Name: ANSWERS

/ 75

BCI SCIENCE

SCH 4U

Test:
Atomic Structure and Properties
Practice Test

Multiple Choice: *Identify the letter of the choice that best completes the statement or answers the question.(30)*

1. Which object(s) would you use to describe the shape of the 2*p* orbital?
(a) a dumb-bell
b. a circle
c. a sphere
d. two perpendicular dumb-bells
e. a doughnut
A

2. Which situation must be true for two electrons to occupy the same orbital?
a. The electrons must have the same principal quantum number, but the other quantum numbers must be different.
b. The electrons must have the same spin.
c. The electrons must have identical sets of quantum numbers.
d. The electrons must have low energy.
(e) The electrons must have the opposite spin.
E

3. An electron has the following set of quantum numbers:
$n = 3, l = 1, m_l = 1, m_s = +\frac{1}{2}$.
In which orbital is this electron found?
a. 3*s*
(b) 3*p*
c. 3*d*
d. 3*f*
e. 4*p*
B

4. Which element contains a full 3*s* orbital?
a. B
b. Na
(c) Mg
d. Be
e. Ne
C

5. Which set of quantum numbers is not possible?
a. $n = 3, l = 0, m_l = 0, m_s = \frac{1}{2}$
b. $n = 5, l = 3, m_l = 2, m_s = \frac{1}{2}$
c. $n = 4, l = 3, m_l = -1, m_s = -\frac{1}{2}$
d. $n = 5, l = 3, m_l = -3, m_s = -\frac{1}{2}$
(e) $n = 4, l = 4, m_l = 2, m_s = -\frac{1}{2}$
E

*** 6. Which set of quantum numbers is not possible?**
a. $n = 5, l = 3, m_l = 0, m_s = -\frac{1}{2}$
b. $n = 1, l = 0, m_l = 0, m_s = \frac{1}{2}$
c. $n = 3, l = 2, m_l = 1, m_s = \frac{1}{2}$
d. $n = 3, l = 3, m_l = -3, m_s = \frac{1}{2}$
e. $n = 5, l = 2, m_l = 0, m_s = -\frac{1}{2}$
D

7. Which scientist postulated that electrons can only move between certain energy levels?
a. Rutherford
b. Dalton
c. Einstein
d. Schodinger
(e) Bohr
E

8. The electron was discovered by
(a) Thomson
b. Rutherford
c. Democritus
d. Goldstein
A

9. In Rutherford's experiment, alpha particles were _____charged
a. negatively
c. neutral
(b) positively
d. both a & b
B

10. What did Heisenberg contribute to the quantum mechanical model of the atom?
(a) Uncertainty principle
b. Hund's rule
c. Aufbau principle
d. Wave equation
A

11. What did Schrodinger contribute to the quantum mechanical model of the atom?
a. Uncertainty principle
b. Hund's rule
c. Aufbau principle
(d) Wave equation
D

12. "A region of space in which there is a high probability of finding an electron" is the definition of a(n)
(a) orbital
b. photon
c. absorption spectrum
d. quanta
A

13. Which electron configuration represents a reactive non-metallic element?
(a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
b. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
c. $1s^2\ 2s^2\ 2p^6\ 3s^2$
d. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
e. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
A

14. How many *p* orbitals are in each energy level, except *n* = 1?
a. 1
(b) 3
c. 5
d. 6
e. 7
B

15. What is the maximum number of electrons in $n = 3$?

E a. 2 d. 9
b. 3 e. 18
c. 6

16. What is the total number of electrons in the 2p orbitals of a sulfur atom at ground state?

B a. 8 d. 3
b. 6 e. 2
c. 4

17. Which sublevel, when full, corresponds to the first row of transition elements?

A a. 3*d* d. 4*f*
b. 3*f* e. 4*p*
c. 4*d*

18. Which sublevel, when full, corresponds to the lanthanide series of elements?

D a. 3*d* d. 4*f*
b. 3*f* e. 5*f*
c. 4*d*

19. Which pair of atoms and/or ions is isoelectric?

B a. O^{2-} and Cl^- d. Li^+ and Na^+
b. Ca^{2+} and Cl^- e. K^+ and Kr
c. F^- and N^{2-}

20. How does atomic radius change from left to right across a period in the periodic table?

B a. It increases. d. It increases and then decreases.
b. It decreases. e. It decreases and then increases.
c. It stays the same.

21. Which element has the highest electron affinity?

D a. Li d. F
b. N e. Ni
c. O

22. Which element has the largest atomic radius?

A a. Mg d. Cl
b. Be e. Si
c. F

23. Which element has the lowest first ionization energy?

B a. Ca d. O
b. Cs e. Ba
c. Br

24. Which forces exist between hydrogen chloride, HCl, particles?

I. Van der Waals (Dispersion Forces) II. metallic bonding
III. hydrogen bonding IV. dipole

B a. I only d. I, III and IV only
b. I and IV only e. I, II and II only
c. I and II only

$\delta^+ \; +17 \; \delta^-$
$H-Cl$

ΔEN
Cl 3.16
H 2.20
0.96

25. In general, the valence electrons of metals are:

B a. few in number and strongly held c. many in number and strongly held
b. few in number and weakly held d. many in number and weakly held

26. Which statement is the best description of chlorine, Cl_2?

C a. polar molecule d. ionic compound
b. polar bonds, non polar molecule e. none of the above
c. non polar molecule

27. Which statement is the best description of potassium chloride, KCl?

D a. polar molecule d. ionic compound
b. polar bonds, non polar molecule e. none of the above
c. non polar molecule

***28. What is the specific shape of a molecule of antimony(III) fluoride, SbF_3?**

C a. linear d. tetrahedral
b. trigonal planar e. angular
c. trigonal pyramidal

N6e⁻
1 Sb x 8e⁻
3 F x 8e⁻
32 Ne⁻

Ve⁻
1 Sb x 5e⁻ = 5e⁻
3 F x 7e⁻ = 21e⁻
26 Ve⁻

# of BONDS
32 − 26 / 2 = 3 BONDS

Le⁻
26 − 6 = 20 Le⁻

29. Why are diamonds so hard?

B a. because they are made of carbon
b. because they are made of a three dimensional array of particles
c. because it is able to conduct electricity
d. because there are covalent bonds between particles
e. none of the above

30. A solid compound consists of ions bound in a crystal lattice. Which property would you not expect this solid to have?

B a. high melting point d. soluble in a polar solvent
b. good conductivity in the solid state e. brittle

31. Molecule Structure Chart. Complete the following chart.(24)

MOLECULE	e⁻ CALCULATIONS (NGe⁻, Ve⁻, Bonds, Le⁻) [/2]	LEWIS STRUCTURES (indicate co-ordinate covalent bonds, δ/δ⁺, dipole moments &/or resonance structures where appropriate) [/2]	3-D DRAWING (indicate correct bond angles) [/2]	NAME (general & Specific) [/2]
XeI_2	NGe⁻ (10e⁻ ☺) 1 Xe × 8e⁻ = 8e⁻ 2 I × 8e⁻ = 16e⁻ 24 NGe⁻ ☹ 26 NGe⁻ Ve⁻ 1 Xe⁻ × 8e⁻ = 8e⁻ 2 I × 7e⁻ = 14e⁻ 22 Ve⁻ # of bonds 24 − 22 = X ☹ 26 − 22 = 2 bonds ☺ 2 Le⁻ 22 − 4 = 18 Le⁻	ΔEN I 2.66 Xe 2.60 0.06 δ⁻ δ⁺ δ⁻ :I – Xe – I:	180° 90°	G: trigonal bipyramidal S: linear
NH_4^+	NGe⁻ 1 N × 8e⁻ = 8e⁻ 4 H × 2e⁻ = 8e⁻ 16 NGe⁻ Ve⁻ 1 N × 5e⁻ = 5e⁻ 4 H × 1e⁻ = 4e⁻ Charge = −1e⁻ 8 Ve⁻ Bonds 16 − 8 / 2 = 4 bonds Le⁻ 8 − 8 = 0 Le⁻	ΔEN N 3.04 H 2.20 0.84 co-ordinate covalent δ⁺ δ⁻ δ⁺ N ∘ H ×	+1 109.5°	G: tetrahedral S: tetrahedral
BBr_3	NGe⁻ (6e⁻ ☺) 1 B × 8e⁻ = 8e⁻ 3 Br × 8e⁻ = 24e⁻ 32 NGe⁻ ☹ 30 NGe⁻ Ve⁻ 1 B × 3e⁻ = 3e⁻ 3 Br × 7e⁻ = 21e⁻ 24e⁻ Bonds 32 − 24 / 2 = 4 bonds ☹ 30 − 24 / 2 = 3 bonds Le⁻ 24 − 6 = 18 Le⁻	ΔEN Br 2.96 B 2.04 0.92 δ⁻ δ⁺ δ⁻ δ⁻	120°	G: trigonal planar S: trigonal planar

Short Answer: *Choose a selection of questions that total out of 21 marks and write the answers on the fullscap provided. (21)*

32. Which of the following would be larger? Explain (4)

a) Mg or Mg^{+2}

- Mg is larger
- both have 12 p^+ but Mg^+ has 2 more e^- than Mg^{+2} so mg's Zeff will be less than Mg^{+2} so its valence e^- will be further from the nucleus.

b) O or O^{-2}

- O^{2-} will be larger
- both have 8 p^+ but O^{2-} has 2 more e
- so its Zeff will be less allowing its valence e^- to be further from the nucleus.

33. Explain emission spectra and absorption spectra by referring how the terms apply to Bohr's Quantum model of the atom. (4)

E.S. is a result of photons of light being released from an atom when e^- fall from higher E levels. A spectra pattern of distinct bands of light is produced which is unique to each element. A.S. is a result of photons of light being absorbed by atoms to promote e^- to higher E levels. As a result black (absence of light) bands are observed to indicate the "jump" of e^-

34. The first and second stage ionization energies of lithium and beryllium are given below:

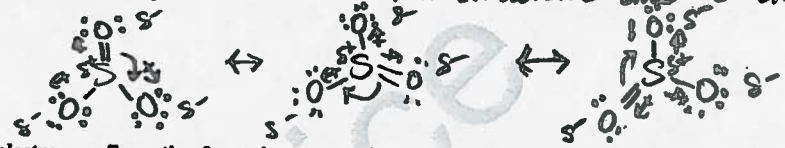
Ionization Energy (kJ/mol)	Lithium	Beryllium
First	520	899
Second	7297	1757

Explain why beryllium has the higher first ionization energy than lithium, but a lower value for the second. (4)

Be $\boxed{\uparrow\downarrow}_{1s}$ $\boxed{\uparrow\downarrow}_{2s}$ vs. Li $\boxed{\uparrow\downarrow}_{1s}$ $\boxed{\uparrow}_{2s}$ Be has full valence orbital 2s ∴ is more stable = more E required to remove e^- than half full 2s of Li (lower E require However Li^+ is isoelectric w/ He, a noble gas (very stable) so it will require a hug amount of E vs. Be^+ $2s^1$.

35. What are resonance structures? Using a SO_3 molecule draw all possible resonance structures to represent its bonding (4)

R.S. = rotation of double bond within a structure due to delocalized e^-.



36. Write the electron configuration for an iron atom with a +2 charge. Also, write the short form of the orbital diagram. Explain how the ~~Pauli Exclusion Principal~~ Hunds Rule applies to this element's electron configuration and also explain the inconsistency in the electron configuration that occurred to achieve a more stable energy arrangement. (5)

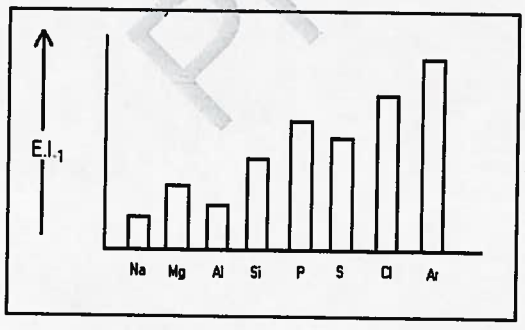
Fe^{2+} $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

[Ar] $\boxed{\uparrow\downarrow}_{4s}$ $\boxed{\uparrow\uparrow\uparrow\uparrow\uparrow}_{3d}$

Hunds Rule: each orbital must be half-filled before it can be paired up.

A 4s e^- was pulled to half-fill 3d to reach a more stable energy arrangement.

37. State the trend for ionization energy across a period. There are two exceptions to that rule in the graph below. Explain the two exceptions using Quantum Theory. (5) I.E. ↑ across a period due to an ↑ in Zeff



Mg vs Al

Mg has a full s^2 vs Al which has $2p^1$ which is less stable than Mg ∴ Mg will require more E to remove e^-

P vs S

P has half-full p orbital vs p^4 of S. ½ full p is more stable than p^4 so P e^- will require more E to remove an e^- vs S.

38. Consider the following electron configurations of neutral atoms: (5)

(i) $1s^2 2s^2 2p^6 3s^2$ (ii) $1s^2 2s^2 2p^6 3s^1$ (iii) $1s^2 2s^2 2p^5$ (iv) $1s^2 2s^2 2p^3$ (v) $1s^2 2s^2 2p^6$

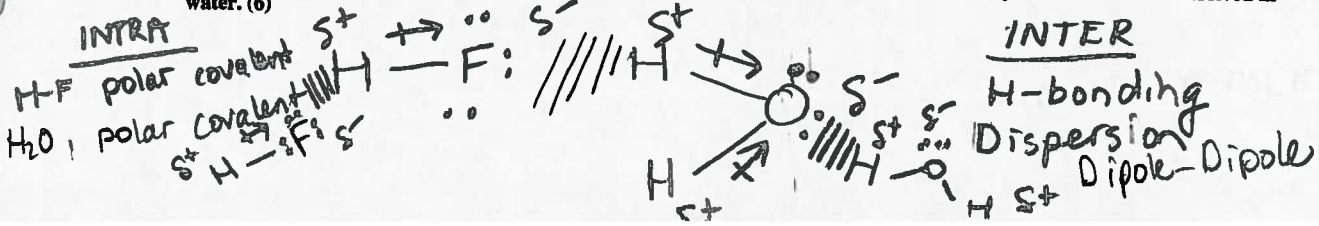
a. Which of these atoms would you expect to have the lowest ionization energy? ii
b. Which atom would you expect to be an inert gas? v
c. List the five atoms in a predicted order of increasing first ionization energies. ii, i, iv, iii, v
d. Predict the atom that should have the highest second ionization energy. ii
e. Predict the atom that should have the lowest second ionization energy. i

39. With use of a diagram, draw and name ALL intra and intermolecular forces present when hydrofluoric acid is dissolved in water. (6)

INTRA

HF polar covalent

H_2O polar covalent



INTER

H-bonding

Dispersion

Dipole-Dipole