Name _____ KEY _____ Lab TA/time _____

1.   (5 pts) For the Hydrogen atom, calculate the **wavelength** (in nm) of a photon for a transition from the 7st shell to the 2nd shell? Is this photon being emitted **or** absorbed by the atom? Is this in the visible portion of the electromagnetic spectrum (and how do you know)? (**Show work and explain your answers.**)

Use the Rydberg Eqn.

$$\frac{1}{\lambda} = R_H \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

This is technically for the wavelength of a photon emitted ($n_f < n_i$) so it comes out positive. For an absorption ($n_f > n_i$) it would come out negative but you ignore the negative sign since the energy of a photon can't be negative (take the absolute value) and just realize it's an absorption.

$$\frac{1}{\lambda} = (1.097 \times 10^7 \text{ m}^{-1}) \left| \left(\frac{1}{2^2} - \frac{1}{7^2} \right) \right| = 2.51\underline{8}62 \times 10^6 \text{ m}^{-1}$$

$$\lambda = 3.97\underline{0}42 \times 10^{-7} \text{ m} = 397 \text{ nm}$$

The photon is being **emitted** since the transition is from n = 7 to n = 2.

This would **NOT** be in the **visible** region of the spectrum. The visible region is 400-750 nm. This would be in the **Ultraviolet** (UV) region.

2.   (3 pts) Which of the following electron transitions requires the smallest energy to be absorbed by the hydrogen atom?

a) from n = 1 to n = 2
b) from n = 2 to n = 4
c) from n = 2 to n = 6
d) from n = 3 to n = 6
e)* from n = 4 to n = 6

Remember, the energy diagram. The further the energy levels get from the nucleus the closer together they get in terms of energy. There's a big gap between n=1 and n=2, the gap between n=2 and n=3 is smaller, and so on and so forth.

From n=4 to n=6 would be the smallest gap, smallest energy needed to be absorbed.

3. (5 pts) Calculate the de Broglie wavelength for a Ne molecule moving at 0.850 m/s? If this wavelength corresponded to a photon of light would we be able to see it in the visible portion of the electromagnetic spectrum (and how do you know)?

$$\lambda = \frac{h}{mv}$$ where h = Planck's constant (6.626 x 10^{-34} J•s), m = mass and v = velocity

Need the mass of 1 Ne arom (20.179 amu) in kg,   (1 J = 1 kg•m^2•s^{-2})

$$? \text{ kg} = 20.179 \text{ amu} \times \frac{1.66 \times 10^{-24} \text{ g}}{1 \text{ amu}} \times \frac{1 \text{ kg}}{10^3 \text{ g}} = 3.3\underline{4}97 \times 10^{-26} \text{ kg}$$

$$\lambda = \frac{h}{mv} = \frac{(6.626 \times 10^{-34} \text{ J•s})[(1 \text{ kg•m}^2\text{•s}^{-2})/1\text{J}]}{(3.3\underline{4}97 \times 10^{-26} \text{ kg}) (0.850 \text{ m/s})}$$

$$\lambda = 2.3\underline{2}71 \times 10^{-8} \text{ m} = 23.\underline{2}7\text{nm} = 23.3\text{nm}$$

You would **NOT** be able to see this in the **visible** region of the spectrum. The **visible region** is **400-750** nm. This would be in the **UV** region (borderline with X-ray region).

4. The uncertainty principle states that

a) matter and energy are really the same thing
b) it is impossible to know how many electrons there are in an atom
c) there can only be one unknown digit in a number
d) it is impossible to know what orbitals the electrons occupy.
e)* it is impossible to know exactly both the position and momentum of an electron.

The Heisenberg Uncertainty Principle states it is impossible to determine exactly both the position (location) and momentum (velocity) of a particle. It applies to all particles but is only important when the masses are extremely small (subatomic particles). It can be stated in the following mathematical way:

$$\Delta x \bullet \Delta(m\bullet v) > \frac{h}{4\pi}$$

Δx = uncertainty in position
Δ(mv) = uncertainty in momentum (p = mv)
m = mass of the particle
v = velocity of particle
h = Planck's constant = 6.626 x 10^{-34} J•s   (1 J = 1 kg•m^2/s^2)

This means if either the position or momentum are known exactly the other has a minimum uncertainty greater than h/4π.

5.   (2 pts) Which of the following sets of quantum numbers are **not** allowed for an electron in an atom?

1)* $n = 2$ $\ell = 2$ $m_\ell = -1$ $m_s = +1/2$
2)* $n = 3$ $\ell = 1$ $m_\ell = +2$ $m_s = +1$
3) $n = 4$ $\ell = 3$ $m_\ell = -2$ $m_s = -1/2$
4) $n = 8$ $\ell = 6$ $m_\ell = 0$ $m_s = +1/2$

See #11 above for a description of the quantum numbers.

$n = 1, 2, 3, ..., \infty$ $\ell = 0, 1, 2, ... , (n-1)$ $m_\ell = -\ell, ..., 0, ..., +\ell$ $m_s = +\frac{1}{2}$ and $-\frac{1}{2}$

Correct q.n.:

For (3) when $n = 4$ and $\ell = 3$ the possible values of m_ℓ are $m_\ell = -3, -2, -1, 0, 1, 2, 3$ and $m_s = -\frac{1}{2}$ ok

For (4) when $n = 8$ and $\ell = 6$ the possible values of m_ℓ are $m_\ell = -6, ..., 0, ..., 6$ and $m_s = +\frac{1}{2}$ ok

Incorrect q.n.:

For (1) when $n = 2$ and $\ell = 0, 1$ (ℓ can't be 2 when $n = 2$)

For (2) when $n = 3$ and $\ell = 1$ can't have $m_\ell = +2$ or $m_s = 1$ (m_s can be only $+\frac{1}{2}$ and $-\frac{1}{2}$)

6.   Select an answer which includes all of the **CORRECT** statements given below.

     1) There are **nine f** orbitals in an f subshell.

     2) A **10p subshell** can have a maximum of **6 electrons**.

     3) Each **p orbital** within a subshell consists of **four lobes** along the axes.

     4) There are 11 **subshells** in the 11th **shell**.

     5) For the ground state of an atom, **within a shell**, all the **orbitals** of a **subshell** are filled **prior** to putting electrons in the **next subshell**.

The correct statements are 2, 4 & 5

2) A **p subshell** can have a maximum of **6 electrons**. (True for all p subshells, no matter what shell).
4) There are 11 **subshells** in the 11th **shell**. (# **subshells in a shell** = **shell #, n**). There are 11² (121) orbitals.
5) In the **ground state** of an atom, all **orbitals within** a **subshell** are **filled prior** to **putting electrons** in the **next subshell**.

The corrected answers for 1 & 3 would be:

1) There are **seven** f **orbitals** in an f subshell.
3) Each **p orbital** within a subshell consists of **two lobes**. (figure eight, ∞; dumbbell)

The following table shows values of q.n. ℓ, corresponding letters, # orb. and max. # e- in subshell

$\ell =$	0	1	2	3	4	5...
subshell letter	s	p	d	f	g	h...
# orbitals in subshell	1	3	5	7	9	11...
max. # e⁻ in subshell	2	6	10	14	18	22...

7.   The electron configuration for an excited state of Gd is [Xe]6s¹4f⁷5d². How many **unpaired** electrons are in this excited atom? (Be careful!)

     [Xe] 6s¹ 4f⁷ 5d²   or   **[Xe] 4f⁷ 5d² 6s¹**
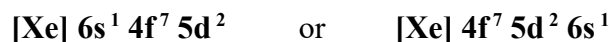
If the # e- is ≤ # e- in a half-filled subshell, then;
the number of unpaired e- = the number of e- in the subshell.
If the # e- is > half filled (or filled), then;
# of unpaired e- = (the maximum # of e- in the subshell) - (the # of e- present in the subshell).

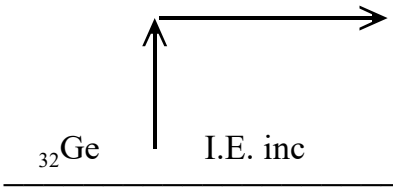There are 7 unpaired e- in the 4f, 2 unpaired e- in the 5d & 1 unpaired e- in the 6s.

This gives a total of **10 unpaired** e-.

_1_ _1_ _1_ _1_ _1_ _1_ _1_   _1_ _1_ __ __ __   _1_
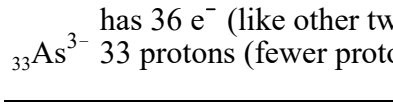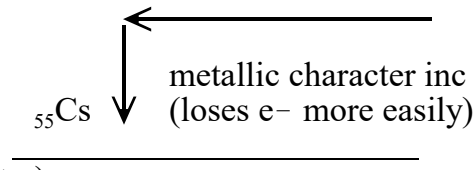   **4f**   **5d**   **6s**

8. Which of the following electron configurations corresponds to an element in the same group as an element with a $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$ electron configuration?

a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$ b) $1s^2 2s^2 2p^6 3s^2 3p^4$

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

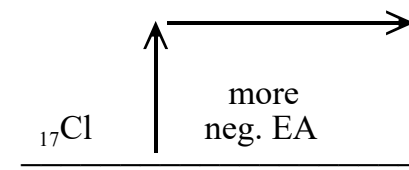e)* $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$

Both $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^2\, 3d^{10}\, 4p^2$ and $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^2\, 3d^{10}\, 4p^6\, 5s^2\, 4d^{10}\, 5p^2$ are in Group 4A (IVA). (There are 4 e– in the valence shell, the one with the highest n value, $4s^2 4p^2$ and $5s^2 5p^2$.)

9. (4 pts) Fill in the blanks with the correct answer to each of the following.

(a) Which has the **smaller** ionization energy, $_{32}$Ge or $_{15}$P?

$_{32}$Ge   I.E. inc

(b) Which is **largest** (size): $_{35}$Br⁻, $_{34}$Se²⁻ or $_{33}$As³⁻?
anions larger than cations in isoelectronic series
(same # e⁻); $_{33}$As³⁻ has higher (–) charge (fewer protons for the same # of electrons - things spread out)

$_{33}$As³⁻ has 36 e⁻ (like other two) but 33 protons (fewer protons)

(c) Which reacts **more readily** with chlorine, $_{37}$Rb or $_{55}$Cs? (i.e. lose an e⁻ more easily)
These metals would lose e– to Cl (more metallic react faster)

$_{55}$Cs   metallic character inc (loses e– more easily)

(d) Which has the **more** negative electron affinity, $_{34}$Se or $_{17}$Cl?(several exceptions)

$_{17}$Cl   more neg. EA

10. (3 pts) Explain how the effective nuclear charge, Z_{eff}, changes going across a row from left to right in the periodic table and why?

Going left to right across a row the Z_{eff} increases (rather rapidly). We are using the following for Z_{eff},

$Z_{eff} = Z - S$ Z = # protons   S = Shielding factor

There are many ways to determine S. We're using the simplest one, S equals the number of core electrons. Going across a row the number of core electrons remains constant. However, the # protons increases by 1 each time you move one column to the right. Since S is constant and Z inc. by 1 when moving over 1 column the Zeff increases by 1 each time you move one column to the right. For example, in row 2 the core electrons are the two electrons in the 1s (He core). As you move over one column to the right Z inc. by 1 but the # of core electrons, S, stays constant at 2 and thus the Z_{eff} inc. by 1. By the time you get to fluorine the Z_{eff} is +7. This has big effects on the size and I.E.

11. (3 pts) Explain why the electron affinity for Beryllium is positive while that for Boron is negative.

EA is the energy which occurs when an electron is added to a gaseous atom. Look at the e⁻ configurations for each atom.

B: [He] $2s^2\ 2p^1$ Be: [He] $2s^2$

In Be the 2s subshell is filled. A filled subshell is pretty stable. The added e⁻ must be put in a previously empty p subshell that is higher in energy. Also, the e- which would go into the 2p is shielded a little by the existing 2s electrons. All this requires energy. In B the extra e⁻ goes in an empty 2p orbital in the 2p subshell which already has 1 e⁻ to begin with and there is no extra shielding due to the existing e⁻ in the 2p subshell.

12. (2 pts) Write the equation corresponding to the first electron affinity of Oxygen.

Electron affinity: **energy change which occurs when an electron is added to a gaseous neutral atom or anion**. The first EA is for the first electron being added. For most neutral atoms energy is released (negative EA) when an electron is added. There are a few cases when the energy change is positive (such as the noble gases, Be, Mg and N, the reasons for this were discussed in lecture and the textbook).

$$O\ (g)\ +\ e^-\ \rightarrow\ O^-\ (g)$$ first E.A. (First electron being added.)

13. (2 pts) Write the equation corresponding to the second ionization energy of Ca.

Ionization: losing one or more electron from a gaseous atom or ion.
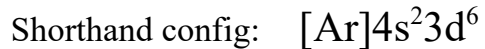
Ionization energy: **energy required** to remove an electron from a gaseous neutral atom or ion.

$$Ca\ (g)\ \rightarrow\ Ca^+\ (g)\ +\ e^-$$ first I.E.

$$Ca^+\ (g)\ \rightarrow\ Ca^{2+}\ (g)\ +\ e^-$$ second I.E.
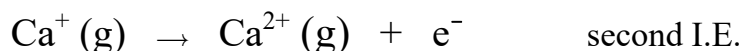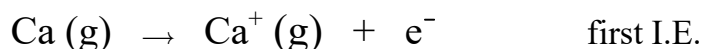
14. (3 pts) Write the electron configuration for the following ion by starting with the electron configuration for the neutral atom and then the ion given. **Show work.**
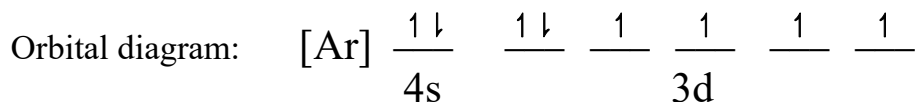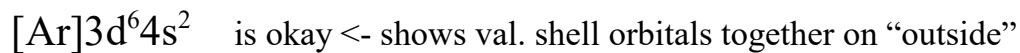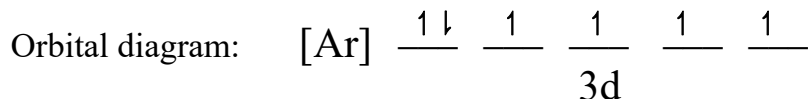
Fe^{2+}

Fe:

Shorthand config: $[Ar]4s^2 3d^6$ Use noble gas from previous period to
                                                            represent the core (inner) e^-.
                                    or

                        $[Ar]3d^6 4s^2$ is okay <- shows val. shell orbitals together on "outside"

Orbital diagram: $[Ar]$ $\frac{\uparrow\downarrow}{4s}$ $\frac{\uparrow\downarrow}{}$ $\frac{\uparrow}{}$ $\frac{\uparrow}{3d}$ $\frac{\uparrow}{}$ $\frac{\uparrow}{}$

In forming ions for transition metals electrons come out of the ns orbital before the (n-1)d orbitals. Remove the two electrons in the 4s. For Fe^{3+} remove one electron from the 3d.

Fe^{2+} $[Ar]3d^6$

Orbital diagram: $[Ar]$ $\frac{\uparrow\downarrow}{}$ $\frac{\uparrow}{}$ $\frac{\uparrow}{3d}$ $\frac{\uparrow}{}$ $\frac{\uparrow}{}$

15. (3 pts) The electrons that are removed from $_{48}Cd$ to form the Cd^{2+} ion are from the ____ subshell and the electrons that are removed from $_{83}Bi$ to form the Bi^{2+} ion are from the _____ subshell.

   a)* 5s; 6p   b) 5s; 5s   c) 4d; 4f   d) 5d; 6p   e) 4s; 5d

Need to write out the e- configuration for each atom and then look at which electrons will come out when it forms an ion. Remember, the **valence** e^- come out first when forming ions.

   Cd $[Kr] 5s^2 4d^{10}$

   Cd^{2+} $[Kr] 4d^{10}$ electrons come out of the 5s (valence shell) first

   Bi $[Xe] 6s^2 4f^{14} 5d^{10} 6p^3$

   Bi^{2+} $[Xe] 6s^2 4f^{14} 5d^{10} 6p^1$ electrons come out of the 6p (valence shell) first