SCH4U1 Atomic & Molecular Structure Test Review

____ 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

____ 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.

____ 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*

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

____ 5. 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. $n = 4$, $l = 3$, $m_l = -3$, $m_s = \frac{1}{2}$
e. $n = 5$, $l = 2$, $m_l = 0$, $m_s = -\frac{1}{2}$

____ 6. Which scientist postulated that electrons can only move between certain energy levels?
a. Rutherford
b. Dalton
c. Einstein
d. Schodinger
e. Bohr

____ 7. 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. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

____ 8. How many *p* orbitals are in each energy level, except $n = 1$?
a. 1
b. 3
c. 5
d. 6
e. 7

____ 9. What is the maximum number of electrons in $n = 3$?
a. 2
b. 3
c. 6
d. 9
e. 18

____ 10. What is the total number of electrons in the 2*p* orbitals of a sulfur atom at ground state?
a. 8
b. 6
c. 4
d. 3
e. 2

____ 11. Which sublevel, when full, corresponds to the first row of transition elements?
a. 3*d*
b. 3*f*
c. 4*d*
d. 4*f*
e. 4*p*

____ 12. Which pair of atoms and/or ions is isoelectronic?
a. O^{2-} and Cl^{-}
b. Ca^{2+} and Cl^{-}
c. F^{-} and N^{2-}
d. Li^{+} and Na^{+}
e. K^{+} and Kr

____ 13. How does atomic radius change from left to right across a period in the periodic table?
a. It increases.
b. It decreases.
c. It stays the same.
d. It increases and then decreases.
e. It decreases and then increases.

____ 14. Which element has the highest electron affinity?
a. Li
b. N
c. O
d. F
e. Ni

____ 15. Which metal is the most reactive?
a. Al
b. K
c. Cu
d. Zn
e. Ca

____ 16. Which element has the largest atomic radius?
a. Mg
b. Be
c. F
d. Cl
e. Si

____ 17. Which element has the lowest first ionization energy?
a. Ca
b. Cs
c. Br
d. O
e. Ba

____ 18. What is the bond angle in a bent molecule, such as water?
a. 90°

b. 104.5°
c. 107.3°
d. 109.5°
e. 120°

____ 19. What is the bond angle in a trigonal pyramidal molecule?
a. 90°
b. 104.5°
c. 107.3°
d. 109.5°
e. 120°

____ 20. What is the bond angle in a trigonal planar molecule?
a. 90°
b. 104.5°
c. 107.3°
d. 109.5°
e. 120°

____ 21. Which compound has polar covalent bonds?
a. $AgCl_2$
b. CH_4
c. Cl_2
d. CF_4
e. B_2H_8

____ 22. Which compound is truly covalent?
a. SO_2
b. MgO
c. NH_3
d. PCl_3
e. P_2S_3

____ 23. Which molecule has a molecular dipole?
a. CCl_4
b. NF_3
c. BeF_2
d. CF_4
e. CO_2

____ 24. Which bond is most polar?
a. C—O
b. C—N
c. B—O
d. B—N
e. S—O

Short Answers/Problems: *For the following questions, write the most appropriate answer in the space provided.*

25. Describe and explain the general trends in atomic radius in the periodic table.

26. What contributions did Schrödinger make to the model of the atom?

27. What contributions did Planck and Einstein make to the current model of the atom?

28. What are the four quantum numbers in the quantum mechanical model of the atom? What do these numbers represent?

29. Explain what the aufbau principle is and how it is used.

30. Explain Hund’s rule and the Pauli exclusion principle. Give an example to show how these two rules are used.

31. What was one problem that the Bohr model of the atom could not explain?

32. What are the allowed values of m_l for an electron with each orbital-shape quantum number.
a) $l = 3$ b) $l = 1$

33. What are the allowed values of l for an electron with each principal quantum number.
a) $n = 4$ b) $n = 6$

34. What is the relationship between atomic radius and ionization energy? Explain your answer.

35. What are the allowed values of l and m_l if $n = 2$? What is the total number of orbitals in this energy level?

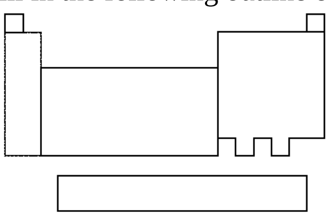
36. List the different types of intermolecular forces, and give an example for each.

37. What are resonance structures? Give an example of a molecule that requires resonance structures to represent its bonding, and draw these resonance structures.

38. Explain the difference between covalent bonding and polar covalent bonding, using an example of each.

39. Using a diagram, explain Rutherford's gold foil experiment. What did this experiment prove?

40. Fill in the following outline of a periodic table to show the four energy sublevel (*s*, *p*, *d*, and *f*) blocks.



41. Sketch the shapes of the three different *p* orbitals on a three-dimensional axis.

42. On a sketch of a periodic table, use arrows to indicate the directions of increasing
a) atomic radius b) ionization energy c) electron affinity.

43. Create a table to summarize the differences in the physical properties of the different types of solids (atomic, molecular, covalent network, ionic, and metallic).

44. The ionization energies of a given atom are IE_1 = 800 kJ/mol, IE_2 = 2400 kJ/mol, IE_3 = 3700 kJ/mol, IE_4 = 25 000 kJ/mol, and IE_5 = 32 800 kJ/mol. Predict the valence electron configuration for the atom, and explain your reasoning.

45. Explain what is wrong with each set of quantum numbers.
a) $n = 3$, l = 3, $m_l = -2$; name: 3*d* b) $n = 5$, l = 3, $m_l = 5$; name: 5*f*

46. Fill in the missing value(s) in each set of quantum numbers.
a) n = ?, $l = 1$, $m_l = -1$; name: 3*p* b) $n = 1$, l = ?, m_l = ?; name: 1*s*

47. What is the name of the orbital that is represented by each set of quantum numbers?
a) $n = 1$, $l = 0$, $m_l = 0$ b) $n = 5$, $l = 3$, $m_l = -3$

48. Write the electron configuration for each element or ion.
a) P b) Te^{2-} c) U

49. Write the condensed electron configuration for each element or ion.
a) Bi b) Zn^{2+} c) Al

50. Draw the orbital diagram for each element or ion.
a) Fe b) Na^+ c) Cl

51. What types of intermolecular forces affect each compound or mixture? Explain your reasoning.
a) HF b) C_6H_{12} c) N_2 dissolved in H_2O d) MgI_2

52. For each molecule,
- draw the Lewis structure
- use the VSEPR theory to predict the shape of the compound
- decide whether or not the molecule is polar
- determine the bond angles
- draw the dipole movement for the molecule if the molecule is polar

a) CS_2 b) PCl_3 c) SiH_4 d) NH_4^+ e) $AsCl_5$
f) ICl g) CH_2Cl_2 h) PH_3 i) SeI_6

Bonding & Atomic Theory Review
Answer Section

MULTIPLE CHOICE

1. ANS: A DIF: easy
2. ANS: E DIF: easy
3. ANS: B DIF: easy
4. ANS: C DIF: easy
5. ANS: D DIF: average
6. ANS: E DIF: easy
7. ANS: A DIF: average
8. ANS: B DIF: easy
9. ANS: E DIF: easy
10. ANS: C DIF: average
11. ANS: A DIF: easy
12. ANS: B DIF: average
13. ANS: B DIF: easy
14. ANS: D DIF: average
15. ANS: B DIF: average
16. ANS: A DIF: average
17. ANS: B DIF: average
18. ANS: B DIF: easy
19. ANS: C DIF: easy
20. ANS: E DIF: easy
21. ANS: D DIF: average
22. ANS: E DIF: average
23. ANS: B DIF: average
24. ANS: C DIF: easy

SHORT ANSWER

25. ANS:

Atomic radius increases as you go down a group of elements in the periodic table. This trend is a result of increasing numbers of electrons occupying increasing numbers of energy levels. The effective nuclear charge changes only slightly and therefore does not offset the increase in size due to the increase in energy levels.

Atomic radius decreases as you go left to right across a period in the periodic table. The valence electrons are found in orbitals of the same energy level. At the same time, the effective nuclear charge is increasing with the increase in nuclear charge, which results in a greater force of attraction pulling the valence electrons closer to the nucleus. Thus, atomic size decreases.

26. ANS:

Schrödinger used mathematics and statistics to combine de Broglie's idea of matter waves and Einstein's idea of quantized energy particles. Schrödinger's mathematical equations and their interpretations, together with Heisenberg's uncertainty principle, resulted in the birth of the field of quantum mechanics.

27. ANS:

Planck proposed that matter at the atomic level can absorb or emit only discrete quantities of energy. In other words, Planck said that the energy of the atom is quantized. Einstein proposed that all forms of electromagnetic energy travel as photons of energy.

28. ANS:

1) The principal quantum number, *n*, indicates the energy level of an atomic orbital and its relative size.
2) The orbital-shape quantum number, *l*, indicates the shape of the orbital.

3) The magnetic quantum number, m_l, indicates the orientation of the orbital.
4) The spin quantum number, m_s, indicates the direction in which the electron is spinning.

29. ANS:
The aufbau principle is the imaginary process of building up the ground state electron structure for each atom, in order of atomic number. When determining the electron configuration of an element, the electrons are written sequentially in orbitals of increasing energy, starting with the electron in the 1*s* orbital.

30. ANS:
The Pauli exclusion principle states that no two electrons in an atom have the same four quantum numbers. (In other words, no two electrons can occupy the same orbital with the same spin.) For example, boron's electron configuration is $1s^2\ 2s^2\ 2p^1$.
Hund's rule states that whenever electrons are added to orbitals of the same energy sub-level, each orbital receives one electron before any pairing occurs. When electrons are added singly to separate orbitals of the same energy sublevel, the electrons must all have the same spin. For example, the electron configuration of nitrogen is $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$.

31. ANS:
The Bohr model successfully explained only one-electron systems. It was unable to explain the emission spectra for atoms with two or more electrons.

32. ANS:
a) $m_l = -3, -2, -1, 0\ 1, 2, 3$
b) $m_l = -1, 0, 1$

33. ANS:
a) $l = 0, 1, 2, 3$
b) $l = 0, 1, 2, 3, 4, 5$

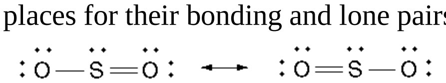
34. ANS:
Ionization energy is the energy that is required to remove an electron completely from a ground state gaseous atom. This energy tends to increase as the atomic radius decreases. The closer the electrons are to the nucleus, the great the force of attraction pulling or holding the electrons in the atom.

35. ANS:
If $n = 2$, then $l = 0, 1$. For $l = 0$, $m_l = 0$. For $l = 1$, $m_l = -1, 0, 1$. There are four orbitals in this energy level.

36. ANS:

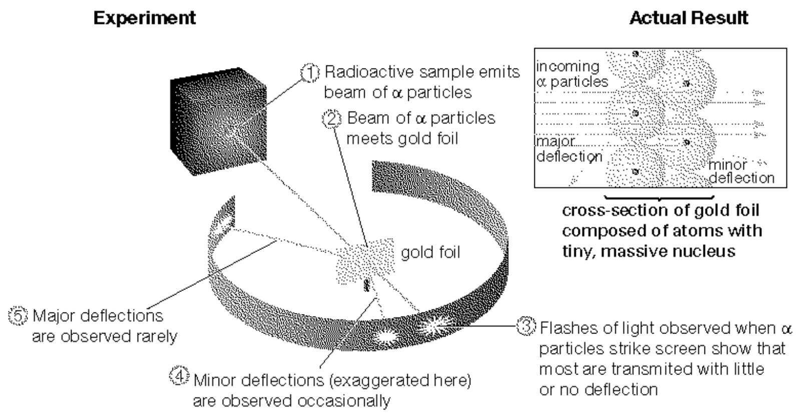
Intermolecular forces	Example
ion-dipole	sodium ions in water
hydrogen bonds	water
dipole-dipole	iodine monochloride
ion-induced dipole	ferrous ions and oxygen molecules
dipole-induced dipole	hydrochloric acid and chlorine
dispersion (London) forces	fluorine gas

37. ANS:
Resonance structures are models that give the same relative position of atoms as Lewis structures, but show different places for their bonding and lone pairs. An example is SO_2.



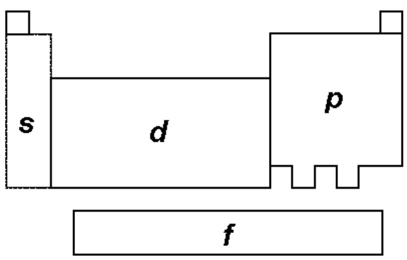
38. ANS:
Covalent bonding exists when the bonding electrons are shared equally or nearly equally, as in molecules of O_2 and N_2. Polar covalent bonding exists when there is unequal sharing of a pair of electrons between two atoms, as in HCl and H_2O.

39. ANS:

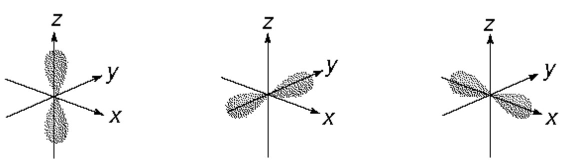


This experiment proved that the atom is made up mainly of empty space, with a small, massive region of concentrated charge at the centre. The charge was soon determined to be positive.

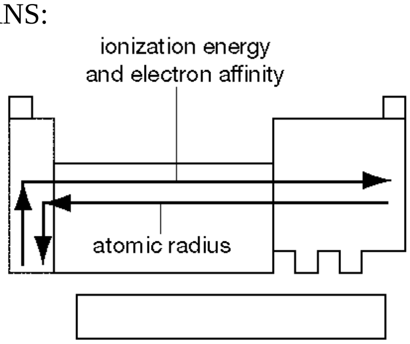
40. ANS:



41. ANS:



42. ANS:



43. ANS:

Type of Crystal Solid	Boiling Point	Electrical Conductivity in Liquid State	Other Physical Properties of Crystals
Atom	low	very low	very soft
Molecular	generally low (non-polar); intermediate polar	very low	non-polar: very soft; soluble in non-polar solvents
			polar: somewhat hard, but brittle; many are soluble in water

Covalent Network	very high	low	hard crystals that are insoluble in most liquids
Ionic	high	high	hard and brittle; many dissolve in water
Metallic	most high	very high	all have a lustre, are malleable and ductile, and are good conductors; they dissolve in other metals to form alloys

44. ANS:
The atom probably has three valence electrons because the increase from the third to the fourth ionization energies is much greater than the increase from the first to the second to the third ionization energies. This large increase can be explained by the fact that the fourth electron is being removed from a filled energy level, which is closer to the nucleus. Thus, the outer (valence) electrons have all been removed.

45. ANS:
a) l cannot equal 3, since the maximum value of l is $(n - 1)$ and d orbitals have an l value of 2.
b) m_l cannot equal 5, since the maximum value of m_l is $+l$.

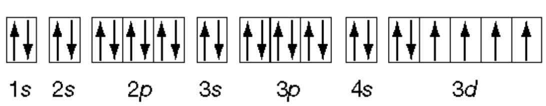
46. ANS:
a) $n = 3$
b) $l = 0$, $m_l = 0$

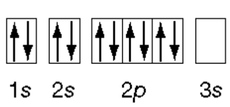
47. ANS:
a) 1s
b) 5f

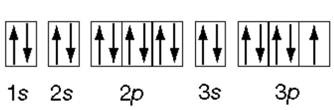
48. ANS:
a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6$
c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2\ 5f^4$

49. ANS:
a) [Xe] $6s^2\ 4f^{14}\ 5d^1\ 6p^3$
b) [Ar] $4s^2\ 3d^{10}$
c) [Ne] $3s^2\ 3p^1$

50. ANS:

a) 

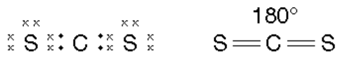
b) 

c) 

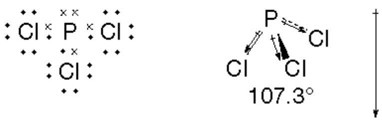
51. ANS:
a) Hydrogen bonds: The bonding of hydrogen to the small electronegative fluorine atom results in the charge distribution that allows the formation of hydrogen bonds between molecules.
b) Dispersion forces: The C_6H_{12} molecule has covalent bonds between atoms.
c) Dipole-induced dipole: The H_2O molecule has a bent shape, so it has a molecular dipole. This dipole induces a dipole in the N_2 molecules, resulting in their mutual attraction.

d) Ion-dipole interactions. MgI_2 dissociates into its ions as it dissolves in H_2O. The H_2O molecule has a molecular dipole. Hence, ion-dipole interaction takes place between the ions of MgI_2 and the polar H_2O molecules.

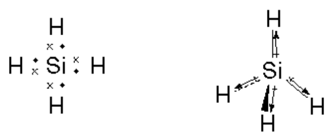
52. ANS:

a)

linear

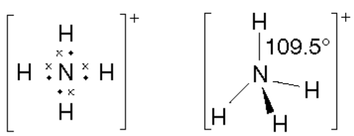
non-polar molecule

b)

trigonal planar

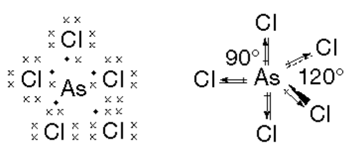
polar molecule

c)

tetrahedral

non-polar molecule

a)

tetrahedral

charged ion

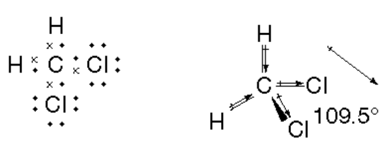
b)

trigonal bypyramidal

non-polar molecule

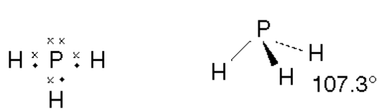
c)

linear

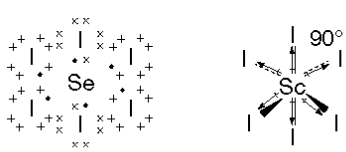
polar molecule

a)

tetrahedral

polar molecule

b)

trigonal pyramidal

non-polar molecule

c)

octahedral

non-polar molecule