. What type of radiation is this?
8. What is the energy of a photon with:

a) a wavelength of 58 nm? What type of radiation is it? → UV ray

$$E = h\left(\frac{c}{\lambda}\right) = (6.626\times10^{-34}\, Js)\left(\frac{3.00\times10^{8}\, m/s}{5.8\times10^{-8}\, m}\right) = 3.4\times10^{-18}\, J$$

b) a wavelength of 0.065 cm? What type of radiation is it?

$$E = h\left(\frac{c}{\lambda}\right) = (6.626\times10^{-34}\, Js)\left(\frac{3.00\times10^{8}\, m/s}{0.00065\, m}\right) = 3.1\times10^{-22}\, J$$

9. Which of the following are directly related?

a) energy and wavelength
b) wavelength and frequency
[✓ c)] frequency and energy

** 6. $\nu = \dfrac{c}{\lambda} = \dfrac{3.00\times10^{8}\, m/s}{1.428\times10^{-6}\, m} = 2.10\times10^{14}\, s^{-1}$

IR ray

7. $\lambda = \dfrac{c}{\nu} = \dfrac{3.00\times10^{8}\, m/s}{1.079\times10^{-4}\, m}$

$= 2.78\times10^{12}\, s^{-1}$

radio wave

* 1. $E = h\nu = (6.626\times10^{-34}\, Js)(7.91\times10^{10}\, s^{-1}) = 5.24\times10^{-23}\, J$

2. $E = (6.626\times10^{-34}\, Js)(2.4\times10^{18}\, s^{-1}) = \dfrac{1.6\times10^{-15}}{1000}\, J = 1.6\times10^{-18}\, kJ$

3. $E = (6.626\times10^{-34}\, Js)\left(\dfrac{3.00\times10^{8}\, m/s}{5\times10^{-7}\, m}\right) = 4\times10^{-19}\, J$

4. $\nu = \dfrac{E}{h} = \dfrac{(2.39\times10^{-18}\, J)}{(6.626\times10^{-34}\, Js)} = 3.61\times10^{15}\, J$

L. Planck recognized that energy is quantized and related the energy of radiation (emitted or absorbed) to its frequency.

$\Delta E = n h v$ where n = integer and h = Planck's constant = 6.626 x 10 -34 J s

$$\Delta E = -R_H Z^2 \left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$$

R_H = Rydberg constant = 2.178×10^{-18}J

Z = nuclear charge = 1 for H, 2 for He

1. What is the energy needed to remove the remaining electron from He⁺ in its ground state? Is it easier or harder to remove the electron from He⁺ than from H?

$$\Delta E = -(2.178\times10^{-18}\text{J})(2)^2\left[\frac{1}{\infty} - \frac{1}{1}\right]$$

$$= \underbrace{8.712\times10^{-18}\text{ J (for He}^+)}_{\text{much greater, because attractive force of nucleus is pulling "harder"}}$$

2.178×10^{-18} J for H

M. Average Atomic Mass

1. Calculate the average atomic (aam) mass of bromine. One isotope of bromine has an atomic mass of 78.92amu and a relative abundance of 50.69%. The other major isotope of bromine has an atomic mass of 80.92amu and a relative abundance of 49.31%.

aam = (0.5069 · 78.92amu) + (0.4931 · 80.92amu) =

2. Lithium-6 is 4% abundant and lithium-7 is 96% abundant. What is the average mass of lithium? Try this WITHOUT a calculator!

3. Here are three isotopes of an element: ^{12}C ^{13}C ^{14}C

a. The element is: carbon

b. The number 6 refers to the atomic #

c. The numbers 12, 13, and 14 refer to the mass number

d. How many protons and neutrons are in the first isotope? 6p, 6n

e. How many protons and neutrons are in the second isotope? 6p, 7n

f. How many protons and neutrons are in the third isotope? 6p, 8n

4. Complete the following chart:

Isotope name	*atomic #*	*mass #*	*# of protons*	*# of neutrons*	*# of electrons*
potassium-37	19	37	19	18	19
oxygen-17	8	17	8	9	8
uranium-235	92	235	92	143	92
uranium-238	92	238	92	146	92
boron-10	5	10	5	5	5
boron-11	5	11	5	6	5

neutral atoms

5. Argon has three naturally occurring isotopes: argon-36, argon-38, and argon-40. Based on argon's reported $\overset{\text{a. a. m}}{\text{atomic mass}}$, which isotope exist as the most abundant in nature? Explain without calculations.

$\downarrow$

39.948amu

^{40}Ar is the most abundant, since it is closer to the a.a.m., and therefore its % must be greater to influence the weighted average.

6. Calculate the atomic mass of lead. The four lead isotopes have atomic masses and relative abundances of 203.973 amu (1.4%), 205.974 amu (24.1%), 206.976 amu (22.1%) and 207.977 amu (52.4%).

$$\text{aam} = (0.014 \cdot 203.973) + (0.241 \cdot 205.974) + (0.221 \cdot 206.976) + (0.524 \cdot 207.977)$$

$$= 207.217 \text{ amu}$$

7. Titanium has five common isotopes: 46Ti (8.0%), 47Ti (7.8%), 48Ti (73.4%), 49Ti (5.5%), 50Ti (5.3%). What is the average atomic mass of titanium?

$$\text{aam} = (0.08 \cdot 46) + (0.078 \cdot 47) + (0.734 \cdot 48) + (0.55 \cdot 49) + (0.053 \cdot 49) + (0.055 \cdot 50)$$

$$= 74{,}875 \text{ amu}$$

8. What is the atomic mass of hafnium if, out of every 100 atoms, 5 have a mass of 176, 19 have a mass of 177, 27 have a mass of 178, 14 have a mass of 179, and 35 have a mass of 180.0?

$$\text{aam} = \left(176 \cdot \frac{5}{100}\right) + \left(177 \cdot \frac{19}{100}\right) + \left(178 \cdot \frac{27}{100}\right) + \left(179 \cdot \frac{14}{100}\right) + \left(180 \cdot \frac{35}{100}\right)$$

$$= 178.55 \text{ amu}$$