Chemistry 12

Solutions Manual Part A

Unit 2 Structures and Properties of Matter

Answers to Unit 2 Preparation Questions

(Student textbook pages 158-61)

1. a. Poisonous and infectious material causing immediate and serious toxic effects

b. Use care when handling hot objects

2. Tie back long hair. Do not have any flammable material near the burner. Secure the burner. For example, attach it to a retort stand. Handle hot glassware with thermal gloves or beaker tongs.

3. The sources must be reliable. Web sites such as university and government sites are most reliable. If the source is posted by an individual, the credentials of the author must be shown to be reliable. Corporate sites are often biased.

4. a

5. Note to the teacher: The numbers in the first printing of the text were accidentally combined. The values in the shaded squares are the data that should be provided and the open squares are the answers. BLM 3-1 has been provided with the correct working chart.

Name of Isotope	Notation for Isotope	Atomic Number	Mass number	Number of Protons	Number of Electrons
bromine-81	$^{81}_{35}Br$	35	81	35	35
neon-22	$^{22}_{10}Ne$	10	22	10	10
calcium-44	$^{44}_{20}Ca$	20	44	20	20
silver-107	$^{107}_{47}Ag$	47	107	47	47

6. Valence electrons are those electrons that exist in the outer energy level. They are the only electrons that are involved in chemical bonds.

7. b

8. c

9. The first four electrons, in a Lewis diagram, are drawn as far apart as possible because their negative charges repel each other.

10. a. The distance from the centre of the atom to the boundary within which the electrons spend 90 percent of their time.

b. The energy absorbed or released when an electron is added to a neutral atom.

c. An indicator of the relative ability of an atom to attract shared electrons.

d. The amount of energy required to remove the outermost electron from an atom or ion in the gaseous state.

11.

a. down a group:	b. across a period:
a. increases	a. decreases
b. decreases	b. increases
c. decreases	c. increases
d. decreases	d. increases

12. a. yellow: halogens; very reactive non-metals

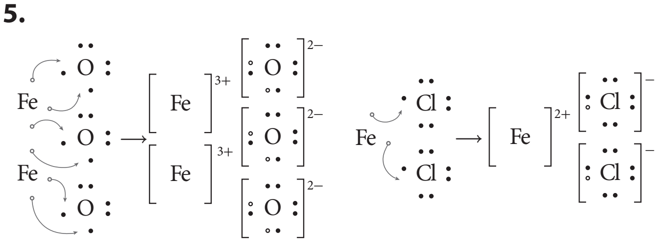
b. purple: transition metals; very hard metals with high boiling points

c. brown; other non-metals of the main group elements; most of these atoms are essential building blocks of molecules found in living organisms. Their chemistry typifies how non-metals combine with other atoms.

13. In both cases, the metal is donating, in one case 1 e^- and in the other 2 e^-, to a non-metal. This leaves a positive metal ion and a negative non-metal ion that are attracted to each other by electrostatic forces and they form an ionic compound.

14. The octet rule states that when bonds form between atoms, the atoms gain, lose, or share electrons in a way that will create an outer shell containing eight (an octet) of electrons. However, the electron configurations of the transition elements are such that they can have more than eight electrons in their outer shell.

15.



16. d

17. Both a lone pair and a bonding pair consist of two paired electrons in the valence energy level of an atom. The lone pair is **not** involved in a covalent bond in the molecule but may interact with a nearby molecule forming an intermolecular bond. A bonding pair consists of 1 electron from the valence energy level of two separate atoms forming a covalent bond between two atoms within the molecule.

18. Ionic: a., d., and e.

Molecular: b. and c.

19. a. ammonium ion

b. cyanide ion

c. chromate ion

d. carbonate ion

e. nitrite ion

20. Titanium can form ions having more than one charge. The Roman numeral IV indicates that the ion has lost four electrons and thus carries a charge of 4^+.

21. a. ionic: iron(III) oxide

b. molecular: phosphorus pentaiodide

c. molecular: dinitrogen monoxide

d. ionic: calcium iodide

e. ionic: sodium cyanate

22. a. ionic: $ZnBr_2$

b. molecular: SF_6

c. ionic: $CuCrO_4$

d. molecular: Cl_2O

e. ionic: CsCl

23.

Number	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

24. chlorine, methanol, potassium oxide:
Ionic compounds such as potassium oxide form crystals with tightly bound positive and negative ions and thus have very high boiling points. Alcohols such as methanol have hydroxyl groups that can form hydrogen bonds among the molecules. Although hydrogen bonds are not nearly as strong as the bonds between ions, they nevertheless increase the boiling point above those of non-polar compounds. Chlorine forms a non-polar molecule. The only attractive forces among the molecules are very weak and therefore non-polar compounds have very low boiling points.

25. c

26. When ionic compounds are dissolved in water, the ions are separated from each other and are free to move independently. Positive ions are attracted to the negative electrode and negative ions are attracted to the positive electrode. The moving ions constitute an electric current. When the ions reach their respective electrode, the negative ions donate electrons to the positive electrode and the positive ions accept electrons from the negative electrode. The electrons move through the external circuit and form a complete loop of electric charge flow.

Each molecule in a molecular compound has an overall neutral charge and thus they do not move towards any electrode. There is no movement of charges.

27. a. Because the chlorine atoms are more electronegative than carbon atoms, the electrons in the carbon-chlorine bond spend more time near the chlorine atoms thus making the bonds polar.

b. Because the CCl_4 molecule is perfectly symmetrical, the polarity of the bonds cancel each other and the overall molecule is not polar.

c. Because the CCl_4 molecules are non-polar, they would likely not dissolve in water.

28. In a sodium chloride crystal, each sodium ion is attracted to six adjacent chloride ions. The distances between the ions and therefore the strength of the attraction, is the same in every case. The same can be said for the chloride ions, each of which is attracted equally to six adjacent sodium ions. Since all of these attractive forces are the same, it is not possible to identify any specific pair of sodium and chloride ions that form a "molecule."

29. Intermolecular forces are those forces that attract molecules to one another. Covalent bonds are the interactions between atoms in a molecule, holding the individual atoms together to form a molecule.

30. Each elliptical shape represents a polar molecule or a dipole. The oppositely charged ends of the dipoles attract each other with electrostatic forces. This type of attraction is called a dipole-dipole force.

Chapter 3 Atomic Models and Properties of Atoms

Answers to Learning Check Questions

(Student textbook page 168)

1. Thomson's discovery of the electron in 1897 invalidated Dalton's atomic theory.
2. Alpha particles would have passed straight through the foil with minimal or no deflection from encounters or collisions with nearby electrons. There would be no deflection caused by the positive charge because Thomson's model postulates a uniform, positive charge spread throughout the atom.
3. Some radioactive elements emit positively charged alpha particles. Rutherford studied them and then used the alpha particles to bombard thin foils including gold foils. This led to the model in which all of the positive charge and most of the atom's mass were confined to a very small region at the centre of the atom, which Rutherford called the nucleus.
4. Diagrams should be based on Figures 3.3 and 3.6. Both models are spherical and include electrons and the positive charge. In Thomson's model, the positive charge is spread throughout the sphere and electrons are embedded in the sphere like raisins in a muffin. In Rutherford's model, the positive charge is found in a tiny, extremely dense nucleus and the electrons orbit the nucleus like planets.
5. Scientists tend to name their models, or other discoveries, after something that is common to their own everyday lives. Rutherford's model is sometimes called the planetary model.
6. In Thomson's model, negative charges were scattered evenly throughout a large positively charged mass. The alpha particles were highly energetic and would not be expected to be deflected very much by such atoms.

(Student textbook page 183)

7. In a hydrogen atom, orbital energy depends only on *n*. For example, electrons in 2*s* and 2*p* have the same energy. In multi-electron atoms, orbitals in different sublevels have different energies associated with them, even if they have the same value of *n*. For example, 2*s* and 2*p* are associated with different energies.
8. **a.** 2*s*, 2*p*, 3*s*, 3*p*
   b. 3*p*, 4*s*, 3*d*, 4*p*
   c. 5*s*, 4*d*, 5*p*, 6*s*, 4*f*, 5*d*, 6*p*, 5*f*
9. An orbital is "full" when it contains two electrons.
10. Method one: There are five possible orbitals for $n = 1$ and $n = 2$: one 1*s* orbital, one 2*s* orbital, and three 2*p* orbitals. Each of these can contain a maximum of two electrons. Therefore, 10 electrons can occupy all possible orbitals with $n = 1$ and $n = 2$. Method two: Using the formula $2n^2$, $n = 1$ can contain two electrons and $n = 2$ can contain eight electrons, for a total of 10.
11. No. Two arrows pointing in the same direction would indicate that two electrons in the same orbital have the same spin quantum number. This violates the statement made in the Pauli exclusion principle that no two electrons can have the same four quantum numbers.
12. **a.** $\boxed{}$ **b.** $\boxed{\uparrow}$ **c.** $\boxed{\uparrow\downarrow}$

(Student textbook page 187)

13. Orbitals fill in order of increasing energy. At energy levels above $n = 3$, the different sublevels overlap. As a result, the 5*s* orbital has a lower energy than the 4*d* orbitals.
14. boron: $1s^22s^22p^1$; $[He]2s^22p^1$
    neon: $1s^22s^22p^6$; $[He]2s^22p^6$
15. $1s^22s^22p^63s^23p^4$
16. $\underset{3s}{\boxed{\uparrow\downarrow}}\ \underset{3p}{\boxed{\uparrow}\boxed{\uparrow}\boxed{}}$
17. **a.** sodium: $1s^22s^22p^63s^1$
    b. vanadium: $1s^22s^22p^63s^23p^64s^23d^3$
18. titanium: $[Ar]4s^23d^2$

Answers to Caption Questions

Figure 3.3 (Student textbook page 165): These descriptions act as models for real structures that are complex or incapable of being seen. Each part of the model is related to a part of the real structure (for example, raisins represent electrons), so that it is easier to visualize the whole structure and its component parts.
Figure 3.16 (Student textbook page 182): 4*s* is higher since it has the higher principal quantum number (even though it is filled before the 3*d*).