First and Last Name___________________________________

Chemistry for Accelerated Students Quiz 10: Chapter 5 (5.1)

Point values for each item are shown in parentheses. Questions on this quiz may be answered a few sentences. Write all responses in complete sentences using correct spelling and grammar.

1. State the law of conservation of mass in chemical reactions. (9)

2. Balance the following equations: (7 each)

   a. $P + O_2 \rightarrow P_2O_3$

   b. $CuO + HCl \rightarrow CuCl_2 + H_2O$

   c. $Pb(NO_3)_2 + KCl \rightarrow PbCl_2 + KNO_3$

   d. $Al + HCl \rightarrow AlCl_3 + H_2$

   e. $FeS + HBr \rightarrow FeBr_2 + H_2S$

   f. $C_3H_7CHO + O_2 \rightarrow CO_2 + H_2O$

3. Determine the oxidation state for the indicated element in each of the following: (7 each)

   a. carbon in CO_2

   b. sulfur in $MgSO_4$

   c. boron in B_2O_3

   d. sodium in $NaHCO_3$

   e. nitrogen in KNO_3

   f. oxygen in H_2SO_4

   g. hydrogen in H_3PO_4

First and Last Name___________________________________

Chemistry for Accelerated Students Quiz 15: Chapter 7 (7.2)

Point values for each item are shown in parentheses. Questions on this quiz may be answered a few sentences. Write all responses in complete sentences using correct spelling and grammar.

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1. A 9.75-L cylinder contains 22.15 mol of nitrogen at 25.0°C. What is the gas pressure in the cylinder? State your answer in atmospheres.

2. A weather balloon is partially filled with 0.285 m^3 of helium at 23.0°C and 757 Torr. Assuming the volume of gas can expand freely in the balloon, what will be the volume when the balloon is at an altitude where the temperature is –45°C and the pressure is 0.640 atm?

3. What is an "ideal gas"?

R Value, with Units	Units for P, V
$8.314\ \frac{\mathrm{L \cdot kPa}}{\mathrm{mol \cdot K}}$	kPa, L
$0.08206\ \frac{\mathrm{L \cdot atm}}{\mathrm{mol \cdot K}}$	atm, L
$8{,}314\ \frac{\mathrm{L \cdot Pa}}{\mathrm{mol \cdot K}}$	Pa, L
$8.314\ \frac{\mathrm{J}}{\mathrm{mol \cdot K}}$ *	Pa, m^3
$62.36\ \frac{\mathrm{L \cdot Torr}}{\mathrm{mol \cdot K}}$	Torr, L
$62.36\ \frac{\mathrm{L \cdot mm\ Hg}}{\mathrm{mol \cdot K}}$	mm Hg, L
$8.314 \times 10^{-5}\ \frac{\mathrm{m^3 \cdot bar}}{\mathrm{mol \cdot K}}$	bar, m^3
$8.206 \times 10^{-5}\ \frac{\mathrm{m^3 \cdot atm}}{\mathrm{mol \cdot K}}$	atm, m^3

These are the MKS units.

First and Last Name__________________________________

Chemistry for Accelerated Students Quiz 18: Chapter 9 (9.1–9.2.3)

Point values for each item are shown in parentheses. Questions on this quiz may be answered a few sentences. Write all responses in complete sentences using correct spelling and grammar.

1. State four properties of acids and four properties of bases.

2. State the names of the following acids:

$HClO_3$	H_2SO_4
H_3PO_4	HNO_3
HBr	$HClO$
H_2CO_3	H_2S

3. State the definitions of *acid* and *base* according to the theories of Arrhenius, Brønsted-Lowry, and Lewis.

4. Explain why water is amphiprotic under the Brønsted-Lowry acid-base theory.

First and Last Name________________________________

Chemistry for Accelerated Students Quiz 26: Chapter 13 (13.1–13.2)

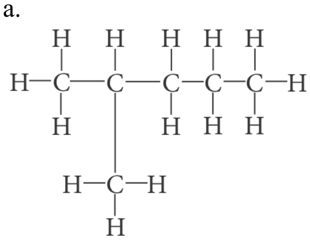
Point values for each item are shown in parentheses. Questions on this quiz may be answered a few sentences. Write all responses in complete sentences using correct spelling and grammar.

1. Explain what allotropes are and give three examples. (20)

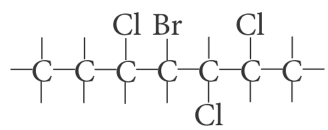
2. Why do hydrocarbon molecules have zig-zag shapes? (10)

3. Name the following molecules: (10 each)

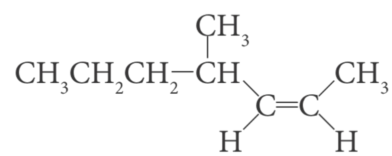
a.



b.



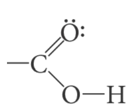
c.



4. Describe two major differences between aliphatic *cyclo*– hydrocarbons, such as cyclohexane, and aromatic hydrocarbons, such as benzene. (20)

5. Identify the following functional groups: (5 each)

a.



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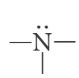
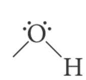


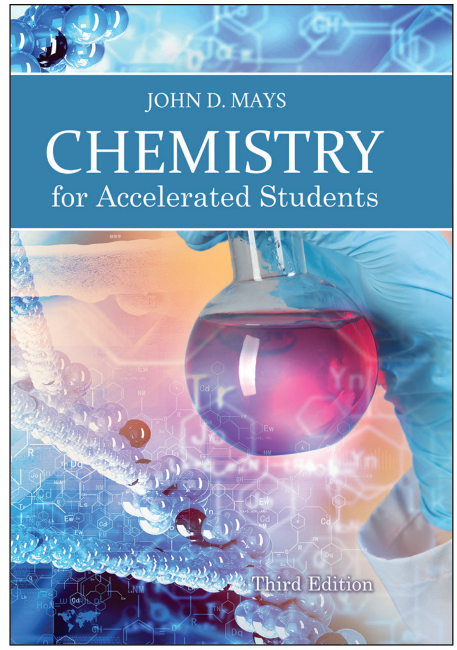
c.



d.

Chemistry for Accelerated Students

All Keys and Sample Answers



Thank you for using *Chemistry for Accelerated Students*. This supplementary document is being provided to aid in situations in which the adult teacher responsible for conducting the course does not possess a background in this subject, or in which a student is studying independently. The answers here are only samples and should not be considered the only correct response to the question.

In environments where there are multiple students in a class or group, it is the recommended method that the student should form their own answers in complete sentences as a homework assignment. These should be graded for completion only, not accuracy. Then in the group setting, students bring their preliminary answers to class where they collaborate with each other and the teacher and improve their answers. The final product will be a useful study tool developed by the group. In such an arrangement, there would be no need for this document, but it is provided for the many home study situations in which there is no collaborative group.

Additional information about how this course should be conducted is provided in the textbook introduction and in documents in the Digital Resources. A full presentation of strategies and techniques for mastery-learning can be found in our book *From Wonder to Mastery*, available from our website.

Thank you!

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Chemistry for Accelerated Students: All Keys and Sample Answers

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Contents

NOTE: Full solutions to the computations in the chapter exercises are available in the *Solutions Manual to Accompany Chemistry for Accelerated Students*. To procure a copy, visit novaresienceandmath.com.

Chapter Exercises

Introduction Exercises

1. Write five brief paragraphs summarizing the main ideas behind the titles of Sections I.1.1 through I.1.5.

Chemistry is all about electrons: Chemical reactions involve interactions between the electrons of atoms. The cloud-like regions where an atom's electrons are are called orbitals, and electrons reside in particular orbitals according to how much energy they have. The orbitals are grouped into energy groupings called shells. Atoms seek to gain, lose, or share electrons so that they have only full shells. Electronegativity also plays a role in electron position. Some atoms attract their electrons more strongly than others (higher electronegativity). This creates polar regions in molecules and leads to attractions and repulsions between molecules.

Chemistry is all about electrical forces: Some forces are stable, like the electrical attractions hold together the atoms in crystals. Others are less stable. Polarity in molecules creates attractions and repulsions between molecules, of which hydrogen bonding is the supremely important example. Hydrogen bonding in water is responsible for many of the unique properties of water.

Chemistry is all about minimizing energy: Two examples of energy minimization are a) a ball in a valley—the ball remains in the valley because its energy there is lower than it would be on the sides of the valley. If a tunnel opens up for the ball to move to another valley lower down, it goes there. b) when released, an inverted cone falls down on its side. This is a lower energy state for a cone than being inverted. In chemical reactions, atoms might be given an energy kick by a flame or other source of energy to get them out of their "energy valley". Then the atoms are free to interact with each other. When they do, energy is released and the atoms combine together to form compounds where the atoms are in lower energy states. Entropy is a measure of the disorder in a system, and in natural processes, entropy always increases. Sometimes the goals of minimizing energy and maximizing entropy pull a system in opposite directions. Such processes go in the most favorable direction. Also mentioned is the fact that Einstein first proposed that the energy atoms possess is quantized.

Chemistry is all about whole-number ratios of atoms: In molecules and in crystals, atoms always combine together in consistent-whole number ratios. This is a powerful computational tool, and allows us to compute the amount of one compound that reacts with another compound.

Chemistry is all about modeling: Since chemistry involves atoms and electrons we cannot see, we must model chemical processes. This is the case in general with science: science is the process of developing mental models, and we call these models theories. The Cycle of Scientific Enterprise is a model of how scientific knowledge works. According to this model, theories are formed to account for known scientific facts in a unified explanatory framework. A scientific fact is statement supported by a lot of evidence that is correct so far as we know, but which can change. From theories, hypotheses are formed. A hypothesis is a specific prediction about what will happen in certain circumstances. Hypotheses are tested in experiments, and the outcome of an experiment (whether positive or negative) adds to the body of known scientific facts. If the hypothesis is confirmed, the theory it came from is strengthened. If disconfirmed, the theory is weakened.

2. Describe two examples, other than those in this Introduction, of a system of some kind spontaneously (without help) moving from a higher energy state to a lower energy state. (Hint: If energy is being released, it means the entities involved in the process are moving to a lower energy state.)

Examples include anything falling or going downhill, any chemical reaction in which light or heat is released, any mechanical system moving from a stretched or compressed state to a relaxed state.

3. Describe two examples, other than those in this Introduction, of a system that will move to a lower energy state if allowed to, but which needs an initial boost of energy to get started (like the ball in Figure I.5 being kicked and then having enough energy to get out of the valley).

Example answers:

1. Popping a balloon with a pin.

2. A cart that can roll downhill but which is stuck, and a small push gets it started.

3. Water that can drain from a container through a hole in the bottom, but the hole is plugged up. Applying a momentary higher pressure to the container can unplug the hole and allow the water to drain out.

4. Describe two examples of processes in which entropy decreases. In each case, describe what source of energy and/or intelligence must be present for the decrease in entropy to occur. Here is an example to assist your thinking: an oxygen tank contains pressurized oxygen gas. The oxygen in this tank is more ordered than the oxygen in air because it has been separated from the air; there is a boundary (the tank) between the oxygen and the air. And if the valve on the tank is opened, the oxygen flows out into the air to increase the entropy (disorder). What we will never see: opening the tank valve and oxygen atoms from the atmosphere spontaneously flow into the tank. But the oxygen is put into the tank somehow, and the process that put it there decreases the entropy of that oxygen.

Example answers:

1. A child stacking blocks. The child is supplying the energy and intelligence to position the blocks in the stack.

2. Building a house. The workers supply the intelligence and some of the energy. Other energy is supplied by various machines and vehicles.

5. What is the ratio of nitrogen atoms to hydrogen atoms in ammonia molecules? What is the ratio of hydrogen atoms to carbon atoms in propane molecules?

In ammonia the N:H ratio is 1:3. In propane, the H:C ratio is 8:3.

6. Why are water molecules polar and what is the significance of this fact?

They are polar because the electronegativity of oxygen is higher than that of hydrogen. The result is that water molecules are polar. This means they stick to each other and to other polar or electrically charged objects. This gives water many of its unique properties, such as the fact that it expands just before freezing so that ice floats.

7. If oppositely charged objects attract, why can't a free electron and a free proton collide into one another and stick together because of their opposite charges?

Energy is quantized, so an electron attached to a proton possesses a specific amount of energy and will go into an orbital for that specific amount of energy.

8. A hydronium ion is a water molecule that has gained an extra proton. (A proton is identical to a hydrogen ion.) Hydronium ions form spontaneously in water, and are formed in greater quantities any time an acid is poured into water. What is the ratio of hydrogen atoms to oxygen atoms in hydronium ions?

The H:O ratio is 3:1.

9. What is hydrogen bonding?

The example of hydrogen bonding discussed here is the attraction of the position H ends of water molecules to the negative O region in the center of the molecule. This attraction makes water molecules stick together in the liquid state. It also make water molecules arrange into hexagonal patterns when they enter the solid state.

10. Distinguish between endothermic and exothermic processes.

In an endothermic process, energy is absorbed. In an exothermic process energy is released. Fires and explosions are exothermic. Photosynthesis is an endothermic process where energy from the sun is absorbed to make the process occur.

11. In a previous course, you may have learned about the "gold foil experiment" conducted by Ernest Rutherford in 1909. (I describe this experiment in Chapter 2.) This experiment led Rutherford to propose that the positive charge in atoms is concentrated in a tiny nucleus in the center of the atom. Think about this experiment and explain why Rutherford had to depend on inference as he interpreted his experimental data.

First, no one can see atoms or the charged particles in atoms. Any time we are trying to explain processes involving things we cannot see or feel we must depend on inference. Second, at the time of this experiment very little was known about atoms. Rutherford had J.J. Thomson's atomic model in mind, in which the positive charge is distributed evenly throughout the atom. Rutherford's high-velocity alpha particles should have blown right through atoms of gold if the positive charge was structured this way. This didn't always happen, which forced Rutherford to come up with a new atomic theory based on inference from his data.

12. Why doesn't oil dissolve in water?

Water molecules are polar, oil molecules are not. Water molecules are attracted to each other and any nonpolar molecules around will simple get squeezed out of the way so that the oil and water separate.

13. Distinguish between facts, theories, and hypotheses.

A fact is a scientific statement supported by a lot of evidence that is correct so far as we know, but which can change if new data comes to light. A theory is a model of how the facts relate together. A theory explains the facts and allows new hypotheses to be formed which can be tested to produce new facts. Neither facts nor theories are truth claims; both are provisional and subject to change. A hypothesis is an informed prediction about what will happen in certain circumstances. All hypotheses are based on specific theories. It is hypotheses that are tested in experiments, and the results of the experiment either strengthen the theory (is the hypothesis is confirmed) or weaken the theory (if the hypothesis is not confirmed).

14. Explain why it is scientifically inappropriate to say, "no theory is true until it is proven."

Theories are not truth claims and are never regarded as "true," and we don't speak of them as true or false. They are models, and all models fall short of being complete descriptions of the thing they are modeling. Also, no theory is ever proven. Instead, theories get stronger if they produce successful hypotheses, i.e., those that are confirmed by experiment. So we speak of strong or weak theories. A strong theory is one that has produced many successful hypotheses and explains most or all of the known related facts.

Chapter 1 Exercises

SECTION 1.2

5. What are two of the limitations of the Bohr model of the atom?

First, it does not account for the specific electron energies possessed by atoms other than hydrogen. Second, it does not account for why electrons don't radiate energy as they orbit the nucleus and gradually spiral in to the nucleus.

6. In the Bohr model, how many electrons would you expect the 5th energy level to be able to hold? Explain your response.

Since there are 18 elements in the 5th period, the Bohr model has 18 electrons in the 5th energy level.

SECTION 1.3

7. A certain atom is in the ground state. The 3*p* subshell of this atom is 2/3 full.

a. Identify the element this atom represents.

The 3*p* subshell can hold 6 electrons. If 2/3 full, it contains 4. This means the atom is sulfur, element 16, the fourth element from the left in the *p*-block region in Period 3.

b. How many unpaired electrons are there in the atom? (A paired electron is one in an orbital with another one possessing opposite spin.)

The first 3 electrons in the 3*p* subshell go singly into the three 3*p* orbitals, $3p_x$, $3p_y$, and $3p_z$. The fourth electron goes into $3p_x$. This leaves the electrons in $3p_y$ and $3p_z$ as the only two unpaired electrons.

8. In a certain ground-state atom, the 4*d* subshell has two electrons in it.

a. Identify the element this atom represents.

This is zirconium, element 40, the second element from the left in the *d*-block region in Period 5.

b. How many unpaired electrons are there in the atom?

The two electrons in the 4*d* subshell are both unpaired.

9. How do the values of the azimuthal quantum number and magnetic quantum number relate to the principle quantum number?

Azimuthal quantum numbers (l) range from 0 to $n-1$. The range of values of the magnetic quantum number, m_l, is the integers going from $-l$ to l, which is the integers from $1-n$ to $n-1$.

10. Demonstrate mathematically that the 4*f* subshell can accommodate 14 electrons.

The number of orbitals in a subshell is determined by the range of values of the magnetic quantum number, m_l, which are the integers going from $-l$ to l. An *f* subshell has $l = 3$, so the possible values of m_l are −3, −2, −1, 0, 1, 2, 3. There are seven values, and thus seven orbitals in this subshell. Each orbital can hold 2 electrons, for a total of 14.

11. Generally speaking, what is the explanation for an atom's electron configuration not following the sequence described by the Madelung rule?

The anomalous configurations are discussed in Section 1.4 and a fuller answer to this question is given there. At this point in Section 1.3, students might anticipate that the explanation must still be due to minimizing energy and that this somehow results in a different configuration than given by the Madelung rule.

After studying Section 1.4, students can add that the *d*-block elements have many orbitals with energies spaced closely together. Violations of the Madelung rule mean that with the particular electron configuration an element has, the minimum energy state is actually slightly different from what the rule predicts.

SECTION 1.4

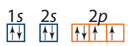
12. For each of the following elements, draw the orbital diagram and write the full-length electron configuration.

a. chlorine

Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$

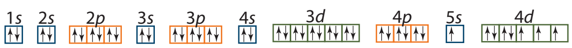
b. oxygen

O: $1s^2 2s^2 2p^4$



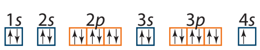
c. rubidium

Ru: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 4d^7$



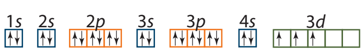
d. potassium

K: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$



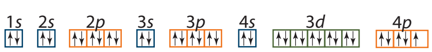
e. vanadium

V: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$



f. bromine

Br: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$



13. For each of the following elements, write the condensed electron configuration.

a. chlorine

Cl: $[Ne]3s^2 3p^5$

b. nitrogen

N: $[He]2s^2 2p^3$

c. aluminum

Al: $[Ne]3s^2 3p^1$

d. yttrium

Y: $[Kr]5s^2 4d^1$

e. strontium

Sr: $[Kr]5s^2$

f. tungsten

W: $[Xe]6s^2 4f^{14} 5d^4$

g. cesium

Cs: $[Xe]6s^1$

h. iodine

I: $[Kr]5s^2 4d^{10} 5p^5$

i: neodymium

Nd: $[Xe]6s^2 4f^4$

14. Compare the electron configurations for beryllium, magnesium, and calcium. Formulate a general rule for the condensed electron configuration of a Group 2 element.

The condensed electron configuration always begins with the noble gas from the previous period, followed by ns^2, where n is the principle quantum number.

15. For which group of elements does the electron configuration always end with np^2? Explain how you know.

Group 14, since the *p*-block elements start with Group 13, so the next group is the one in which atoms contain two 2-*p* electrons.

16. Write the condensed electron configurations for ytterbium, einsteinium, and nobelium.

Yb: $[Xe]6s^2 4f^{14}$

Es: $[Rn]7s^2 5f^{11}$

No: $[Rn]7s^2 5f^{14}$

SECTION 1.5

17. Which two nuclides in Table 1.5 have 20 neutrons?

calcium-40 (40 – 20 = 20) and chlorine-37 (37 – 17 = 20)

18. In Table 1.5, how many neutrons are there in the heaviest nuclide listed? How many neutrons are there in the lightest nuclide listed?

The heaviest is U-238, which has 238 – 92 = 146 neutrons. The lightest is H-1, which has 0.

Chapter 2 Exercises

SECTION 2.1

1. What entire group of elements did not appear in Mendeleev's original periodic table? Why were they left out and how were they put in?

The noble gases, Group 18 were left out because they don't react with other elements, so no one knew they existed. When they are finally discovered, they were added to the table simply by adding a new group to the right end of the table.

2. Write a paragraph explaining the general structure and arrangement of the periodic table.

The metals are on the left, nonmetals on the right, with a few elements called metalloids positioned in between the metals and nonmetals. Elements in the same group (column) have similar chemical properties. In elements in the table are ordered by atomic number, which represents the number of protons atoms of the element have in their nuclei.

SECTION 2.2

3. State the chief chemical property that distinguishes the metals from the nonmetals.

Metals ionize by losing electrons to become cations. Nonmetals ionize by gaining electrons to become anions.

4. Distinguish between cations and anions.

Cations are positively charged ions. Anions are negatively charged ions.

5. What is the "long form" of the periodic table, and why are there two forms (see Section 1.4.1).

The long form of the table has the inner transition elements shown in place rather than separated out at the bottom the way they are typically shown. The long form of the table is inconveniently wide, so the inner transition metals are often separated out and shown under the table to make the table easier to fit on pages in books.

SECTION 2.3

7. Describe the trend in atomic radius going down a group and across a period.

Radii get larger going down the group, and smaller going from left to right across the period.

8. Using the concept of the shield effect, write a description accounting for the trends in atomic size in the periodic table.

Going across the table from left to right, atoms have more electrons. But they also have more protons, and the increasing attraction between the nucleus and the electron cloud pulls all the electron orbitals in tighter, reducing the size of the atom going left to right across the table. Going down a group, each lower element has more electrons and they are in higher energy levels, which makes the atoms larger. Also, valence electrons are not attracted to the nucleus as strongly as we would think because of the shield effect in which the core electrons screen off the attraction of the positive nucleus from the valence electrons. The core electrons are tightly held, but the valence electrons are further out because they are not attracted as strongly because of the shield effect. This helps explain the large jump in size from an element in a given period to the one just below it.

9. From your knowledge of the periodic table, put the elements rubidium (Rb), silver (Ag), xenon (Xe), and yttrium (Y) in order of increasing atomic radius. Explain your order by referring to trends in the periodic table.

Xe < Ag < Y < Rb

10. From your knowledge of the periodic table, put the elements sodium (Na), barium (Ba), cesium (Cs), and magnesium (Mg) in order of increasing atomic radius. Explain your order by referring to trends in the periodic table.

Mg < Na < Ba < Cs

11. Based on your knowledge of trends in the periodic table, place the following atoms and ions in order of decreasing size: Be^{2+}, Mg, Ca, and Mg^{2+}.

Ca > Mg > Mg^{2+} > Be^{2+}

12. Based on your knowledge of trends in the periodic table, arrange the following atoms and ions in order of size from largest to smallest: S^{2-}, Ar, K^+, Cl^-, and Ca^{2+}.

$S^{2-} > Cl^- > Ar > K^+ > Ca^{2+}$

SECTION 2.4

13. Aluminum and scandium both ionize to +3, even though scandium is in Group 3 and aluminum is in Group 13. Explain why this is.

Each has 3 valence electrons. Aluminum is third from the left in Period 3, scandium is third from the left in Period 4.

14. Estimate the effective nuclear charge, Z_{eff}, for vanadium (V), magnesium (Mg), chlorine (Cl), and arsenic (As).

5, 2, 7, 15

15. Define ionization energy and describe the trends for ionization energy in the periodic table across periods and down groups.

Ionization energy is the energy required to remove an electron from an atom. This energy increases from left to right across the table, and decreases down the groups.

16. Referring again to Table 2.1, explain the large increase in ionization energy that occurs in the yellow shaded region of that table.

The high ionization energies are for core electrons, which are bound much more tightly to the nucleus than the valence electrons are.

17. Write a description accounting for the trends in ionization energy in terms of the shield effect and other factors.

The strong upward trend moving across each period is accounted for by the decreasing efficacy of the shield effect that occurs moving from left to right in the period. Moving from left to right, the strength of the shielding provided by the atom's core electrons remains the same, but the positive charge of the nucleus increases with each additional proton. This means the shield effect becomes less effective and the attraction between the nucleus and the valence electrons increases from left to right. The higher the nuclear attraction is, the stronger the nucleus attracts the outermost electrons and the greater the energy required to remove one of them.

The trend down a group is accounted for by the fact that the elements in each new period have their valence electrons in a shell with a higher principle quantum number. These electrons are thus farther from the nucleus, have a less negative energy, and are easier to remove.

18. Write the condensed electron configurations for Cu^{2+}, As^{5+}, Ag^+, and Au^{3+}.

Cu^{2+}: [Ar]$3d^9$

As^{5+}: [Ar]$3d^{10}$

Ag^+: [Kr]$4d^{10}$

Au^{3+}: [Xe]$4f^{14}5d^8$

19. Why are ionization energies so much higher when core electrons are involved than they are when only valence electrons are involved?

Core electrons feel the full attraction of the nucleus. Valence electrons are held much less tightly because of the shield effect of the core electrons screening off the attraction of the positive nucleus from the valence electrons.

20. Based on your knowledge of trends in the periodic table, place the following atoms in order of increasing ionization energy: Ar, Sr, P, Mg, and Ba.

Ba < Sr < Mg < P < Ar

21. Distinguish between ionization energy and electron affinity.

Ionization energy is the energy required to remove an electron from an atom. Electron affinity is the energy released when an electron is added to an atom.

22. Based on your knowledge of trends in the periodic table, place the following atoms in order of increasing electron affinity: Br, Rb, and S.

Rb < S < Br

23. Explain what the electronegativity scale is used for and how it arose.

The scale was invented by Linus Pauling in 1932 while he was researching chemical bonding. The scale describes the relative strength of the attraction of atoms of different elements for shared electrons in molecular bonds. Differences in electronegativity account for the polarization that occurs in molecules due to the different strength of attraction in the molecule of different atoms for the shared electrons in the chemical bonds.

24. Of the following cations, which is least likely to form: Ca^{3+}, Mg^{2+}, K^{+}? Explain your response.

The Ca^{3+} is least likely. Ca normally ionizes to Ca^{3+}. To ionize to Ca^{3+} a core electron would have to be removed from the atom, which does not normally occur.

25. Why do the chalcogens form ions with a charge of –2?

Main group elements ionize so as to end up with either an empty or full valence shell. Chalcogens need two more electrons to have a full valence shell, so they ionize by gaining two electrons, acquiring a change of –2 in the process.

26. Distinguish between electron affinity and electronegativity.

Electron affinity is the energy released when an electron is added to an atom. Electronegativity is the relative strength of attraction of an atom for shared electron pairs in molecular bonds.

27. Referring to the periodic table, describe the chemical properties of potassium (K), sulfur (S), xenon (Xe), iodine, (I), and manganese (Mn).

potassium: has 1 valence electron, ionizes to +1, has a very low electronegativity

sulfur: has 6 valence electrons, ionizes to –2, has a very high electronegativity

xenon: is a noble gas, so has a full valence shell and is unreactive

iodine: has 7 valence electrons, ionizes to –1, has a very high electronegativity

manganese: has 2 valence electrons, ionizes to +2, has average electronegativity

28. Which of the following is likely to have the greatest difference between the third and fourth ionization energies: Cl, Sc, Na, C?

Scandium (Sc), due to the fact that it has 3 valence electrons, so loses these easily. A fourth ionization would require losing a core electron, which requires a much higher energy to accomplish.

29. Using only the periodic table as a reference (without electronegativity data), predict the relative electronegativities of these elements and put them in order from least to greatest: Ni, Ta, Se, F, Cs, Cl.

Cs < Ta < Ni < Se < Cl < F

30. Develop an explanation for why the electron affinity values for the chalcogens are each significantly lower than those of their halogen neighbors.

The further to the right an element is in the table, the less pronounced the shield effect is. The shield effect reduces the attraction of a nucleus for the atom's electrons, but also affects the attraction of the nucleus for other atoms' electrons. Halogens are to the right of the chalcogens, so the shield effect in the halogens is less pronounced and they thus their nuclei attract neighboring atoms' electrons more strongly. This means they have higher electronegativities.

31. Why don't the noble gases have electronegativity values listed in Figure 2.18?

These elements have full valence shells and are unreactive. They do not normally ionize.

SECTION 2.5

32. Describe the chemical properties that place hydrogen in Group 1, and the chemical properties hydrogen shares with Group 17 elements.

Hydrogen's electron is in an *s*-subshell, like the Group 1 metals and it ionizes to +1 like the Group 1 metals. In solution, hydrogen atoms from acids dissociate to form cations, just like Group 1 metals in ionic crystals. But hydrogen tends to share one pair of electrons to form molecules, just like halogens. Also, hydrogen makes H_2 molecules with itself, like halogens. Finally, hydrogen can ionize by gaining an electron as halogens do.

GENERAL REVIEW EXERCISES

35. Identify the block, period, and group for the elements represented by each of the following condensed electron configurations:

a. [Ne]$3s^23p^3$

P: *p*, 3, 15

b. [Xe]$6s^24f^{14}5d^{10}6p^1$

Tl: *p*, 6, 13

c. [Kr]$5s^14d^5$

Mo: *d*, 5, 6

d. [Ar]$4s^23d^3$

V: *d*, 4, 5

36. How many orbitals are there in the shell associated with $n = 4$? How many electrons can this shell hold?

16, 32

39. Identify some specific differences between the chemical properties of the alkali metals and those of the transition metals.

Alkali metals all ionize to +1 because they have a single electron in an *s*-subshell of the valence shell. Most of the transitions metals can take on multiple oxidation states because they have valence electrons in *s*- and *d*-subshells. Also, the alkali metals all have very low electronegativities, but the most of the transition metals have mid-level or average electronegativities.

43. What is meant by the phrase, "chemistry is all about modeling"?

Chemistry deals with quantities we cannot see, even with a microscope, so to describe nature we must propose theories or models to explain the facts we have discovered. These models are always provisional.

44. Why is it that scientists, when they are being accurate in their speech, avoid using the term truth? What are they likely to say instead?

Scientific facts are said to be correct so far as we know. Strong theories are regarded as our best model or explanation at the time. Both are provisional and subject to change as new knowledge is gained, so neither are regarded as actually the truth. A truth claim is a strong claim about the way something really is, and can be known by direct observation, valid logic, or divine revelation. We do not have this kind of certain knowledge about the way nature works, so scientific statements (facts and theories) are not regarded as truth claims.

45. Is there a difference between scientific facts and historical facts? If so, what is it?

Both are said to be "correct so far as we know." But scientific facts are based on repeatable experiments, so they can always be verified by repeating an experiment. Historical facts are based on non-repeatable history and cannot be verified in the same way.

46. What is the difference between molar mass and molecular mass?

Molar mass is the mass of one mole (6.022×10^{23} molecules or formula units) of a substance. Molecular mass is the mass of a single molecule.

47. Why must a calculation of molecular mass based on periodic table data necessarily be an average mass and not the mass of a specific molecule?

Because the atoms of which molecules are composed all come in various isotopes which have different masses due to the different numbers of neutrons they possess. The masses shown in the table are based on the weighted average of all an elements isotopes.

48. Distinguish between the two definitions for the mole.

The first is: A mole is the amount of a substance that contains the same number of particles as there are atoms in 12 grams of carbon-12. This value is equal to Avogadro's number, which leads to the second definition: A mole is the amount of a pure substance (element or compound) that contains Avogadro's number of particles of the substance.

Chapter 3 Exercises

SECTION 3.1

1. Classify each of the following as element, compound, homogeneous mixture, or heterogeneous mixture.

a. water

compound

b. cesium chloride

compound

c. pond water

heterogeneous mixture

d. methane

compound

e. a soft drink

homogeneous mixture

f. nitric acid

homogeneous mixture

g. black coffee

homogeneous mixture

h. argon

element

i. air

homogeneous mixture

j. hydrogen nitrate

compound

k. exhaust fumes

heterogeneous mixture

l. quartz

compound

m. brass

homogeneous mixture

n. hydrogen gas

element

o. hydrogen cyanide

compound

p. mouthwash

homogeneous mixture

q. platinum

element

r. dirt

heterogeneous mixture

s. radon

element

t. a smoothie

heterogeneous mixture

2. Explain why salt water and sugar water are homogeneous mixtures while automotive paint, which contains invisible particulates, is not.

Salt water and sugar water are aqueous solutions, and thus homogeneous mixture. Since paint contains particulates, it is not uniform down to the molecular level; it contains large particles, even though they are not visible to the naked eye. Thus paint is a heterogeneous mixture.

3. Explain the octet rule.

Adding the two possible electrons in an *s* subshell to the six possible electrons in a *p* subshell gives eight electrons—an *octet*. The noble gas electron configuration is the most stable of possible electron configurations, where all electrons are in a stable state that minimizes their energies. Other main group elements (and many of the transition metals) tend to ionize or share electrons with other atoms in such a way that they acquire a noble gas electron configuration, the stablest way to be. In other words, *atoms tend to gain or lose electrons to achieve an octet of electrons in the highest-energy occupied shell*. This is characteristic is the *octet rule*.

SECTION 3.2

4. Explain what lattice energy is and why the lattice energy of ionic compounds is so large.

The lattice energy is the energy that binds the ions together in the crystal lattice. Lattice energies are large because of the enormous number of ions that attract together to form the lattice. Each ion releases energy as it rushes to connect electrically to another ion in the lattice (just as a falling ball releases energy as it rushes toward the ground). Because of the large number of ions in a mole of any crystal (at least 1.2×10^{24} ions), the lattice energy is large.

5. Distinguish between ionic and covalent bonds.

In ionic bonds, electrons are transferred from metal atoms to nonmetal atoms during ionization. The resulting ions form a crystal lattice held together by electrical attraction between ions. In covalent bonds, atoms share electrons in pairs so that each achieves an octet (noble gas configuration). Each group of atoms sharing electron pairs is a molecule, and the molecule is held together by the separate attractions of atomic nuclei for the same pair of electrons.

6. Write the formula for the ionic compound formed by each of the following cation-anion pairs:

a. Sr and Br

$SrBr_2$

b. Mg and O

MgO

c. Cu(II) and Cl

$CuCl_2$

d. Li and Se

Li_2Se

e. Ni^{2+} and P

Ni_3P_2

f. K and N

K_3N

g. Be and Cl

$BeCl_2$

h. Na and I

NaI

i. Cr^{3+} and O

Cr_2O_3

j. Ag and Se

Ag_2Se

k. Mn(IV) and F

MnF_4

l. Sb^{3+} and Br

$SbBr_3$

m. Rb and S

Rb_2S

n. Sc and O

Sc_2O_3

o. Al and F

AlF_3

p. Mg and F

MgF_2

q. Zn and Cl

$ZnCl_2$

r. Li and S

Li_2S

s. Co(III) and O

Co_2O_3

t. V^{5+} and I

VI_5

7. Write the names for each compound in the previous question. For multivalent cations, write the name in both the Latin-based system and the Stock system.

a. Sr and Br

strontium bromide

b. Mg and O

magnesium oxide

c. Cu(II) and Cl

copper(II) chloride; cupric chloride

d. Li and Se

lithium selenide

e. Ni^{2+} and P

nickel phosphide

f. K and N

potassium nitride

g. Be and Cl

beryllium chloride

h. Na and I

sodium iodide

i. Cr^{3+} and O

chromium(III) oxide; chromous oxide

j. Ag and Se

silver selenide

k. Mn(IV) and F

manganese(IV) fluoride; manganic fluoride

l. Sb^{3+} and Br

antimony(III) bromide; antimonous bromide

m. Rb and S

rubidium sulfide

n. Sc and O

scandium oxide

o. Al and F

aluminum fluoride

p. Mg and F

magnesium fluoride

q. Zn and Cl

zinc chloride

r. Li and S

lithium sulfide

s. Co(III) and O

cobalt(III) oxide; cobaltic oxide

t. V^{5+} and I

vanadium(V) iodide; vanadic iodide

8. Explain what a hydrate is and the conditions under which a hydrate forms.

In some ionic compounds, it is common for water molecules to integrate into the crystal lattice if the crystals are grown in an aqueous (water) solution. These are hydrates.

9. Explain why ionic compounds are brittle and cannot be bent into shapes the way metals can.

Ionic crystals can flex to some extent. But when subjected to a high enough stress, a crystal lattice breaks apart instead of flexing. This is due to the strong electrical forces in the crystal lattice. These forces make the crystal strong, but when subjected to high enough stress, the attractions holding the crystal together are overcome by repulsions between misaligned, identically charged ions.

10. Explain why ionic compounds generally have high melting points.

This is due to the strong electrical forces in the crystal lattice.

11. Since energy is required to separate the ions in a crystal (the lattice energy), why doesn't the temperature of the water drop when salt is dissolved in water?

Because this energy is not supplied from the thermal energy in the water. It is instead supplied by the polar water molecules rushing in to surround the ions separated from the crystal lattice.

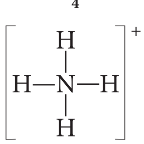
SECTION 3.3

12. Draw the Lewis structures for the following compounds and ions. Indicate which of the molecules have resonance structures and show them. (In each of the items shown, when H and O are both present, the H bonds to O.)

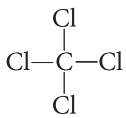
a. C_2HCl

H—C≡C—Cl

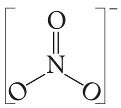
b. NH_4^+

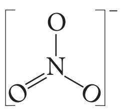


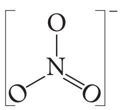
c. CCl_4



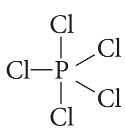
d. NO_3^-



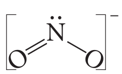


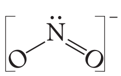


e. PCl_5



f. NO_2^-

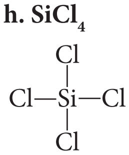




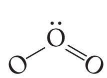
g. CO

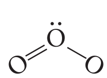
$C\equiv O$

h. $SiCl_4$

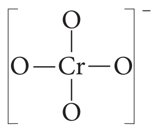


i. O_3

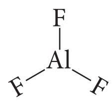




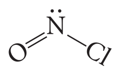
j. CrO_4^{2-}



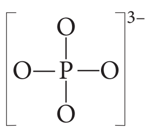
k. AlF_3



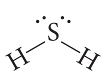
l. NOCl



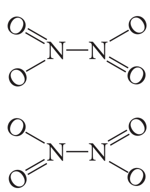
m. PO_4^{3-}



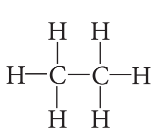
n. H_2S



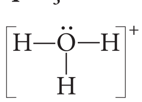
o. N_2O_4



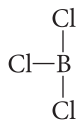
p. C_2H_6



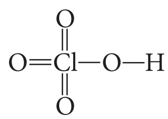
q. H_3O^+



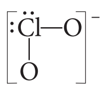
r. BCl_3



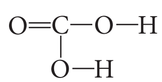
s. $HClO_4$



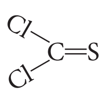
t. ClO_2^-



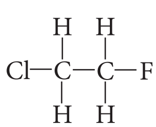
u. H_2CO_3



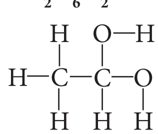
v. CCl_2S



w. C_2H_4ClF



x. $C_2H_6O_2$



13. State the names for each of the binary compounds listed in the previous exercise. (Several of the compounds are not binary compounds.)

c. CCl_4

carbon tetrachloride

e. PCl_5

phosphorus pentachloride

g. CO

carbon monoxide

h. $SiCl_4$

silicon tetrachloride

k. AlF_3

aluminum trifluoride

n. H_2S

dihydrogen sulfide (common name hydrogen sulfide)

o. N_2O_4

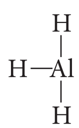
dinitrogen tetroxide

r. BCl_3

boron trichloride

14. Draw the Lewis structures for the following compounds and ions. Identify (i) those that do not obey the octet rule (explain why), (ii) resonance structures (show all), and (iii) free radicals (explain why).

a. AlH_3

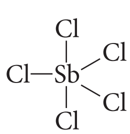


Exception to the octet rule. Aluminum and boron can bond with only three electron pairs instead of four. This is due to energy minimization in the molecule.

b. CS_2

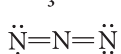
S=C=S

c. $SbCl_5$



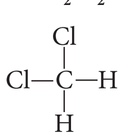
Exception to the octet rule. The *p*-block elements in periods 3 and higher can form bonds with five electron pairs instead of four. The usuual explanation is that since the central atom is a *p*-block element, it has a completely unfilled *d* subshell that can take on the extra pair.

d. N_3

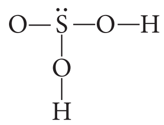


Free radical. There is an odd number of valence electrons to share, so one of the nitrogen atoms on the ends will always have an unpaired electron.

e. CH_2Cl_2



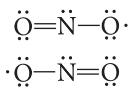
f. H_2SO_3



g. OH

Free radical. There is an odd number of valence electrons to share, so the oxygen atom has an unpaired electron.

h. NO_2



Free radical. There is an odd number of valence electrons to share, so one of the oxygen atoms on the ends will always have an unpaired electron.

15. State the names of the following acids.

a. H_2SO_4
sulfuric acid

b. HF

hydrofluoric acid

c. H_2CO_3
carbonic acid

d. $HClO_4$
perchloric acid

e. HNO_2
nitrous acid

f. $HBrO_2$
bromous acid

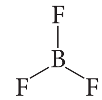
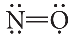
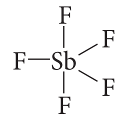
g. HBr
hydrobromic acid

h. H_3PO_4
phosphoric acid

16. Explain why covalent compounds tend to be gases or liquids at room temperature, while ionic compounds tend to be solids.

The lattice energy in ionic compounds binds the entire crystal structure together. But the bond energy in covalent compounds only holds together the atoms in the individual molecules. The forces holding one molecule to another are much weaker. One of the results of the weak forces holding molecules together is that many common covalent compounds are gases or liquids at room temperature.

17. Describe the three general cases of exceptions to the octet rule (other than hydrogen and helium), and give an example for each. Show the Lewis structures of your examples.

First, boron forms three bonds, leaving no nonbonding pair on the central atom. BF_3 is an example.  Boron does this because it is not energetically favorable to achieve a full octet. The second case is when free radicals form, as with nitric oxide, NO.  . The third case occurs with elements that form more bonds than are required to achieve an octet, such as antimony, which forms compounds such as antimony pentafluoride, SbF_5.  The elements that bond this way are *p*-block elements in periods three and higher. The explanation normally given for why the central atom in these compounds accepts more than an octet of electrons is that since the central atom is a *p*-block element, it has a completely unfilled *d* subshell that can take on the extra pair. Some chemists are now of the opinion that there are different factors behind this type of exception to the octet rule.

18. Write the formulas for the ionic compounds formed from each of the ion pairs given below, and state the name of each using the Stock system. (Of course, naming these is a snap.)

a. magnesium and carbonate

$MgCO_3$ magnesium carbonate

b. sodium and chromate

Na_2CrO_4 sodium chromate

c. calcium and phosphate

$Ca_3(PO_4)_2$ calcium phosphate

d. ammonium and phosphate

$(NH_4)_3PO_4$ ammonium phosphate

e. copper(II) and sulfate

$CuSO_4$ copper(II) sulfate

f. sodium and bicarbonate

$NaHCO_3$ sodium bicarbonate

g. iron(III) and chlorate

$Fe(ClO_3)_3$ iron(III) chlorate

h. aluminum and hydroxide

$Al(OH)_3$ aluminum hydroxide

i. chromium(VI) and sulfite

$Cr(SO_3)_3$ chromium(VI) sulfite

j. calcium and acetate

$Ca(CH_3COO)_2$ calcium acetate

19. Explain what resonance structures are.

Resonance structures are molecules that can form in two or more different ways that are identical except for the position of one or more double bonds. The molecular structure is understood to be a blend of the separate resonance structures. The electrons in such structures are said to be delocalized, being shared in orbitals that extend across several atoms.

SECTION 3.4

21. Explain why water molecules are polar and chlorine molecules are not.

The electronegativity of oxygen is much higher than that of hydrogen, so in the water molecule the electrons are pulled toward the oxygen atom and away from the hydrogen atoms. This makes the region around the oxygen atom more negative and the region around the hydrogen atoms more positive. In a chlorine molecule, the electronegativities of the two atoms are the same, so the electrons are balanced evenly in the molecule and there is no polarity.

23. When chlorine bonds to the three ions below, it forms compounds that are yellow, orange, or red in aqueous solution. Predict the color each compound displays.

a. As(III)

red

b. Cd(II)

yellow

c. Sb(III)

orange

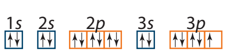
GENERAL REVIEW EXERCISES

26. In Bohr's model of the atom, the first four energy levels could hold 2, 8, 8, and 18 electrons, respectively. The values 2, 8, 18, and 32 were later found to be the correct values for *n* = 1, 2, 3, and 4. Explain where Bohr's numbers came from, and what determines the numbers of electrons in the shells. Hint: compute the difference between each of the electron quantities and relate that difference to the quantum model of the atom.

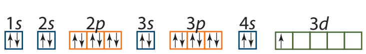
Bohr's numbers came from the numbers of elements in the periods of the periodic table. The quantum model has a single *s* orbital in $n = 1$, which holds two electrons. In $n = 2$ there are *s* and *p* subshells, which hold 2 and 6, respectively, for 8 total. In $n = 3$, there are *s*, *p*, and *d* subshells, which hold 2, 6, and 6, respectively, for 18 total. In $n = 4$, there are *s*, *p*, *d*, and *f* subshells, which hold 2, 6, 10, and 14, respectively, for 32 total.

27. Write the electron configurations and draw the orbital diagrams for Cl, Sc, Ag, and Gd.

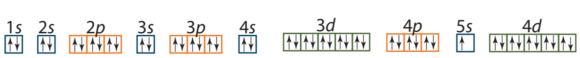
Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$



Sc: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$



Ag: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 4d^{10}$



Gd: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^7 5d^1$

30. Why are the noble gases unreactive?

They all have full octets and exhibit no tendency to share or exchange electrons with other atoms. The full octet is the most stable, minimum energy electron configuration.

Sample Course Lesson Schedule **Chemistry for Accelerated Students** **Fall Semester**

Lesson Number	Activity or Topic	Text Key	Assignment	Notes
1	Welcome/What is Chemistry All About	I.1.1-I.1.2		
2	What is Chemistry All About	I.1.3-I.1.4		
3	What is Chemistry All About	I.1.5	Intro study questions	
4	Quiz 1 (Intro) Electromagnetic Spectra, Planck relation	1.1.1		Quiz 1
5	Lab Day: Identification of Substances by Physical Properties		Lab Report	This is Experiment 1 in *Chemistry Experiments for High School (CEHS) and Chemistry Experiments for High School at Home (CEHSH)*
6	Hydrogen atom, Bohr model	1.1.2-1.2	1.1-1.2 exercises	
7	Quantum atomic model	1.3.1-1.3.2		
8	Quiz 2 (1.1-1.2) Quantum atomic model	1.3.2		Quiz 2
9	Aufbau principle, Madelung rule, Hund's rule	1.3.3	1.3 exercises	
10	Discussion			
11	Lab Day: Flame tests and metal cation identification		Lab Report	This is Experiment 3 in *CEHS* and *CEHSH*
12	Electron configurations, empirical formulas	1.4-1.5	1.4, 1.5 exercises	
13	Quiz 3 (1.3-1.4) Periodic Table	2.1-2.2	2.1-2.2 exercises begin ch 2 general review exercises	Quiz 3
14	Lab Day: Determining an empirical formula		Lab Report	This is Experiment 4 in *CEHS* and *CEHSH*
15	Lab Day: Determining an empirical formula (continued)			
16	Test Intro & Chapter 1			Test Intro & Chapter 1
17	Physical periodic properties	2.3.1-2.3.2	2.3 exercises	
18	core/valence electrons, ionization energy	2.4.1-2.4.3		
19	Quiz 4 (2.1-2.3) electron affinity	2.4.4		Quiz 4
20	electronegativity, hydrogen	2.4.4-2.5	2.4-2.5 exercises	
21	Discussion			
22	Quiz 5 (2.4) Substances	3.1.1-3.1.2		Quiz 5
23	Test Chapter 2			Test Chapter 2
24	Octet rule, ionic bonds, naming ionic compounds, energy, hydrates, intensive & extensive properties	3.1.3-3.2.5	3.1 exercises begin ch 3 general review exercises	
25	Lab Day: Separation of Components in a Mixture		Lab Report	This is Experiment 2 in *CEHS* and *CEHSH*
26	Discussion			
27	Quiz 6 (3.1-3.2) Ionic physical properties	3.2.6	3.2 exercises	Quiz 6
28	Covalent bonds, molecules, polyatomic ions, polyatomic ion names, acid names	3.3.1-3.3.5	begin 3.3 exercises	
29	Lewis structures	3.3.6		
30	Discussion			
31	Octet Exceptions, resonance structures, naming binary compounds, energy in covalent bonds, bond number, physical properties	3.3.7-3.3.12	3.3 exercise	
32	Polarity and dipoles; nature of the bond	3.4	3.4 exercises	
33	Quiz 7 (3.3)			Quiz 7
34	Discussion			
35	Test Chapter 3			Test Chapter 3
36	Covalent Bond theory and VSEPR theory	4.1.1-4.1.2	begin ch 4 general review exercises	
37	Nonbonding domains and bond angle	4.1.3		
38	Orbital hybridization theory, valence bond theory	4.1.4-4.1.5	4.1 exercises	
39	Discussion			
40	Quiz 8 (4.1) Metallic bonding, Physical properties of metals	4.2.1-4.2.2	4.2 exercises	
41	Intermolecular forces, hydrogen bonding, Van der Waals forces	4.3.1-4.3.4	4.3 exercises	
42	Quiz 9 (4.2)			Quiz 9
43	Lab Day: Intermolecular Forces		Lab Report	This is Experiment 7 in *CEHS* and *CEHSH*
44	Discussion			
45	Test Chapter 4			Test Chapter 4
46	Law of conservation of mass and balancing equations	5.1.1-5.1.4	begin ch 5 general review exercises	
47	Oxidation states	5.1.5	5.1 exercises	
48	Reaction types, activity series of metals	5.2.1-5.2.6		
49	Quiz 10 (5.1)	5.2.7-5.2.8	5.2 exercises	Quiz 10

	Reaction types			
50	Discussion			
51	Lab Day: Activity Series of Metals		Lab Report	This is Experiment 5 in *CEHS* and *CEHSH*
52	Stoichiometry	5.3.1		
53	Limiting reagent, theoretical yield, percent yield	5.3.2-5.3.3	5.3 exercises	
54	Quiz 11 (5.2)			Quiz 11
55	Lab Day: Limiting Reactant and Percent Yield		Lab Report	This is Experiment 6 in *CEHS* and *CEHSH*
56	Discussion			
57	Test Chapter 5			Test Chapter 5
58	Kinetic molecular theory of gases	6.1.1-6.1.3	begin ch 6 general review exercises	
59	Gas pressure, pressure units	6.1.4	6.1 exercises	
60	States of matter	6.2.1-6.2.5		
61	Quiz 12 (6.1) Phase transitions	6.2.6	begin 8.2 exercises	Quiz 12
62	Heat capacity, heat of fusion/vaporization	6.2.7		
63	Discussion			
64	Evaporation, vapor pressure	6.2.8-6.2.9	6.2 exercises	
65	Quiz 13 (6.2)			Quiz 13
66	Discussion			
67	Test Chapter 6			Test Chapter 6
68	Lab Day: Metathesis Reactions		Lab Report	This is Experiment 10 in *CEHS* and *CEHSH*
69	Review			

Course Lesson Schedule **Chemistry for Accelerated Students** **Spring Semester**

Lesson Number	Topic	Text Key	Assignment	Notes
1	Boyle's law, Charles' law, Avogadro's law	7.1.1-7.1.3	7.1 exercises	
2	STP, ideal gas law	7.2.1-7.2.2	begin ch 7 general review exercises	
3	Quiz 14 (7.1) ideal gas law	7.2.2	begin 7.2 exercises	Quiz 14
4	Discussion			
5	Using ideal gas law to find molar mass and density	7.2.3	7.2 exercises	
6	Quiz 15 (7.2) Law of partial pressure	7.3.1		Quiz 15
7	Collecting gas over water, stoichiometry of gases	7.3.2-7.4.1	7.3 exercises, begin 7.4 exercises	
8	Lab Day: Mole Amount of a Gas		Lab Report	This is Experiment 9 in *CEHS* and *CEHSH*
9	Diffusion and effusion	7.4.2	7.4 exercises	
10	Discussion			
11	Test Chapter 7			Test Chapter 7
12	Suspensions, colloids, dissolution	8.1-8.2.1	8.1 exercises begin ch 8 general review exercises	
13	Entropy of solution, entropy, free energy	8.2.2-8.2.3		
14	Electrolytes, Solubility of ionic solids	8.2.4-8.3.2	8.2 exercises	
15	Solutions of polar and nonpolar liquids	8.3.3-8.3.4		
16	Quiz 16 (8.1-8.2)			Quiz 16
17	Solid solutions, gases in solution, Le Chatelier's principle, Henry's law, temperature effect on solubility	8.3.5-8.3.7	8.3 exercises	
18	Discussion			
19	Molarity, molality	8.4.1-8.4.2	8.4 exercises	
20	Ionic equations, precipitates, net ionic equations, spectator ions	8.5.1-8.5.2		
21	Quiz 17 (8.3-8.4) Colligative properties, vapor pressure lowering	8.6.1		Quiz 17
22	Lab Day: Molarity		Lab Report	This is Experiment 8 in *CEHS* and *CEHSH*
23	Boiling point elevation, freezing point depression, osmotic pressure	8.6.2-8.6.3	8.6 exercises	
24	Discussion			
25	Test Chapter 8			Test Chapter 8
26	Acid base intro, names, formulas	9.1.1-9.1.3	9.1 exercises begin ch 9 general review exercises	
27	Acid-base theories: Arrhenius	9.2.1		
28	Acid-base theories: Bronsted-Lowry	9.2.2		
29	Bronsted-Lowry, Lewis, acid-base strength	9.2.3-9.2.4	9.2 exercises	
30	Discussion			
31	Self-ionization of water, $[H_3O^+]$, $[OH^-]$	9.3.1-9.3.2	begin 9.3 exercises	
32	Quiz 18 (9.1-9.2.3) pH	9.3.3		Quiz 18
33	pH	9.3.3		
34	pH indicators, titration	9.3.4		
35	Quiz 23 (9.2-9.3.2) Titration	9.3.5		Quiz 19
36	Titration	9.3.5		
37	Determining $[H_3O^+]$ and $[OH^-]$ from titration data	9.3.6	9.3 exercises	
38	Lab Day: Acid-base titration		Lab Report	This is Experiment 11 in *CEHS* and *CEHSH*
39	Discussion			
40	Lab Day: Effectiveness of antacids		Lab Report	This is Experiment 12 in *CEHS* and *CEHSH*
41	Test Chapter 9			Test Chapter 9
42	Lab Day: Titration curves and K_a			This is Experiment 16 in *CEHS* and 15 *CEHSH*
43	Enthalpy	10.1.1-10.1.3	begin ch 10 general review exercises	
44	Enthalpy of combustion, and formation	10.1.4-10.1.5	10.1 exercises	
45	Hess' law	10.2.1		
46	Quiz 20 (10.1) General enthalpy change equation	10.2.2	10.2 exercises	Quiz 20
47	Lab Day: Calorimetry and Hess' Law			This is Experiment 13 in *CEHS* and *CEHSH*
48	Gibb's free energy	10.3.1-10.3.3	10.3 exercises	
49	Discussion			
50	Quiz 21(10.2-10.3) Reaction kinetics, collision theory	10.4.1	begin 10.4 exercises	Quiz 21
51	Reaction rate factors, reaction mechanisms	10.4.2-10.4.3		
52	Activation energy, activation complex, reaction rate laws	10.4.4-10.4.5		

53	Rate laws and reaction mechanisms	10.4.6	10.4 exercises	
54	Discussion			
55	Lab Day: Rate Law Determination			This is Experiment 14 in *CEHS* and *CEHSH*
56	Test Chapter 10			Test Chapter 10
57	Physical and chemical equilibria	11.1.1-11.1.2	begin 11.1 exercises begin ch 11 general review exercises	
58	Law of chemical equilibrium, equilibrium constant	11.1.3		
59	Le Chatelier's principle and equilibrium shifts	11.1.4		
60	Reactions that go to completion, common-ion effect	11.1.5-11.1.6	11.1 exercises	
61	Discussion			
62	Lab Day: Le Chatelier's principle		Lab Report	This is Experiment 15 in *CEHS* (not included in *CEHSH*)
63	Quiz 22 (11.1) Acid-dissociation constant	11.2.1		Quiz 22
64	Base dissociation constant, buffering	11.2.2-11.2.3		
65	Hydrolysis of salts	11.2.4	11.2 exercises	
66	Solubility product, calculating K_{sp}	11.3.1-11.3.2	begin 11.3 exercises	
67	Quiz 23 (11.2) Calculating solubility from K_{sp}	11.3.3		Quiz 23
68	Using K_{sp} to predict precipitation	11.3.4	11.3 exercises	
69	Discussion			
70	Test Chapter 11			Test Chapter 11
71	Redox reactions, oxidation states	12.1.1-12.1.2	begin ch 12 general review exercises	
72	Strength of oxidizing and reducing agents	12.1.3	12.1 exercises	
73	Quiz 24 (12.1) Redox half-reactions	12.2.1	begin 12.2 exercises	Quiz 24
74	Balancing redox equations	12.2.2		
75	Balancing redox equations	12.2.2	12.2 exercises	
76	Discussion			
77	Lab Day: Redox titration		Lab Report	This is Experiment 17 in *CEHS* and 16 *CEHSH*
78	Quiz 25 (12.2)			Quiz 25
79	Electrochemistry	12.3.1-12.3.2	begin 12.3 exercises	
80	Electrochemical cells	12.3.3		
81	Electrode potentials	12.3.4		
82	Lab Day: Electrochemical series		Lab Report	This is Experiment 18 in *CEHS* and 17 *CEHSH*
83	Electrochemical applications	12.3.5	12.3 exercises	
84	Discussion			
85	Test Chapter 12			Test Chapter 12
86				