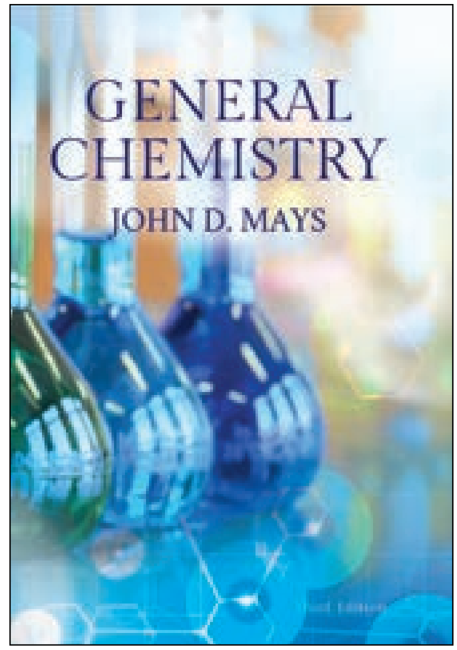
General Chemistry

All Keys and Sample Answers



Thank you for using *General Chemistry*. This document contains sample answers to all verbal questions in the text, on quizzes, and on exams. Also included are the solutions to the computations on the quizzes and tests. Solutions to computational exercises in the text are available separately in the solutions manual, available on our website. (Answers to all exercises are found in the book.)

The sample answers in this document are 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 written answers provided here are only samples and should not be considered the only correct responses to the questions.

In environments where there are multiple students in a class or group, it is recommended that students form their own answers to chapter exercises 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 to improve their answers. The final product is a useful study tool developed by the group. In such a setting, there is little need for the written answers in the present 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 on the Resource CD. A full presentation of strategies and techniques for mastery-learning can be found *From Wonder to Mastery*, available from our website.

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NOTE: Full solutions to the computations in the chapter exercises are available in the *Solutions Manual to Accompany General Chemistry*. To procure a copy, visit novarescienceandmath.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 1.1

1. Write a paragraph distinguishing between matter and mass.

Matter consists of anything composed of atoms or parts of atoms. Mass is a variable used to quantify the amount of inertia (a property of matter) an object has.

2. Distinguish between base units and derived units in the SI system of units and give three examples of each.

The entire SI system of units is based on seven base units of measure—meter, kilogram, second, ampere, kelvin, candela, and mole. All other units of measure are derived from one or more of the base units. Examples of derived units are joule, newton, cubic meter, watt, and pascal.

3. Describe the advantages the SI system has over the USCS system for scientific work.

The SI system does not use random numbers such as 3, 12, and 5280. Also, there is only one major unit of measure for each kind of quantity. Instead of using several different units for measurements of different sizes, the SI system uses prefixes to scale up or down the unit used for the quantity in question.

4. Why does the SI system use prefixes on the units of measure?

Instead of using several different units for measurements of different sizes, the SI system uses prefixes to scale up or down the unit used for the quantity in question.

5. Re-express the quantities in the following table using only a single numerical digit followed by an SI unit symbol, with a metric prefix if necessary. Example: 5 thousand liters = 5 kL

a. 8 Pa

b. 5 cm

c. 3 MA

d. 2 km

e. 4 ms

f. 6 kN

g. 8 kg

h. 7 μL

i. 1 mJ

6. Re-write the quantities in the following table by writing out the unit names without symbols. Example: 5 km = 5 kilometers

a. 14 cubic meters

b. 164.1 kilograms

c. 250 megapascals

d. 16.533 milliseconds

e. 160 kiloamps

f. 19.55 centiliters

g. 31.11 microjoules

h. 2300 kelvins

i. 13.0 millimoles

SECTION 1.2

7. Why must equations be used instead of conversion factors for most temperature unit conversions?

Most temperature scales are not absolute scales. If they were, conversion factors could be used.

SECTION 1.3

11. Distinguish between accuracy and precision.

Accuracy relates to error—the greater the accuracy, the lower the error. Error is the difference between a measurement of a quantity and the quantity's true value. Accuracy can be increased by eliminating the sources of error. Precision is the amount of fine-ness or resolution in a measurement, and is determined by the instrument used to take the measurement. To increase precision, a more precise measurement instrument must be used. The precision in a measurement is indicated by the number of significant digits in the measurement.

12. Describe the measurements you would obtain from an instrument that was very precise but not very accurate.

You would have lots of digits and/or decimal places in the order, but the more precise digits would be meaningless. For example, in a measurement such as 44.1234, if the accuracy was off by 5% there would be an error of around 2.2. Thus, the actual value could be anywhere from about 41.9 to 46.3 and all the decimal places in the measurement have no value.

13. Which is more important on the speedometer of a car—accuracy or precision?

No one needs to know their speed to a precision of greater than the nearest 1 mile per hour, so precision isn't really that important. However, accuracy is important because if your speedometer was off by a few percent you could be driving faster than you knew. This could lead to unsafe driving or speeding violations.

14. Explain why accuracy is important on a heart rate monitor but precision is not.

This is similar to the previous question. The precision of a heart rate does not need to be greater than about 1 beat per minute, but accurate readings are important for medical personnel to monitor patients' condition.

15. Sometimes we want high accuracy in a measurement, but are not too concerned about high precision. Sometimes we want both high accuracy and high precision. Explain why no one wants low accuracy and high precision.

This is discussed in question 12. Having lots of digits and/or decimal places is of no value if the measurement is not accurate to begin with.

16. On the package of a digital stopwatch I once purchased was the phrase: "1/100th second accuracy." The stopwatch readings in seconds contained two decimal places, but the values the stopwatch actually displayed were spaced 0.03 seconds apart. Thus, it could read 12.31 s, 12.34 s, 12.37 s, etc. Comment on the accuracy and precision of this stopwatch with respect to the claim on the package.

What is probably meant is that the watch is precise to 1/100th of a second, which means the stopwatch shows seconds, tenths of seconds, and hundredths of seconds. In this case, the term the package uses is incorrect. It could actually be a statement about the accuracy, in which case decimals more precise than a hundredth of a second would be unreliable. But since possible readings are spaced 0.03 s apart, the accuracy is not to the nearest 1/00 of a second. If the time was 12.325 would be displayed as 12.34 and would have an error of 0.015 s, which is more than 0.01 s.

17. Using the correct number of significant digits and the correct units of measure, record the measurements represented by the following instruments.

Note: The last digit in measurements (a) and (b) must be estimated, and thus could vary by 1 or 2 from the answers given here. However, measurements should have the same number of digits as the answers shown.

a) 17.9 psi

b) 0.28 mL

c) 1.320 × 103 gal

d) 1.0 × 101 mph

Chapter 2 Exercises

SECTION 2.1

1. Write paragraphs describing the experiments performed by J.J. Thomson, Robert Millikan, and Ernest Rutherford.

Thomson placed electrodes from a high-voltage electrical source inside a sealed-glass vacuum tube. This apparatus can generate a cathode ray from the negative electrode, called the cathode, to the positive one, called the anode. The anode inside Thomson's vacuum tube had a hole in it for some of the electrons to escape through, which created a beam of cathode rays heading toward the other end of the tube. Thomson placed the electrodes of another voltage source inside the tube, above and below the cathode ray, and discovered that the beam of electrons deflected when this voltage was turned on. He also placed magnetic coils on the sides of the tube and discovered that the electrons also deflected as they passed through the magnetic field produced by the coils. The deflection of the beam toward the positive electrode led Thomson to theorize that the beam was composed of negatively charged particles. By trying out many different arrangements of cathode ray tubes, Thomson confirmed that the ray was negatively charged. Then using the scale on the end of the tube to measure the deflection angle, he was able to determine the charge-to-mass ratio of the individual electrons he had discovered.

Inside a heavy metal drum about the size of a 5-gallon bucket, Millikan placed a pair of horizontal metal plates connected to an adjustable high-voltage source. The upper plate had a hole in the center and was connected to the positive voltage, the lower plate to the negative. He used an atomizer spray pump to spray in a fine mist of watchmaker's oil above the positive plate. Some of the oil droplets would fall through the hole in the upper plate and move into the region between the plates. Connected through the side of the drum between the two plates was a telescope eyepiece and lamp so that Millikan could see the oil droplets between the plates. The process of squirting in the oil droplets with the atomizer sprayer caused some of the droplets to acquire a charge of static electricity. This means the droplets had excess electrons on them and carried a net negative charge. They picked up these extra electrons by friction as the droplets squirted through the rubber sprayer tube. As Millikan looked at an oil droplet through the eyepiece and adjusted the voltage between the plates, he could make the charged oil droplet hover when the voltage was just right. Millikan took into account the weight of the droplets and the viscosity of the air as the droplets fell and was able to determine that every droplet had a charge on it that was a multiple of 1.6 × 10–19 C. From this he deduced that this must be the charge on a single electron.

Rutherford created a beam of α-particles by placing some radioactive material (radium bromide) inside a lead box with a hole in one end. The α-particles from the decaying radium atoms streamed out the hole at very high speed. Rutherford aimed the α-particles at an extremely thin sheet of gold foil only a few hundred atoms thick. Surrounding the gold foil was a ring-shaped screen coated with a material that glows when hit by α-particles. Rutherford could then determine where the α-particles went after encountering the gold foil. When Rutherford began taking data, he had Thomson's plum pudding model in mind and was expecting results consistent with that atomic model. In the plum pudding model, the atom's positive charge is spread throughout the atom and the negative electrons are embedded in the positive material. Rutherford expected the massive and positively charged α-particles to blow right through the gold foil. What Rutherford found was astonishing. Most of the α-particles passed straight through the foil and struck the screen on the other side, just as Rutherford expected. However, occasionally an α-particle (one particle out of every several thousand) deflected with a small angle. And sometimes the deflected particles bounced almost straight back. From thus Rutherford theorized that the positive charge in the atom is all concentrated in a tiny nucleus and that most of the atom is empty space.

2. Describe the main points or features in the atomic models proposed by John Dalton, J.J. Thomson, and Ernest Rutherford.

Dalton's 1803 model contained the following five points:

1. All substances are composed of tiny, indivisible substances called atoms.
2. All atoms of the same substance are identical.
3. Atoms of different elements have different weights.
4. Atoms combine in whole-number ratios to form compounds.
5. Atoms are neither created nor destroyed in chemical reactions.

In 1897, Thomson theorized that electrons came from inside atoms. He developed a new atomic model that envisions atoms as tiny clouds of massless, positive charge sprinkled with thousands of the negatively charged electrons. Thomson's model is usually called the plum pudding model.

Rutherford's 1911 model contained the following points:

1. The positive charge in atoms is concentrated in a tiny region in the center of the atom, which Rutherford called the nucleus.
2. Atoms are mostly empty space.
3. The electrons, which contain the atoms' negative charge, are outside the nucleus.

3. Explain why Ernest Rutherford found the reflection of alpha particles off gold foil so astonishing.

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, but sometimes one of the particles would bounce almost straight back. Rutherford said this was like firing an artillery shell at a piece of tissue and having it bounce back.

SECTION 2.2

4. Write paragraphs distinguishing between these pairs of terms:

compounds and elements

mixtures and compounds

heterogeneous mixtures and homogeneous mixtures

suspensions and colloids

a. There are 118 elements, distinguished from one another by the number of protons in the nucleus. Compounds are formed when atoms of two or more elements are chemically bonded together.

b. Mixtures are formed when two or more different substances are combined without a chemical reaction occurring. If a chemical reaction does occur, then compounds are formed. The components of mixtures may be separated by physical means. The elements in compounds can only be separated by chemical means, that is, by a chemical reaction.

c. Homogeneous mixtures are solutions—one substance is dissolved in another. These mixtures have a uniform composition all the way down to the level of individual atoms or molecules. These individual particles are too small to be seen with any microscope. Heterogeneous mixtures contain larger lumps of material. These lumps are visible either with the eye or with a microscope.

d. In a suspension, larger particles are temporarily suspended in a fluid medium (liquid or gas), but the particles will settle out eventually. A colloid is formed when substances in two different physical states are combined.

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

a. water	**b. cesium chloride**	**c. pond water**	**d. methane**
e. a soft drink	**f. nitric acid**	**g. black coffee**	**h. argon**
i. air	**j. hydrogen nitrate**	**k. exhaust fumes**	**l. quartz**
m. brass	**n. hydrogen gas**	**o. hydrogen cyanide**	**p. mouthwash**
q. platinum	**r. dirt**	**s. radon**	**t. a smoothie**

a. compound	b. compound	c. heterogeneous mixture	d. compound
e. homogeneous mixture	f. homogeneous mixture (acids are typically solutions in water)	g. homogeneous mixture	h. element
i. solution	j. compound	k. heterogeneous mixture	l. compound
m. homogeneous mixture	n. element	o. compound	p. homogeneous mixture

q. element	r. heterogeneous mixture	s. element	t. heterogeneous mixture

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

The invisible particulates would be visible under a microscopes because they are really just tiny lumps. Since paint is not uniform down to the atomic level, it is not a solution.

7. Write a paragraph describing the two basic types of structures atoms can take when bonding together.

When bonding chemically, atoms form two basics types of structures—crystals and molecules. In crystals, atoms of two or more elements are arranged in a regular, repeating geometric pattern. In molecules, atoms bond together in small individual clusters.

8. Select three pure substances not mentioned in the chapter. For each substance, list at least eight physical properties and three chemical properties.

Example answers:

substance 1: lead (element)

physical properties: silver color, high density of 11.34 g/cm3, malleable, conducts electricity, melts at 327°C, boils at 1749°C, relatively soft, ductile

chemical properties: oxidizes to form lead oxide, non-flammable, reacts with fluorine and chlorine

substance 2: naphthalene (compound)

physical properties: strong odor (moth balls), density of 1.145 g/cm3, melts at 78.2°C, boils at 218°C, white solid at room temperature, can form as flakes, specific heat capacity of 165.7 J/m·K, nonconductive

chemical properties: molecule is made of two benzene rings joined together, reacts with chlorine to form chloronaphthalene, reacts with bromine to form bromonaphthalene

substance 3: nitrogen (element)

physical properties: gas at room temperature, condenses at –196°C, freezes at –210°C, colorless, density at STP of 1.25 g/L, odorless, heat of vaporization 5.56 kJ/mol, heat of fusion 0.72 kJ/mol

chemical properties: diatomic gas, does not burn, forms triple bonds

9. Explain why colloids reflect light. What is this effect called?

Particles in colloids are large enough to reflect light, even though they may not be otherwise visible. Headlight beams scattering in fog is an example of this effect, which is called the Tyndall effect.

10. Identify each of the following as a physical change or a chemical change. For each, explain your choice.

a. an avalanche	**b. a cigar burning**	**c. spilling a glass of milk**
d. digesting your food	**e. swatting a fly**	**f. stirring cream into coffee**
g. firing a pop gun	**h. firing a real gun**	**i. boiling mercury**
j. welding steel	**k. filling a helium balloon**	**l. allowing molten iron to harden**
m. frying chicken	**n. snow melting**	**o. a car exhaust pipe rusting**
p. paint "drying"	**q. wood rotting**	**r. a ball rolling down a hill**

a. physical	b. chemical	c. physical
d. chemical	e. physical	f. physical
g. physical	h. chemical	i. physical

j. physical	k. physical	l. physical
m. chemical	n. physical	o. chemical
p. chemical	q. chemical	r. physical

SECTION 2.3

11. How is the unified atomic mass unit, u, defined?

A mass of 1/12 the mass of an atom of carbon-12.

12. Referring to Table 2.6, calculate the atomic mass for silicon, calcium, and uranium. Compare your results to the values shown in the periodic table.

silicon:

0.92223·27.9769 + 0.04685·28.9765 + 0.03092·28.9738 = 28.0855 (Value in table 28.0855, so isotopes shown must be virtually all of the silicon.)

calcium:

0.96941·39.9626 + 0.00647·41.9586 = 39.012 (Value in table is 40.078, but percentages shown only add up to 97.588%, so there must be other heavier isotopes not listed.)

uranium:

0.007204·235.0439 + 0.992742·238.0508 = 238.016 (Value in table is 238.0289, so there must be a small amount of heavier isotopes not shown.)

13. Which two nuclides in Table 2.6 have 20 neutrons?

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

SECTION 2.4

25. How is the mole defined?

The new definition that went into effect in 2019 and will appear in the text in the next printing after 2020:

The amount of pure substance that contains Avogadro's number of particles.

The definition from the Second Edition text is:

A mole is the amount of substance that contains the same number of particles as there are atoms in 12 grams of carbon-12. Alternatively, a mole is the mount of pure substance that contains Avogadro's number of particles.

Chapter 3 Exercises

SECTION 3.1

1. What does it mean for an atom to be in the ground state?

It means all the electrons in the atom are in the lowest energy levels available.

2. What does it mean for an atom to be in an excited state, and in what ways can it occur?

When an atom is excited, one or more of its electrons are in energy levels higher than the lowest available energy level. This can occur when the atom absorbs a photon, or by the atom absorbing energy from a collision with another particle, such as an electron.

3. What conditions have to be met for a photon to ionize an atom?

Each electron has a specific negative amount of energy associated with the energy level it is in inside the atom. To ionize the atom means to give the electron the amount of energy necessary to bring its energy up to zero. For a photon to do this, its energy must be absorbed by the atom and its energy must be equal to the specific amount of energy the electron needs to have its energy raised to zero.

SECTION 3.2

8. 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.

9. 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 3.3

10. Describe the difference between the orbital energies in the hydrogen atom and those of other atoms.

In the hydrogen atom, all the orbitals with the same principle quantum number represent the same amount of energy. In all other atoms, different orbitals with the same principle quantum number represent different energies.

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

a. Identify the element this atom represents.

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

a. 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. 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.

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

a. Identify the element this atom represents.

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

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

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

13. 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 3.4 and a fuller answer to this question is given there. At this point in Section 3.3, 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 3.4, 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.

14. Describe the two main ways that atoms can possess energy.

One way is by the atom possessing kinetic energy. This can be manifest as vibrations or translations, or as it rotates inside a molecule. The second way is for one or more of the electrons to be in an energy level higher than the lowest energy level available. In this case, the atom is said to be in an excited state.

15. State and describe the four quantum numbers required to describe the quantum state of an electron.

principle quantum number—represents the main cluster of energy levels an electron is in.

azimuthal quantum number—designates the subshell (*s*, *p*, *d*, *f*, *g*) the electron is in.

magnetic quantum number—indicates the particular orbital an electron is in inside the subshell.

spin projection quantum number—indicates the spin of the electron (up or down).

16. For the first three principal quantum numbers, describe the orbitals available for electrons.

For $n = 1$, the 1*s* orbital is available.

For $n = 2$, 1*s*, 2*s*, $2p_x$, $2p_y$, and $2p_z$.

For $n = 3$, 1*s*, 2*s*, $2p_x$, $2p_y$, $2p_z$, 3*s*, $3p_x$, $3p_y$, and $3p_z$.

17. Write a description of the principles governing the quantum state of electrons in atoms.

The Aufbrau principle states that electrons fill up orbitals from lowest energy states to higher energy states. The Madelung rule specifies the order for increasing energy of the different orbitals, which is 1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, etc. The third rule is Hund's rule, which states that when equal energy orbitals are available, electrons fill each one up singly before any of them takes on a second electron. (Each orbital can hold at most two electrons.)

18. State the Pauli exclusion principle, and explain its relationship to the placement of electrons in atoms.

The Pauli exclusion principle states that no two electrons in an atom can occupy the same quantum state. This means that each electron in an atom has a unique combination of the four quantum numbers. Thus, electrons must be located in different orbitals in an atom, except for the two electrons that may occupy an orbital if they have opposite spins.

SECTION 3.4

21. 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.

22. For which group of elements does the electron configuration always end with np2? 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.

GENERAL REVIEW EXERCISES

39. Why is it important to show the correct number of significant digits in measurements?

The precision of a measurement is limited by the instrument used to make the measurement. This precision is represented in the number of significant digits the measurement has when it is written down. When dealing with measurements and calculations from them, the significant digits indicate how precise our values are.

40. Why is it important to show the correct number of significant digits in calculations performed on measurements?

As stated in the previous response, the significant digits indicate the precision we have in our measurements and calculations. In most cases, a calculation cannot increase precision because precision is set by the instrument used to make the measurements in the first place. The only exception is in the case of the addition rule, when values precise to the same place value are added to give a result that is still precise to that place value but has additional digits to the left due to the addition.

41. What is the definition for the unified atomic mass unit, u?

A mass of 1/12 the mass of an atom of carbon-12.

42. Distinguish between theories and hypotheses.

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. 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).

43. Distinguish between theories and truth.

Theories are mental models we use to describe nature. They are not truth claims; they are simply our best explanation at the time. They are provisional and subject to change and improvement. By contrast, truth is the way things really are and does not change.

44. Distinguish between matter and mass.

Matter consists of anything composed of atoms or parts of atoms. Mass is a variable used to quantify the amount of inertia (a property of matter) an object has.

45. Why is the Kelvin temperature scale referred to as an "absolute" scale?

Because there are no negative values on the Kelvin scale, so the 0 of the scale is at the theoretical lowest temperature.

Course Lesson Schedule General Chemistry 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) Measurements, Mass	1.1.1-1.1.2		Quiz 1
5	Metric System, Unit Conversions	1.1.3-1.2.3	1.1-1.2 exercises	
6	Quiz 2 (1.1-1.2.2) Accuracy, Precision, % Diff	1.3-1.4	1.3-1.4 exercises	Quiz 2
7	History of Atomic Models	2.1.1-2.1.2	2.1 exercises	
8	Quiz 3 (1.2.3) Substances	2.2.1-2.2.3	2.2 exercises	Quiz 3
9	Discussion			
10	Test Intro and Chapter 1			Test Chapter 1
11	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)*
12	Lab Day: Identification of substances by physical properties			
13	Isotopes and Atomic Mass	2.3.1-2.3.3	2.3 exercises	
14	Quiz 4 (2.1-2.2.2) Density	2.4.1		Quiz 4
15	Lab Day: Separation of components in mixtures		Lab Report	This is Experiment 2 in *CEHS* and *CEHSH*
16	Mole, Avogadro Constant, Molar mass	2.4.2-2.4.3		
17	Discussion			
18	Quiz 5 (2.2.3-2.3) masses of atoms and molecules	2.4.4	2.4 exercises	Quiz 5
19	Electromagnetic Spectra, Planck relation	3.1.1	begin ch 3 general review exercises	
20	Discussion			
21	Test Chapter 2			Test Chapter 2
22	Lab Day: Flame tests and metal cation identification		Lab Report	This is Experiment 3 in *CEHS* and *CEHSH*
23	Hydrogen atom, Bohr model	3.1.2-3.2	3.1-3.2 exercises	
24	Quantum atomic model	3.3.1-3.3.2		
25	Quiz 6 (3.1-3.2) Quantum atomic model	3.3.2		Quiz 6
26	Aufbau principle, Madelung rule, Hund's rule	3.3.3	3.3 exercises	
27	Discussion			
28	Electron configurations	3.4	3.4 exercises	
29	Empirical formulas	3.5	3.5 exercises	
30	Quiz 7 (3.3-3.4) Periodic Table	4.1-4.2	4.1-4.2 exercises begin ch 4 general review exercises	Quiz 7
31	Lab Day: Determining an empirical formula		Lab Report	This is Experiment 4 in *CEHS* and *CEHSH*
32	Lab Day: Determining an empirical formula (continued)			
33	Discussion			
34	Test Chapter 3			Test Chapter 3
35	Physical periodic properties	4.3.1-4.3.2	4.3 exercises	
36	core/valence electrons, ionization energy	4.4.1-4.4.2		
37	Quiz 8 (4.1-4.3) electron affinity	4.4.3		Quiz 8
38	electronegativity, hydrogen	4.4.4-4.5	4.4-4.5 exercises	
39	Octet rule, ionic bonds	5.1.1-5.2.1	5.1 exercises	
40	Quiz 9 (4.4) Naming ionic compounds	5.2.2	begin ch 5 general review exercises	Quiz 9
41	Discussion			
42	Test Chapter 4			Test Chapter 4
43	Energy, hydrates, intensive & extensive properties	5.2.3-5.2.5		
44	Discussion			
45	Ionic physical properties	5.2.6	5.2 exercises	
46	Covalent bonds, molecules	5.3.1		
47	Quiz 10 (5.1-5.2) Polyatomic ions	5.3.2	begin 5.3 exercises	Quiz 10
48	Polyatomic ion names, acid names	5.3.4-5.3.5		
49	Lewis structures	5.3.6		
50	Discussion			
51	Octet Exceptions, Resonance structures	5.3.7-5.3.8		
52	Naming binary compounds, energy in covalent bonds, physical properties	5.3.9-5.3.11	5.3 exercises	
53	Polarity and dipoles; nature of the bond	5.4	5.4 exercises	

54	Quiz 11 (5.3)			Quiz 11
55	Discussion			
56	Test Chapter 5			Test Chapter 5
57	Covalent Bond theory and VSEPR theory	6.1.1-6.1.2	begin ch 6 general review exercises	
58	Nonbonding domains and bond angle	6.1.3	6.1 exercises	
59	Discussion			
60	Metallic bonding	6.2.1		
61	Quiz 12 (6.1) Physical properties of metals	6.2.2	6.2 exercises	Quiz 12
62	Intermolecular forces	6.3.1-6.3.2		
63	Quiz 13 (6.2) Hydrogen bonding	6.3.3	begin 6.3 exercises	Quiz 13
64	Van der Waals forces	6.3.4	6.3 exercises	
65	Lab Day: Intermolecular Forces		Lab Report	This is Experiment 7 in *CEHS* and *CEHSH*
66	Discussion			
67	Test Chapter 6			Test Chapter 6
68	Review			
69	Review			

Course Lesson Schedule General Chemistry Spring Semester

Lesson Number	Topic	Text Key	Assignment	Notes
1	Law of conservation of mass and balancing equations	7.1.1-7.1.4	begin ch 7 general review exercises	
2	Oxidation states	7.1.5	7.1 exercises	
3	Reaction types, activity series of metals	7.2.1-7.2.6		
4	Quiz 14 (7.1) Reaction types	7.2.7-7.2.8	7.2 exercises	Quiz 14
5	Discussion			
6	Lab Day: Activity Series of Metals		Lab Report	This is Experiment 5 in *CEHS* and *CEHSH*
7	Stoichiometry	7.3.1		
8	Limiting reagent, theoretical yield, percent yield	7.3.2-7.3.3	7.3 exercises	
9	Quiz 15 (7.2)			Quiz 15
10	Lab Day: Limiting Reactant and Percent Yield		Lab Report	This is Experiment 6 in *CEHS* and *CEHSH*
11	Discussion			
12	Test Chapter 7			Test Chapter 7
13	Kinetic molecular theory of gases	8.1.1-8.1.3	begin ch 8 general review exercises	
14	Gas pressure, pressure units	8.1.4	8.1 exercises	
15	States of matter	8.2.1-8.2.5		
16	Quiz 16 (8.1) Phase transitions	8.2.6	begin 8.2 exercises	Quiz 16
17	Heat capacity, heat of fusion/vaporization	8.2.7		
18	Discussion			
19	Evaporation, vapor pressure	8.2.8-8.2.9	8.2 exercises	
20	Quiz 17 (8.2)			Quiz 17
21	Discussion			
22	Test Chapter 8			Test Chapter 8
23	Lab Day: Metathesis Reactions		Lab Report	This is Experiment 10 in *CEHS* and *CEHSH*
24	Boyle's law, Charles' law, Avogadro's law	9.1.1-9.1.3	9.1 exercises	
25	STP, ideal gas law	9.2.1-9.2.2	begin ch 9 general review exercises	
26	Quiz 18 (9.1) ideal gas law	9.2.2	begin 9.2 exercises	Quiz 18
27	Discussion			
28	Using ideal gas law to find molar mass and density	9.2.3	9.2 exercises	
29	Quiz 19 (9.2) Law of partial pressure	9.3.1		Quiz 19
30	Collecting gas over water	9.3.2	9.3 exercises	
31	Stoichiometry of gases	9.4.1	begin 9.4 exercises	
32	Lab Day: Mole Amount of a Gas		Lab Report	This is Experiment 9 in *CEHS* and *CEHSH*
33	Diffusion and effusion	9.4.2	9.4 exercises	
34	Discussion			
35	Test Chapter 9			Test Chapter 9
36	Dissolution	10.1	begin ch 10 general review exercises	
37	Heat of solution, entropy, free energy	10.1.2-10.1.3		
38	Electrolytes	10.1.4	10.1 exercises	
39	Solubility of ionic solids	10.2.1-10.2.2		
40	Quiz 20 (10.1)			Quiz 20
41	Solutions of polar and nonpolar liquids	10.2.3-10.2.4	begin 10.2 exercises	
42	Solid solutions	10.2.5		
43	Gases in solution, temperature effect on solubility	10.2.6-10.2.7	10.2 exercises	
44	Discussion			
45	Molarity, molality	10.3.1-10.3.2	10.3 exercises	
46	Ionic equations, precipitates	10.4.1		
47	Net ionic equations, spectator ions	10.4.2	10.4 exercises	
48	Quiz 21 (10.2-10.3) Colligative properties (intro)	10.5		Quiz 21
49	Lab Day: Molarity		Lab Report	This is Experiment 8 in *CEHS* and *CEHSH*
50	Vapor pressure lowering	10.5.1		
51	Boiling point elevation, freezing point depression	10.5.2	10.5 exercises	
52	Discussion			
53	Test Chapter 10			Test Chapter 10
54	Acid base intro, names, formulas	11.1.1-11.1.3	11.1 exercises begin ch 11 general review exercises	
55	Acid-base theories: Arrhenius	11.2.1		
56	Acid-base theories: Bronsted-Lowry	11.2.2		
57	Bronsted-Lowry, Lewis, acid-base strength	11.2.3-11.2.4	11.2 exercises	
58	Discussion			
59	Self-ionization of water, $[H_3O^+]$, $[OH^-]$	11.3.1,	begin 11.3 exercises	

		11.3.2		
60	Quiz 22 (11.1-11.2) pH	11.3.3		Quiz 22
61	pH	11.3.3		
62	pH	11.3.3		
63	pH indicators, titration	11.3.4		
64	Quiz 23 (11.2-11.3.2) Titration	11.3.5		Quiz 23
65	Titration	11.3.5		
66	Determining [H_3O^+] and [OH^-] from titration data	11.3.6	11.3 exercises	
67	Lab Day: Acid-base titration		Lab Report	This is Experiment 11 in *CEHS* and *CEHSH*
68	Discussion			
69	Lab Day: Effectiveness of antacids		Lab Report	This is Experiment 12 in *CEHS* and *CEHSH*
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	Redox half-reactions	12.2.1		
74	Quiz 24 (12.1) Redox half-reactions	12.2.1	begin 12.2 exercises	Quiz 24
75	Balancing redox equations	12.2.2		
76	Balancing redox equations	12.2.2	12.2 exercises	
77	Discussion			
78	Lab Day: Redox titration		Lab Report	This is Experiment 17 in *CEHS* and 16 *CEHSH*
79	Quiz 25 (12.2)			Quiz 25
80	Electrochemistry	12.3.1-12.3.2	begin 12.3 exercises	
81	Electrochemical cells	12.3.3		
82	Electrode potentials	12.3.4		
83	Lab Day: Electrochemical series		Lab Report	This is Experiment 18 in *CEHS* and 17 *CEHSH*
84	Electrochemical applications	12.3.5	12.3 exercises	
85	Discussion			
86	Test Chapter 12			Test Chapter 12

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General Chemistry Quiz 8: Chapter 4 (4.1–4.3)

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. Who is usually credited with discovering the Periodic Table of the Elements, and why? (25)

2. On the periodic table on the other side, identify the following: (25)
   a. transition metals
   b. inner transition metals
   c. noble gases
   d. alkali metals
   e. halogens
   f. alkaline-earth metals
   g. chalcogens

3. What is the chief chemical property that differentiates the metals from the nonmetals? (25)

4. Describe the trends in atomic size in the periodic table, and briefly describe three factors that influence atomic size. (25)

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H Hydrogen 1.0079																	2 He Helium 4.0026
3 Li Lithium 6.941	4 Be Beryllium 9.0122											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.0067	8 O Oxygen 15.9994	9 F Fluorine 18.9984	10 Ne Neon 20.1797
11 Na Sodium 22.9898	12 Mg Magnesium 24.3050											13 Al Aluminum 26.9815	14 Si Silicon 28.0855	15 P Phosphorus 30.9738	16 S Sulfur 32.066	17 Cl Chlorine 35.4527	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.9559	22 Ti Titanium 47.88	23 V Vanadium 50.9415	24 Cr Chromium 51.9961	25 Mn Manganese 54.9380	26 Fe Iron 55.847	27 Co Cobalt 58.9332	28 Ni Nickel 58.6934	29 Cu Copper 63.546	30 Zn Zinc 65.39	31 Ga Gallium 69.723	32 Ge Germanium 72.61	33 As Arsenic 74.9216	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.80
37 Rb Rubidium 85.468	38 Sr Strontium 87.62	39 Y Yttrium 88.9059	40 Zr Zirconium 91.224	41 Nb Niobium 92.9064	42 Mo Molybdenum 95.94	43 Tc Technetium 98.9072	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.9055	46 Pd Palladium 106.42	47 Ag Silver 107.8682	48 Cd Cadmium 112.411	49 In Indium 114.82	50 Sn Tin 118.710	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.9045	54 Xe Xenon 131.29
55 Cs Cesium 132.905	56 Ba Barium 137.327	71 Lu Lutetium 174.967	72 Hf Hafnium 178.49	73 Ta Tantalum 180.9479	74 W Tungsten 183.85	75 Re Rhenium 186.207	76 Os Osmium 190.2	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.9665	80 Hg Mercury 200.59	81 Tl Thallium 204.3833	82 Pb Lead 207.2	83 Bi Bismuth 208.9804	84 Po Polonium 208.9824	85 At Astatine 209.9871	86 Rn Radon 222.0176
87 Fr Francium 223.0197	88 Ra Radium 226.0254	103 Lr Lawrencium 262.11	104 Rf Rutherfordium 261.11	105 Db Dubnium 262.114	106 Sg Seaborgium 263.118	107 Bh Bohrium 262.12	108 Hs Hassium (265)	109 Mt Meitnerium (266)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (281)	112 Cn Copernicium (285)	113 Uut Ununtrium (284)	114 Fl Flerovium (289)	115 Uup Ununpentium (288)	116 Lv Livermorium (293)	117 Uus Ununseptium (294)	118 Uuo Ununoctium (294)

57 La Lanthanum 138.9055	58 Ce Cerium 140.115	59 Pr Praseodymium 140.9077	60 Nd Neodymium 144.24	61 Pm Promethium 144.9127	62 Sm Samarium 150.36	63 Eu Europium 151.965	64 Gd Gadolinium 157.25	65 Tb Terbium 158.9253	66 Dy Dysprosium 162.50	67 Ho Holmium 164.9303	68 Er Erbium 167.26	69 Tm Thulium 168.9342	70 Yb Ytterbium 173.04
89 Ac Actinium 227.0278	90 Th Thorium 232.0381	91 Pa Protactinium 231.0359	92 U Uranium 238.0289	93 Np Neptunium 237.0482	94 Pu Plutonium 244.0642	95 Am Americium 243.0614	96 Cm Curium 247.0703	97 Bk Berkelium 247.0703	98 Cf Californium 251.0796	99 Es Einsteinium 252.083	100 Fm Fermium 257.0951	101 Md Mendelevium 258.10	102 No Nobelium 259.1009

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General Chemistry Quiz 12: Chapter 6 (6.1)

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. Distinguish between bonding domains and nonbonding domains. (20)

2. Briefly describe VSEPR theory. In your answer, state what the letters in the acronym mean. (20)

3. For the following molecules, draw the Lewis structure and identify the electron domain geometry and the molecular geometry (5 points per cell in the table)

formula	Lewis structure	electron domain geometry	molecular geometry
SiF_4			
H_2S			
AlF_3			

4. Briefly explain why the angle in the water molecule is less than 109.5°. (15)

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General Chemistry Quiz 21: Chapter 10 (10.2–10.3)

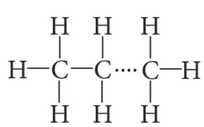
Point values for each item are shown in parentheses. Questions on this quiz may be answered a few sentences. Write all responses in complete sentences using correct spelling and grammar.

1. Consider a saturated solution of sugar in water with excess sugar at the bottom of the container. Describe the equilibrium that exists in this solution. (28)

2. Identify the following compounds as soluble or insoluble in water. (2 points each, 18 total)

barium carbonate	Li_2CO_3	$Cu(OH)_2$
PbI_2	CaS	ammonium phosphate
$Mg(NO_3)_2$	potassium sulfate	$ZnCl_2$

3. Water and liquid HF are good solvents for many ionic salts. However, oils and waxes with structures such as shown in the accompanying Lewis structure do not dissolve in water. Explain why. (27)



4. 259.5 g of potassium chloride are dissolved in enough water to make 992 mL of solution. Determine the molarity of the solution. (27)

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General Chemistry Quiz 25: Chapter 12 (12.2)

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. Complete and balance each of the following half-reactions. In each case, indicate if the reaction is an oxidation or reduction. (25 each)

    a. $ClO_3^-(aq) \rightarrow Cl^-(aq)$ (acidic solution)

    b. $N_2(g) \rightarrow NH_3(g)$ (basic solution)

2. Use the half-reaction method to balance the following redox reaction, which takes place in acidic solution: (50)

    $I_2(s) + HClO(aq) \rightarrow HIO_3(aq) + HCl(aq)$