Unit 3 Review Packet

Name Key

1) Draw the Bohr model for the following atoms. Include the full electron configuration under your diagram.

a) Phosphorous

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

b) Fluorine with a -1 charge

add 1 e⁻ to last sublevel

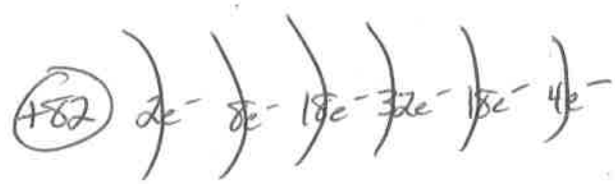
$1s^2 2s^2 2p^6$

c) Calcium with a +2 charge

remove 2e⁻ (make sure they are valence!)

$1s^2 2s^2 2p^6 3s^2 3p^6 \underline{4s^0}$

d) Lead



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

2) Identify the neutral element based up on the following electron configurations:

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

P

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

Fe

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

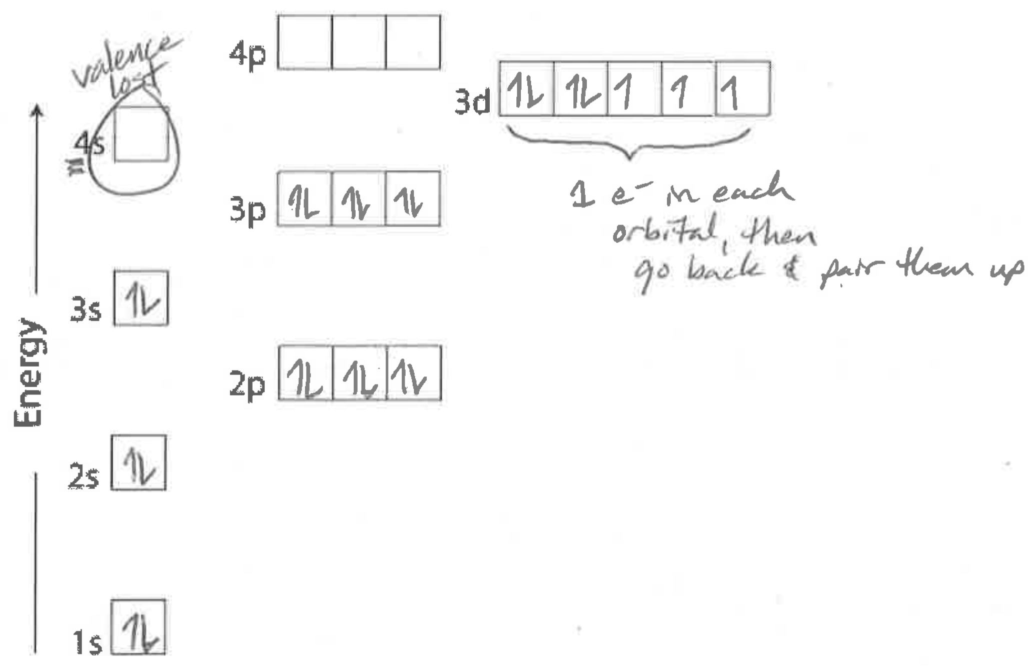
I

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

Na

3) Fill in the following AUFBAU diagram for cobalt with a +2 charge.

$\hookrightarrow 27\ e^- = 25\ e^-$

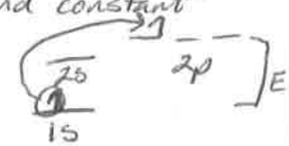


4) Discuss the Quantum Mechanical Model for atoms and explain how it adapts ideas from both Bohr's and Rutherford's models.

Rutherford's electron cloud idea has been adapted to show multiple clouds (orbitals), which describe specific locations for electrons using probability maps. They predict, with 95% certainty, where a pair of electrons are likely to be found. Electrons can pair up in orbitals as their opposite spin creates opposing magnetic poles whose attraction overcomes the repulsion of their like charge.

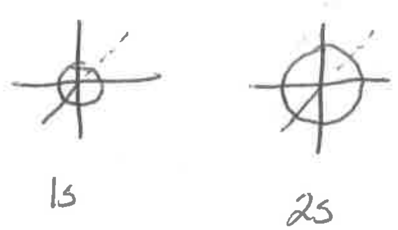
energy levels
↓
sublevels (s,p,d,f)
↓
orbitals (1,3,5,7)
↓
e^- (2,6,10,14)

Bohr's quantized energy levels have been adapted by showing they can be further broken down into sublevels, which are made up of a certain shape and number of orbitals, each of which can be occupied by a pair of electrons. Electrons occupy the lowest energy sublevel available and only pair up if they have to (1 in each cloud first). The sublevels are quantized, meaning it always requires a specific amount of energy for an e^- to excite between sublevels and constant light emission waves are observed for specific elements

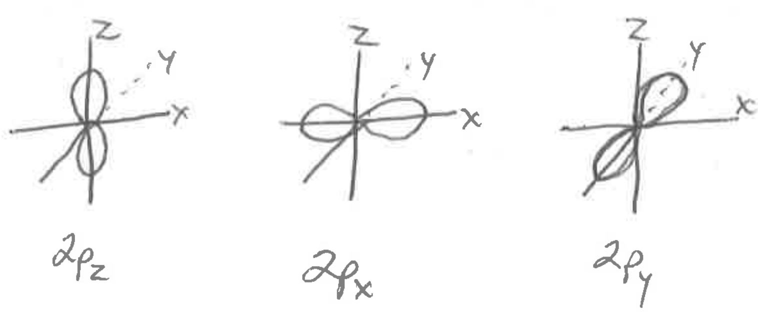


5) Draw cloud diagrams to help explain the following.

a. The difference between a 1s orbital cloud and a 2s orbital cloud.



b. The 2p sublevel is made up of three orbital clouds.



6) Explain why each orbital only holds 2 electrons and describe the orientation of those electrons.

- Opposite spin creates opposing magnetic poles
- Attraction between the poles counteracts the repulsion of their like charge.
- No other e^- could join the orbital (3rd or 4th ...) because they would strongly repel due to their like charge and like spin/poles.

7) Describe the relationship between the following terms: energy levels, orbitals, electrons and sublevels.

Energy Levels (1, 2, 3, 4, etc.) are made of...

↓

Sublevels (s, p, d, f) which are made of...

↓

orbitals (1, 3, 5, 7) each of which can be occupied by...

↓

$2e^-$ (2, 6, 10, 14) Which is why each energy level can hold the number of e^- that it does.

8) Explain why is the 4s sublevel filled before the 3d sublevel.

4s is a lower energy sublevel, despite being part of the next energy level, than 3d. It is therefore occupied by e^- first.

9) Be able to discuss excitation/emission and electromagnetic waves (wavelength, frequency, energy) using the quantum mechanical model.

a. Describe what happens during excitation and emission using the quantum mechanical model.

The QMM, means the electrons are located in quantized energy levels. If you can excite an electron with enough energy it can jump from one energy level to another. It will then emit excess energy as a light wave.

b. An electron undergoes excitation and emission between the 2s sublevel and 3s sublevel, emitting a light wave with a wavelength of 412nm. Would excitation/emission between 4s and 5s create a wave of shorter or longer wavelength? Explain.

Longer, it takes less energy to excite these electrons as they nucleus ~~has~~ is holding these electrons less tightly. Allowing them to jump from one energy level to another with less energy.

c. 4s and 5s are $1.2x10^{-18}$J of energy apart, what wavelength of light would need to be absorbed to allow an electron to jump up from 4s to 5s?

$E = \frac{hc}{\lambda}$

$1.2 \cdot 10^{-18} J = \frac{(6.63 \cdot 10^{-34} Js)(3.0 \cdot 10^{8} m/s)}{\lambda}$

$E = hf$

$\boxed{\lambda = 1.7 \cdot 10^{-7} m}$

$h = 6.63 \cdot 10^{-34}$ Js

$c = 3.0 \cdot 10^{8}$ m/s

10) Define the following terms:

a) Atomic Radius

The distance from the nucleus out to the valence shell

b) Electronegativity

The ability of an atom to attract anothers atom's valence electrons towards itself

c) Ionization Energy

The energy required to remove a valence electron from an atom

11) Describe how atomic radius is theoretically measured vs how scientists actually measure the distance.

Theoretically: can ~~mease~~ measure from nucleus to valence electrons

Actually: Take the distance from nucleus to nucleus of 2 bonded atoms and divide by 2.

12) Explain why the following trends occur:

a) Atomic radius decreases across a period

As you go across a period the elements gain more protons allowing them to better pull in their electrons, which decreases atomic radius.

b) Atomic radius increases down a group

When you go down a group, you add more energy levels. The more energy levels you have the larger the atomic radius will be.

c) Electronegativity increases across a period

Electronegativity is inversely related with atomic radius. If an atom has a small atomic radius it can use its nucleus a positive attractive to steal valence electrons. If there is a larger atomic radius, it is harder for it to use its nucleus to steal electrons.

d) Ionization energy decreases down a group

As the atomic radius gets bigger, ionization energy decreases. This is due to the nucleus having less of a hold on valence electrons making them eaiser to take. So, since the atomic radius gets larger as you go down a group, the ionization energy decreases.

13) What does it mean when two atoms are isoelectric? Provide an example.

That they have the same number of electrons.

Ex: F^- and ~~Na~~ Ne

14) Explain why noble gases are not reactive.

Their energy levels are all full, which means they have a very small radius and a high ionization energy. They also have a 0 for electronegativity, as they do not want anymore electrons. Therefore, they are not reactive.

15) Circle the atom with the largest radius.

a) Na or Na^+ b) F or F^- c) P or Cl d) Br^- or Kr

16) Circle the atom with the lowest electronegativity.

a) Na or Al b) F or Br c) P or In d) Br^- or Br

17) Circle the atom with the highest ionization energy.

a) Si or O b) Se or Ca c) Cs or Mg d) K^+ or K

18) Do metals and non-metals tend to lose or gain electrons? Does this form a cation or an anion? Circle the right answers and use ionization energy and electronegativity to explain your answers.

Metals (gain / lose) electrons, forming (cations / anions).

Explain: It will lose electrons, as It has low ionization energys. This means it is easy for other elements to steal their electrons. They also have low electronegativities, so they will not gain electrons. Since they are losing electrons they have a positive charge, making them cations.

Non-Metals (gain / lose) electrons, forming (cations/anions).

Explain:

They have very high ionization levels, which means it is unlikely for them to lose electrons. They also have high electronegativitys, meaning it is easy for them to steal electrons. When something has an extra electron, it gains a negative charge, which is an anion.