GET 800
CHOOSE YOUR COLLEGE

320 SAT CHEMISTRY SUBJECT TEST PROBLEMS

arranged by **Topic** and **Difficulty** Level

Christopher Bozza
Dr. Steve Warner

320 Level 1, 2, 3, 4, and 5 Chemistry Problems

Legal Notice

320 SAT Chemistry Subject Test Problems arranged by Topic and Difficulty Level

320 Level 1, 2, 3, 4, and 5
Chemistry Problems

Christopher Bozza
Dr. Steve Warner

BOOKS BY DR. STEVE WARNER FOR COLLEGE BOUND STUDENTS

28 New SAT Math Lessons to Improve Your Score in One Month
- Beginner Course
- Intermediate Course
- Advanced Course

New SAT Math Problems arranged by Topic and Difficulty Level
320 SAT Math Problems arranged by Topic and Difficulty Level
SAT Verbal Prep Book for Reading and Writing Mastery
320 SAT Math Subject Test Problems
- Level 1 Test
- Level 2 Test

The 32 Most Effective SAT Math Strategies (old SAT)
SAT Prep Official Study Guide Math Companion (old SAT)
Vocabulary Builder
320 ACT Math Problems arranged by Topic and Difficulty Level
320 GRE Math Problems arranged by Topic and Difficulty Level
320 AP Calculus AB Problems
320 AP Calculus BC Problems
Physics Mastery for Advanced High School Students
SHSAT Verbal Prep Book to Improve Your Score in Two Months
555 Math IQ Questions for Middle School Students
555 Advanced Math Problems for Middle School Students
555 Geometry Problems for High School Students
Algebra Handbook for Gifted Middle School Students

CONNECT WITH DR. STEVE WARNER

Table of Contents

INTRODUCTION
THE PROPER WAY TO PREPARE

There are many ways that a student can prepare for the SAT Chemistry Subject Test. But not all preparation is created equal. With my own students, I always emphasize the methods that will give the maximum result with the minimum amount of effort.

The book you are now reading is self-contained. Each problem was carefully created to ensure that you are making the most effective use of your time while preparing for the test. By grouping the problems given here by level and topic, we have ensured that you can focus on the types of problems that will be most effective to improving your score.

This book can be utilized at any point in your preparation for the SAT Chemistry Subject Test. Early in your studies going over the Level 1 and 2 problems and reading the remarks can help you identify and review important concepts that went by quickly in class. After several weeks of studying, if you feel comfortable with the topics at the 2^{nd} and 3^{rd} Levels you can be confident you are on track. As a final review in the last weeks before you take the SAT Chemistry Subject Test, you can review definitions and get help tackling the Level 4 and Level 5 difficulty problems. Alternately, if you're in the last stretch of preparation and need to focus on practice problems, this book will help you to hone in on those problems and concepts at whatever difficulty level and areas you need. I generally recommend four to six months of preparation for the SAT Chemistry Subject Test. You can minimize this preparation time by planning to take the test in the Spring of the year when you take Chemistry. Your classwork will then help to reinforce your studying, build in structured practice, and review for most, if not all, areas of the test. Depending on how closely your teacher's lessons match the material for the exam, you could start preparing for the Chemistry SAT 3 months before your test date.

1. Using this book effectively

- Begin studying at least one month before test day.
- Practice Chemistry Subject Test problems twenty minutes each day.
- Choose a consistent study time and location.

You will retain much more of what you study if you study in short bursts rather than if you try to tackle everything at once. So try to choose about a twenty-minute block of time that you will dedicate to the SAT Chemistry Subject Test each day. Make it a habit. The results are well worth this small time commitment.

- Every time you get a question wrong, **mark it off, no matter what your mistake**.
- Begin each study session by first redoing problems from previous study sessions that you have marked off and then reviewing the remarks.
- If you get a problem wrong again, **keep it marked off**.

Being able to solve any specific problem is of minimal importance. Understanding the concepts that enable you to solve each problem will help you to be more prepared, help you to tackle a problem you have never seen before, and allow you to solve similar problems more quickly. To support this, read all of the remarks following each problem. Remarks explain and expand upon important concepts in each question. However, if your reviewing phase is done and you’re prioritizing practice questions, note that the quickest solution to any problem will always be marked with an asterisk (*).

2. The magical mixture for success

A combination of three components will maximize your SAT Chemistry Subject Test score with the least amount of effort.

- Learning test-taking strategies that work specifically for standardized tests.
- Practicing SAT Chemistry Subject Test problems for a small amount of time each day for at least three months before the test.
- Taking about two practice tests before test day to make sure you are applying the strategies effectively under timed conditions.

I will discuss each of these three components in a bit more detail.

Strategy: The more SAT specific strategies you know, the better off you will be. Throughout this book you will see many strategies being used. Some examples of basic strategies are "balancing equations using boxes," "using $M_1V_1 = M_2V_2$," and "picking numbers." Some more advanced strategies include "mole map conversions," and "determining rate law." Pay careful attention to as many strategies as possible and try to internalize them. Even if you do not need to use a strategy for that specific problem, you will certainly find it useful for other problems in the future. Roughly 20% of the problems on the SAT Chemistry Subject Test will require you to simply know and understand terminology. Roughly 45% of the test will require you to know the terminology for a single concept being discussed and apply it to a specific scenario. To achieve mastery for each of these types of problems, it is important to read and understand the remarks and definitions associated with each problem. The remaining 35% of the test will require you to process multiple concepts, which means the previous two problem types should be mastered first in order to perform well on these harder problems.

Practice: The problems given in this book are more than enough to vastly improve your current SAT Chemistry Subject Test score. All you need to do is work on these problems for about twenty minutes each day over a period of at least three months and the final result will far exceed your expectations.

Let me further break this component into two subcomponents – **topic** and **level**.

Topic: There are a total of 15 topics on the test. We have grouped these into four general subject areas that contain similar concepts in order to focus your studying. The objective is to practice each of the four general chemistry subject areas given on the SAT chemistry test and improve in each independently. The four subject areas with their topics are:

Structure of Matter

- Atomic structure, Molecular Structure, Bonding

States and Reactions of Matter

- Gases, Liquids & Solids, Solutions

Product and Reactant Relationships

- Acids & Bases, Oxidation-Reduction, Precipitation, Mole Concept, Chemical Equations, Equilibrium Systems, Rates of Reaction, Thermochemistry

Descriptive Chemistry & Laboratory

- Descriptive Chemistry & Laboratory

Level: You will make the best use of your time by primarily practicing problems that are at and slightly above your current ability level. For example, if you are struggling with Level 2 Structure of Matter problems, then it makes no sense at all to practice Level 5 Structure of Matter problems. Keep working on Level 2 until you are comfortable, and then slowly move up to Level 3. Maybe you should never attempt those Level 5 problems. You can actually get an 800 on the Chemistry Subject Test without answering any of them.

Tests: You want to take about two practice tests before test day to make sure that you are implementing strategies correctly and using your time wisely under pressure. For this task you should use the only available authentic SAT Chemistry Subject Test found in "*The Official Study Guide for All SAT Subject Tests, Second Edition*." Additional practice questions can be found on the College Board website.

3. Practice problems of the appropriate level

In each of the two parts of this book the test questions have been split into 5 levels. Roughly speaking, the questions increase in difficulty as you progress from Level 1 to Level 5. The first 32 problems are Level 1, the next 32 are Level 2 and so on.

Keep track of your current ability level so that you know the types of problems you should focus on. If you are still getting most Level 2 States of Matter questions wrong, then do not move on to Level 3 States of Matter until you start getting more Level 2 States of Matter questions right on your own.

If you really want to refine your studying, then you should keep track of your ability level in each of the four major categories of problems:

- **Structure of Matter**
- **States and Reactions of Matter**
- **Product and Reactant Relationships**
- **Descriptive Chemistry & Laboratory**

For example, if you are stronger in Product and Reactant Relationships than Descriptive Chemistry & Laboratory, then it might make sense to practice Level 4 Product and Reactant Relationships problems while you continue to practice Level 2 Descriptive Chemistry & Laboratory problems.

4. Practice a small amount every day

Ideally you want to practice doing SAT Chemistry Subject Test problems about twenty minutes each day beginning at least three months before the exam. You will retain much more of what you study if you study in short bursts than if you try to tackle everything at once.

The only exception is on a day you do a practice test. You should do at least two practice tests before you take the test. Ideally you should do your practice tests on a Saturday or Sunday morning.

So try to choose about a twenty-minute block of time that you will dedicate to practice each day. Make it a habit. The results are well worth this small time commitment.

5. Redo the problems you get wrong over and over and over until you get them right

If you get a problem wrong, and never attempt the problem again, then it is extremely unlikely that you will get a similar problem correct if it appears on the SAT Chemistry Subject Test.

Most students will read an explanation of the solution, or have someone explain it to them, and then never look at the problem again. This is *not* how you optimize your score. To be sure that you will get a similar problem correct on the actual test, you must get the practice problem correct before taking the real test—and without actually remembering the problem.

This means that after getting a problem incorrect, you should go over and understand why you got it wrong, read all the remarks, wait at least a few days, then attempt the same problem again. If you get it right, you can cross it off your list of problems to review. If you get it wrong, keep revisiting it every few days until you get it right. Your score *does not* improve by getting problems correct. **Your score improves when you learn from your mistakes.**

6. Check your answers properly

When you go back to check your earlier answers for careless errors *do not* simply look over your work to try to catch a mistake. This is usually a waste of time. Always redo the problem and re-read the answers without looking at any of your previous work. When applicable, you want to use a different method than you used the first time, or evaluate answers in the reverse direction.

For example, if you solved the problem by picking numbers the first time, try to solve it algebraically the second time, or at the very least pick different numbers. If you do not know, or are not comfortable with a different method, then use the same method, but do the problem from the beginning and do not look at your original solution. If your two answers do not match up, then you know that this is a problem you need to spend a little more time on to figure out where your error is.

This may seem time consuming, but that's okay. It is better to spend more time checking over a few problems than to rush through a lot of problems and repeat the same mistakes.

7. Guess when appropriate

Answering a multiple choice question wrong will result in a $\frac{1}{4}$ point penalty. This penalty is to discourage random guessing. If you have no idea how to do a problem, no intuition as to what the correct answer might be, and you cannot even eliminate a single answer choice, then *DO NOT* just take a guess. Omit the question and move on.

If, however, you can eliminate even one answer choice, you should take a guess from the remaining four. You should of course eliminate as many choices as you can before you take your guess. This strategy generally leads to higher scores than avoiding guesses for any question.

8. Pace yourself

Do not waste your time on a question that is too hard or will take too long. After you've been working on a question for about 30 to 45 seconds you need to make a decision. If you understand the question and think that you can get the answer in another 30 seconds or so, continue to work on the problem. If you still do not know how to do the problem or you are using a technique that is going to take a long time, mark it off and come back to it later if you have time.

If you do not know the correct answer, but you can eliminate at least one answer choice, then take a guess. But you still want to leave open the possibility of coming back to it later. Remember that every problem is worth the same amount. Do not sacrifice problems that you may be able to do by getting hung up on a problem that is too hard for you.

9. Attempt the right number of questions

Many students make the mistake of thinking that they have to attempt every single question on the SAT Subject Tests. There is no such rule. In fact, most students will increase their score by *reducing* the number of questions they attempt.

This is particularly true for the relationship analysis questions. This question type is a difficult one for many students since each question is a set of three binary options: two true/false statements and a determination of the statements' relationship. If you are unsure whether both of the two statements are true or false, you should not guess.

10. Understand the relationship analysis questions before you take the test

The set of relationship analysis questions are one of the defining features of the SAT chemistry subject test. Since this is a new question type for most students, it is important to learn how to recognize, evaluate, and when necessary, guess most effectively on it. Relationship analysis questions appear as a pair of statements as seen in the following example:

I		II
Atoms are the smallest units of matter that cannot be split and retain their properties.	BECAUSE	Protons are found in the nucleus of the atom.

There are two statements labeled I and II, each of which will have a corresponding pair of True/False bubbles on the grid-in sheet. The above example would have statement I marked as T and statement II marked as T. Once both statement answers are bubbled in, you evaluate whether statement II is a correct explanation of statement I. If it is, you will fill in the bubble labeled CE (Correct Explanation). This example does not have a correct explanation and would not have CE selected.

There are several important characteristics to know about these questions. First, <u>no question with a false statement I or II can be a CE</u>. Only questions that have two true statements can possibly be a CE. Second, there is no bubble for 'not CE' on the answer key, you simply bubble in CE or leave it empty. This leads to 5 possible answer choices:

- TT, CE
- TT, Not CE
- FT
- TF
- FF

These 5 choices effectively mean that relationship analysis questions are really (A) through (E) multiple choice questions with a different look! Third, if you are unsure whether both statements are true or false, you should not guess. However, if you are sure one of the statements is true, you should guess the remaining statement as T or F. Then, if you guess the remaining statement to be true, you will need to guess at the CE as well. All of these choices come together for a 33% chance at a correct response when you know that one statement is true. On the other hand, if you are sure that one of the statements is false you have a 50% chance to guess the correct answer by putting T or F for the other statement. You should definitely guess the remaining statement when you know one statement is false. Only in the case where you are unsure about both statements should you leave these questions blank.

PROBLEMS BY LEVEL AND TOPIC WITH FULLY EXPLAINED SOLUTIONS

Note: The quickest solution will always be marked with an asterisk *.

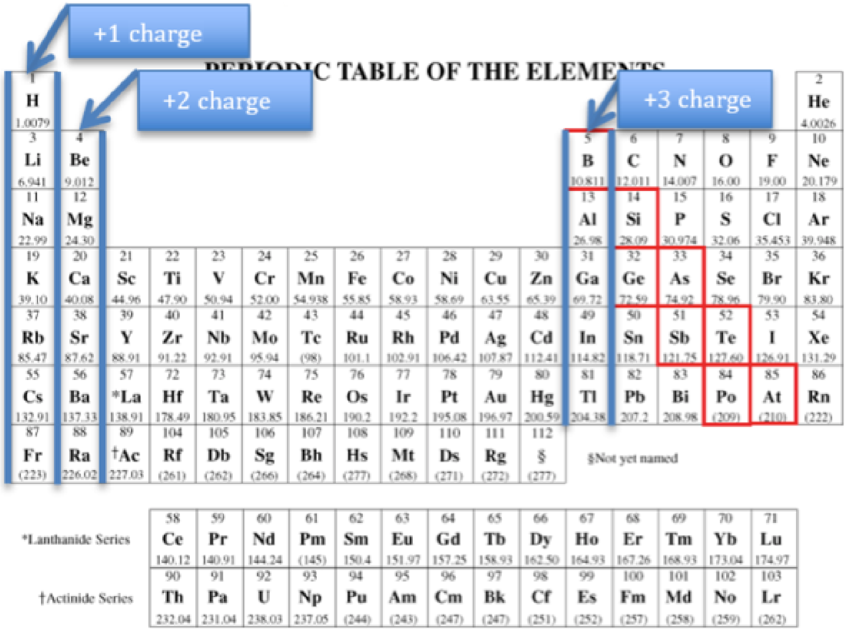
LEVEL 1: DESCRIPTIVE CHEMISTRY AND LABORATORY

1. Identify a type of ion found in magnesium nitride from the following species of ions.

   (A) X^{+}
   (B) X^{2+}
   (C) X^{3+}
   (D) XO_3^{2-}
   (E) XO_4^{3-}

***Solution:** This ionic compound has a magnesium cation (Mg^{2+}) and a nitrogen anion (N^{3-}). By comparing each answer choice to Mg^{2+} and N^{3-}, we see that answer choice (B) has the positive charge matching magnesium.

Remarks: (1) Group 1 metals (alkali metals, first column) all have ions with a +1 charge, group 2 metals (alkaline earth metals, second column) all have ions with a +2 charge. Group 3A metals all form ions with a +3 charge. Below are columns of ions that could represent each of the answer choices A through C.



PERIODIC TABLE OF THE ELEMENTS

1 H 1.0079																	2 He 4.0026
3 Li 6.941	4 Be 9.012											5 B 10.811	6 C 12.011	7 N 14.007	8 O 16.00	9 F 19.00	10 Ne 20.179
11 Na 22.99	12 Mg 24.30											13 Al 26.98	14 Si 28.09	15 P 30.974	16 S 32.06	17 Cl 35.453	18 Ar 39.948
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.938	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.91	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 *La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.2	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.02	89 †Ac 227.03	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (277)	109 Mt (268)	110 Ds (271)	111 Rg (272)	112 § (277)	§Not yet named					

*Lanthanide Series	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.4	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
†Actinide Series	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

(2) There is no “X” abbreviation on the periodic table. When X is used in a formula, it indicates a general class or type of element. However, do watch out for the noble gas Xenon: Xe.

(3) This problem requires you to use your knowledge of ionic compounds and the provided periodic table. Magnesium nitride is composed of a magnesium cation and a nitride anion. Magnesium is found in group 2, and all group 2 metals form ions with a +2 charge. Nitrogen is found in group 5 and all group 5 non-metals form anions with a -3 charge. For a further discussion of why this happens see problem 59.

(4) X^{3-} would be another acceptable answer to this question, but it does not appear as an answer choice.

(5) Magnesium nitride is a **binary ionic compound**. The word *binary* means ‘having two parts,’ and therefore every binary ionic compound contains only two elements. An easy way to identify a binary ionic compound is to confirm that it contains only two elements and that one is a metal and the other is a nonmetal. Some ionic compounds contain a polyatomic anion or cation, and will have three or more elements. Examples and descriptions of these can be found in problem 33.

(6) The advantage of knowing whether something is a binary ionic compound is that we can determine the charge of the metal and nonmetal pretty easily based on the columns of the elements. The periodic table above shows the columns where positive ions of +1, +2, and +3 can be found. Negative ions can be found in the columns to the right of carbon. Nitrogen's column is -3, oxygen's column is -2, and fluorine's column is -1. When an ionic charge is not clear for the metals found in the middle of the periodic table, we can determine it using a commonly called 'criss-cross method' or a charge balancing method. For an example where this method is employed, see problem 102, remark (4).

(7) Answer choice (E) has the same charge as a nitrogen ion, however choice (E) also includes oxygen atoms which magnesium nitride does not have. Answer choice (D) is also eliminated because it contains oxygen in addition to having a -2 charge, which is not found on either ion in the problem. For patterns in naming related oxygen-containing polyatomic ions see problem 116. For a list of common polyatomic ions (and for the formal definition of 'polyatomic ion') see problem 33.

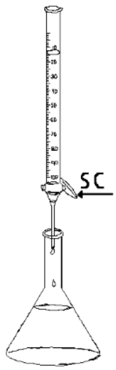
Definitions: Ionic compounds are always composed of an **anion** and a **cation**. **Ions** are atoms that have either gained electrons to become negatively charged **anions**, or lost electrons to become positively charged **cations**. A **binary ionic compound** is an ionic compound that contains exactly two elements.

Metals are found to the left of the staircase elements, **semi-metals** (also known as **metalloids**) make up the staircase elements (outlined in the periodic table above), and **non-metals** are found to the right of the staircase elements. In a binary ionic compound, the metal forms the cation, or positive ion, and the non-metal forms the anion, or negative ion.

2.

I		II
If a buret's volume can be accurately read to the nearest 0.001 mL, then when removing exactly 2 mL of a solution, the recorded change in volume should be 2.000.	BECAUSE	Volumes of liquids are recorded to three decimal places in standard laboratory procedure.

***Solution:** When reading a buret, you always record to the most accurate measurement possible, so statement I is **True**. The accuracy of a piece of measuring equipment is determined by its markings, so II is **False**. With a False statement this cannot be a CE.



Remarks: (1) A **buret** is a piece of laboratory equipment that can be used to accurately add small volumes of liquid to a container. It is marked with lines to measure volume and has an open top and a narrow bottom. To the right is a picture of a buret dispensing a solution into an **Erlenmeyer flask** with the **stopcock** labeled SC.

A buret is generally used for **titration** experiments.

(2) When taking measurements using a piece of laboratory equipment, the measurement should have as many decimal places as the container is listed as having plus one additional estimated digit. For example, if a graduated cylinder is accurate to the 0.001 place, then a valid measurement could be 0.3255 or 0.9882. Both numbers go one spot past the listed accuracy and are valid estimates. This problem is worded to say "accurately read to the nearest 0.001 mL" which means the last digit is estimated and the markings are in increments of 0.01 mL.

Definitions: A **buret** looks like a graduated cylinder with a **stopcock** at the bottom. The buret dispenses its solution from the bottom due to gravity. The **stopcock** is a valve with a small handle that can be opened or closed like a faucet to control the amount of volume that is dispensed. An **Erlenmeyer flask** is used to hold solutions and has a cone shaped bottom portion with a narrow neck as seen above.

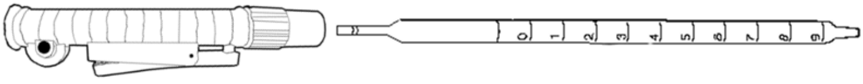
Titration is a laboratory method that involves adding a solution with a known concentration to a second solution of an unknown concentration. The goal is to determine the number of moles of the unknown chemical concentration and therefore the molarity of the second solution. The most common titration you will encounter on the SAT Chemistry test is the acid / base titration. When titrating an acid with a base, you will be adding a basic solution with a known concentration to an acid. This will raise the pH until the **equivalence point** is reached. The equivalence point is the point at which the moles of base added is equal to the moles of acid in the unknown solution. This results in complete neutralization of both acid and base, generally producing a neutral pH. For more information about neutralization see problem 50. For titration involving a weak acid or weak base where the final pH is not neutral, see problem 146.

3. Which of these pieces of laboratory equipment is commonly used to transfer exact volumes of liquid from one container to another?

    (A) Balance
    (B) Condenser
    (C) Pipettete
    (D) Thermometer
    (E) Funnel

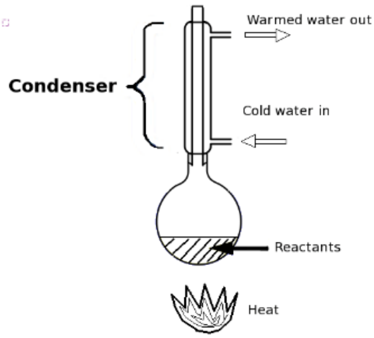
***Solution:** A **pipettete** is the exact device for this from this list, choice (C).

Remarks: Below is an illustration of a pipettete and a pipettete pump. This type of pipettete needs vacuum suction in order to bring the liquid up. A rubber bulb or pipettete pump, shown at left, can be used to create the vacuum for suction. The left end of the pipettete plugs into the pipettete pump and the right side is used for dispensing. The volumetric markings go to the bottom of the pipettete so you can always tell which end should be used to dispense the liquid. Though depicted horizontally, pipettetes are always held vertically to avoid contaminating the pipettete pump with the liquid or solution being measured and dispensed.

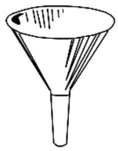


Definitions: A **balance** is for measuring mass. It is common practice to "**tare**" (or "zero out") the balance after placing the container that will hold the chemicals on it. **Taring** resets the balance reading to zero. This ensures you do not have to do any subtraction to determine the mass of the substance you are weighing.

A **condenser** is used to capture a gaseous chemical product using low temperatures to condense the gas into a liquid. In the picture below, cold water is used to cool the gas that is produced by the reaction.



A **funnel** is helpful to prevent spills when transferring chemicals poured from beakers or flasks and is generally not needed when using a pipette.



4. Which of these pieces of laboratory equipment is commonly used when separating a mixture containing a solid and a liquid?

   (A) Balance
   (B) Condenser
   (C) Pipette
   (D) Thermometer
   (E) Filter funnel

***Solution:** A **filter funnel** is the piece of equipment from this list that would be used for separating the solid and liquid, choice (E).

Remarks: This question's wording to 'separate a solid and liquid' describes a **filtration**. A pipettete would not be used because it is specifically for dispensing liquid. **Precipitates** could clog the narrow tip of the pipette. Separation of states of matter is accomplished when liquids are poured from **beakers** or **flasks** into a funnel that contains a piece of **filter paper**. The filter paper catches the large solids and allows liquids to drain through.

Definitions: A **filtration** is a type of separation experiment that is used when solids are in a mixture with a liquid. Filtrations are generally performed to capture a **precipitate** from a reaction. **Precipitates** are solid compounds (or molecules) that were produced in a liquid environment. **Beakers** and **flasks** are containers used to contain and transfer liquids. They may be **volumetric**, and have markings to show specific volumes that may be transferred. When using **filter paper** to separate a solid from a liquid, the paper works in a similar manner to a coffee filter. The solid particles are too large to go through the **pores** or holes of the paper, and therefore the particles stay on the paper as the liquid is drained.

5. Which of the following is a product for the complete combustion of a hydrocarbon?

   (A) NO_2
   (B) C_2
   (C) O_2
   (D) CO_2
   (E) O_3

***Solution:** Hydrocarbons produce CO_2 and H_2O when completely combusted. Therefore, the correct answer is choice (D).

Remarks: (1) A **hydrocarbon** is made up of only hydrogen and carbon. The process of **combustion** will combine oxygen with each of the individual elements. Complete combustion will produce only H_2O and CO_2.

(2) A hydrocarbon will not contain nitrogen, so NO_2 is not possible. C_2 has not been combined with oxygen so it has not undergone combustion, while O_2 and O_3 have not been combined with either hydrogen or carbon.

Definitions: Combustion is a type of chemical reaction which combines oxygen gas with an element or compound, usually a **hydrocarbon**, to produce CO_2 and H_2O. Incomplete combustion can produce carbon monoxide, CO.

6. Which of the following observations indicates that a chemical reaction occurred after a colorless liquid and a blue liquid were mixed?

I. A change in color to red
II. Appearance of a precipitate
III. The resulting solution evaporates

(A) I only
(B) II only
(C) II and III only
(D) I and II only
(E) I, II, and III

***Solution:** Both I and II can happen when there is a chemical change. Therefore, the answer is choice (D).

Remarks: (1) **Evaporation** is an example of a physical change. A liquid evaporating is not a chemical reaction. It is simply a change of **phase**.

(2) Chemical changes occur when **bonds** within a molecule are broken or new ones are created. New bonds create new compounds that may be seen with changes in color. The color change of the solution in this problem indicates that a new compound, which reflects red light, has been formed.

(3) A chemical change may also result in the production of new states of matter. The most common change in state of matter when mixing two solutions is for a gas (bubbles) or a solid (**precipitate**) to be produced.

(4) Another indicator of chemical change is the production of light. Light is a form of energy that electrons may release as they change within a molecule. An example would be when wood is being burned and we see the light of the fire.

Definitions: Phase refers to the state of matter that a chemical is in. The common states of matter are solid, liquid, and gas. A **precipitate** occurs when a solid is produced in a liquid solution. The solid will settles to the bottom of the solution.

Evaporation is the change from the liquid to gaseous state at a temperature below the boiling point. Evaporation does not involve the breaking of chemical bonds, but does involve the breaking of **intermolecular bonds**, which are the bonds between different molecules. For a further discussion of intermolecular bonds see problems 88 and 127.

Chemical **bonds** refer to the sharing of electrons for covalent bonds or the stealing and magnetic attraction of ionic bonds. For a further discussion of covalent and ionic bonds, see problems 26 and 94.

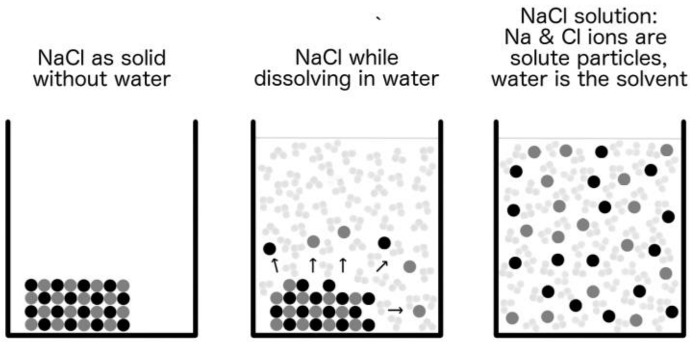
7. Which of the following observations would not be classified as a physical change?

   (A) 1 gram of $C_6H_{12}O_6$ being dissolved in water
   (B) 2 grams of $C_6H_{12}O_6$ being divided into two 1 gram piles
   (C) 1 gram of $C_6H_{12}O_6$ burning to produce CO_2 and H_2O
   (D) 2 grams of $C_6H_{12}O_6$ being warmed by 10ºC
   (E) 2 grams of ice being melted and brought to a boil

***Solution:** A question asking about a non-physical change can only mean a chemical change. The only answer here that deals with a chemical change is the burning of a substance, choice (C).

Remarks: (1) Burning is another way of indicating combustion took place, and combustion produces new compounds. In this case $C_6H_{12}O_6$, which is the formula for a single sugar molecule (in this case glucose), has its bonds broken when CO_2 and H_2O are produced during combustion.

(2) **Dissolving**, separating, and changing the temperature of a substance are examples of physical changes. An illustration of dissolving is shown below. **Melting** and **boiling** are both descriptions of matter as it changes state.



(3) Changes in states of matter, which all count as physical changes, are the following: **condensation, vaporization, sublimation, deposition, melting,** and **freezing.**

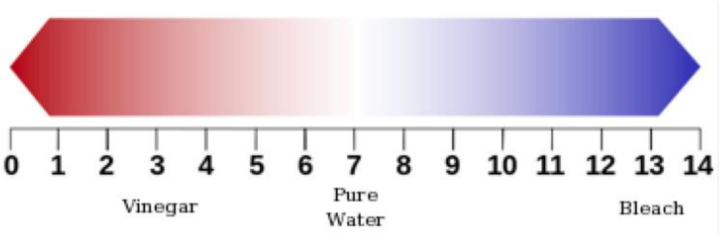
Definitions: Dissolving a substance means changing it from a solid state to an **aqueous** state. The aqueous state has individual molecules of a substance surrounded by the molecules of the liquid that is dissolving it. For a further discussion of dissolving, see problem 55. **Boiling** is the observable release of gas bubbles from a liquid that is spontaneously changing into a gas. You can think of what happens to water when you heat it up on the stove as a reference. This phase change when the liquid is at the boiling point is called **vaporization**. **Condensation** is the transition from the gas state to the liquid state. An example is when the hot steam from a shower condenses on a mirror. **Sublimation** is the change from solid to gas. The most common example is with **dry ice,** which is solid CO_2. Dry ice appears similar to normal ice, but it does not melt. Instead, dry ice seems to disappear because it goes directly from the solid to the gaseous state, so it never gets anything wet! **Melting** is the change from solid to liquid, and **freezing** is the change from liquid to solid. Again, thinking about these transitions with water in mind is a good way to remember them. **Deposition** is the change from a gas directly to a solid. In colder months, you might see a thin layer of ice on windows, but the windows were never wet! This is called frost. Frost on windows is the result of water vapor in the air freezing when coming in contact with the cold glass.

8. A solution with a pH of 8.0 is

(A) strongly acidic
(B) weakly acidic
(C) neutral
(D) weakly basic
(E) strongly basic

***Solution:** The pH scale dictates that this solution is weakly basic, choice (D).

Remarks: (1) The pH scale is a measure of the **acidity** or **basicity** of a solution. It is **neutral** at 7, basic above 7, and acidic below 7. The pH scale has a normal range from 0 to 14 with 0 representing the strongest acidic solution and 14 representing the strongest basic solution. A scale is shown here with household solutions for reference.



(2) The pH scale is a logarithmic scale (like the Richter scale for earthquakes). This means it is based on powers of 10 for each step. For example, a solution of pH 5 is 10 times more acidic than a solution of pH 6. A solution of pH 4 is 100 times more acidic than a solution of pH 6.

(3) Even a neutral solution of pH 7 has some acid and basic qualities. The reason it is called neutral is because it is equally acidic and basic. Comparing a pH 7 solution to a pH 8, we would say the pH 7 solution is 10 times more acidic than the pH 8.

(4) Just as we can describe solutions by their acidity, we can also describe them by their basicity. We could say that a pH 8 solution is 10 times more basic than a pH 7 solution.

Definitions: The **acidity** of a solution is a measure of its hydrogen ion (H^+) concentration. Sometimes H_3O^+, called **hydronium**, is used to represent the H^+ concentration since this is more representative of how H^+ ions interact with water.

The **basicity** of a solution is a measure of its hydroxide ion (OH^-) concentration. There is no equivalent for hydroxide like hydronium is for H^+. At acidic pH's, the H^+ concentration is greater than the OH^- concentration. At basic pH's, the OH^- concentration is greater than the H^+ concentration. At the neutral pH of 7, the concentrations of H^+ and OH^- are equal.

LEVEL 1: PRODUCT AND REACTANT RELATIONSHIPS

9. The number of oxygen atoms in 1.5 mole of $NaHCO_3$ is

(A) 4.5
(B) 2.7×10^{23}
(C) 6.0×10^{23}
(D) 27×10^{24}
(E) 2.7×10^{24}

***Solution by mole conversion:** The first step here is to recall that the value of a mole is 6.02×10^{23}. Each molecule of $NaHCO_3$ has three oxygen atoms, so that 1.5 moles of this molecule will have:

$$1.5 \text{ moles} \times \frac{6.02\times10^{23}\text{molecules}}{1 \text{ mole}} \times \frac{3 \text{ O atoms}}{\text{molecule}} = 27.09 \times 10^{23}\text{oxygen atoms}$$

which is closest to 2.7×10^{24}, choice (E).

Remarks: (1) This problem relies primarily on your knowledge of the mathematical definition of a mole, secondly on your understanding of the definition of atom vs. molecule, and lastly on your familiarity with exponents.

(2) Some students strongly feel that this problem requires the use of a calculator. However, once you are comfortable with the definition of a mole, the commutative property of multiplication and rounding numbers, and exponents, you can rewrite the problem as follows:

$$1.5 \text{ moles of } \cancel{\text{molecules}} \times \frac{3 \text{ O atoms}}{\cancel{\text{molecule}}} = 4.5 \text{ moles of O atoms}$$

$$4.5 \ \cancel{\text{moles}} \text{ of O atoms} \times \frac{6 \times 10^{23}}{\cancel{\text{mole}}} = 27 \times 10^{23} \text{ O atoms}$$

This boils down to computing $1.5 \times 3 \times 6$ to get 27, and then 'attaching' 10^{23} to the end. 27×10^{23} is not proper scientific notation since it starts with a number with two digits, so we divide the first number by 10 to make it 2.7 and balance this out by multiplying the second number by 10, making it 10^{24}.

(3) Within a chemical formula, subscripts that appear after an element indicate the number of atoms of that element. The only exception is that when there is just one atom of that element, no number appears.

For example; P_3 is three phosphorous atoms as a molecule, F_2 is two fluorine atoms as a molecule, and Ar is a single argon atom. Each molecule of $NaHCO_3$ contains one atom of sodium (Na), one atom of hydrogen (H), one atom of carbon (C), and three atoms of oxygen (O).

(4) A mole is a number of something, just like a dozen is. However, while a dozen is simply 12, a mole is 6.02×10^{23}, a much larger number. When approaching a problem using moles, if you find yourself getting confused, simply replace 'moles' with 'dozens' and then substitute backwards. An example follows in remark (5).

1.0 mole of $NaHCO_3$ would contain 3 moles of oxygen, and 1.5 moles will have 4.5 moles of oxygen. This problem is asking for atoms of oxygen. However, after this step we have only moles of oxygen atoms, so the conversion between moles and atoms needs to be applied. This is similar to saying we have 1.5 dozen atoms and we need to figure out the exact number ($12 \times 1.5 = 18$ atoms).

(5) If you find moles confusing and are not 100% sure when to divide and when to multiply for problems involving moles, use the trick demonstrated on the following practice problem.

1.5 moles of $NaHCO_3$ would be how many molecules?

a) Change moles to dozen.

1.5 **dozen** of $NaHCO_3$, would be how many molecules?

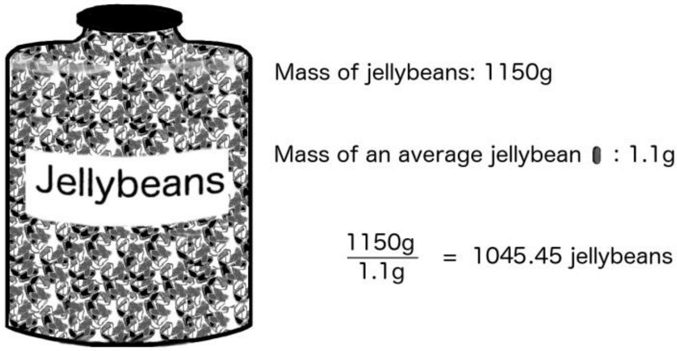
This problem is much easier. You would naturally perform the following calculation:

$$1.5 \text{ dozen } \times \frac{12 \text{ molecules}}{1 \text{ dozen}} = 18 \text{ molecules}$$

b) Now, switch out the terms and corresponding numbers with their mole counterparts to get the correct answer.

$$1.5 \textbf{ moles } \times \frac{\mathbf{6.02 \times 10^{23}} \textbf{ molecules}}{1 \textbf{ mole}} = 9.03 \times 10^{23} \text{ molecules}$$

(6) The original problem also involves the idea of **significant figures**. The answer (2.709×10^{24}) has a decimal component that ends up limited to only 2.7×10^{24}. This is because the number "1.5" limits the specificity that we can determine the number of atoms to. The core idea behind significant figures is that you want the number you state to be 'meaningfully accurate.' For example, let's pretend you win a competition where you had to guess the number of jellybeans in 1150g of jellybeans by being the closest to the actual number. You did this by looking up the average mass of a jellybean (1.1g/jellybean) and estimating that there were 1045 jellybeans because you rounded down. You don't know the actual number, but the judge determined your guess was the closest.



When claiming the prize, which is all 1150 grams of jellybeans, you find an extra jellybean next to the jar. Then your chemistry teacher asks you how many jellybeans you won, you should not say "1045 plus 1, so 1046!" Your guess was just the closest one to the actual number, it could've been 1044 or 1047 or even less accurate! The judge never told you the exact number. So you would be wrong to report the exact number of jellybeans with the new one added since you cannot be completely sure. This seems trivial because it is an example with jellybeans, but it is very important to know exactly how much of something you have when it is a strong medicine or potentially dangerous pesticide. Not over-reporting your number is what is meant by meaningful accuracy. In general, unless a question specifically asks for significant figures, you can select the answer choice closest to your own without worrying about the details of the calculation.

Definitions: A **mole** is 6.02 x10^{23} 'of something', exactly like a dozen is 12 'of something.' Moles can be converted in several different ways. See problem 43 for a figure describing the different conversions to and from moles.

Significant figures are the number of digits that you should use to report a scientific result. When using multiplication or division in a calculation, you use the least number of significant figures found in one of the factors of the mathematical setup.

Significant figures explanation

Example: $3.5 \times 98.002 = \mathbf{340}$ (Calculator answer: 343.007)

Determining the number of significant figures to use is simple when you know the rules, though sometimes the calculations you will produce are not intuitive (such as the one above).

Rule 1: Every non-zero number is significant

a) 3.5 has two significant figures.

Rule 2: *Some* zeros are significant.

Zeros to the left of the decimal are only significant if there is a decimal. 56**0** has *two* significant figures. 56**0.** has *three* significant figures.

b) If there is no decimal, zeros to the left of the ones place are only significant if they are "sandwiched" between other numbers. 56**00** has *two* significant figures. 5600. has *four* significant figures. 56**0**1 also has *four* significant figures.

c) Zeroes that are to the *right* of the decimal are **always** significant 98.**00** has four significant figures.

3.5 has two significant figures, 98.002 has 5 significant figures. When multiplying or dividing you always report the answer to the least number of significant figures appearing in one of the factors, which in this case is two (because "3.5" has two significant digits). The exact product here is 343.007, but that number has 6 significant digits. If we trim down to 343, we still have a number with 3 significant figures, and so we must use one of the 'zero rules' (rule 2a) to get the correct number. We round the 3 down (as per the usual rounding rules) and **340** is the correct answer.

Remark: It may seem counter intuitive to not use all of the numbers a calculator would show you, but this method preserves a reasonable level of accuracy. If you won the lottery today and you knew it was *roughly* 10 million dollars, you might excitedly post to your social media about the windfall. However, if on your way to pick up your prize you found a dime on the sidewalk, would you update your daily total to $10,000,000.10? You certainly could, but since you do not know the exact amount of your winnings, the additional ten cents should not affect your estimate of your earnings. This is similar to the previous jellybean example.

A **molecule** is an arrangement of atoms that are bound together as a single unit. Just like a car always comes with 4 wheels and an engine, every molecule of a particular substance will always have the same number of atoms within it.

10. Select the true statement about catalysts from the choices below.

    (A) They decrease the value of K_{eq}.
    (B) They increase the concentration of products at equilibrium.
    (C) They shift the position of equilibrium to the right.
    (D) They decrease the activation energy of the reaction.
    (E) They are consumed as the reaction proceeds and more catalyst constantly needs to be added to finish the reaction.

***Quick solution:** Catalysts have three major characteristics, one of which is that they decrease the activation energy of a reaction, choice (D).

Remarks: (1) Catalysts have three major aspects: (i) they increase the reaction rate (both forward and reverse), which they do by (ii) lowering the activation energy for the reaction, and (iii) they are not produced or consumed by the reaction (whatever catalyst molecule you started with, you also find with the same formula at the end).

(2) Each of the wrong answers A, B, and C are observed when dealing with **equilibrium** and **Le Chatelier's principle**, but not for catalysts.

(3) Changes in the value of the equilibrium constant K_{eq} (mentioned in answer choice A) are possible only by changing the temperature.

Increasing the concentration of products at equilibrium (answer B) is accomplished by applying Le Chatelier's principle (see problem 78), generally by changing a reaction condition that would favor the formation of products. This would require knowledge of a specific chemical equation, which is not provided here.

Shifting equilibrium (answer choice C) generally refers to an increase in products (shift right) or an increase in reactants (shift left) by applying Le Chatelier's principle.

Definitions: A **catalyst** is a substance that increases the rate of a chemical reaction without itself undergoing any permanent chemical change.

The **equilibrium constant (K_{eq})** of a reaction is a number that describes the ratio of products to reactants at equilibrium.

Equilibrium is the point in a reaction when the product and reactant concentrations are held constant due to equal rates of the forward and reverse of the reaction. You can think of a crowded cafe around lunch time. At 12'o clock it is the busiest and new customers can sit down to eat only when other people leave. Before 'reaching equilibrium' or 12 o'clock, more people are heading into the café to sit down for lunch than are leaving. At 12 o'clock or 'equilibrium' all the tables are full and the amount of people entering is equal to the amount leaving. As soon as one table is empty it is filled with new customers. After lunchtime, it would be 'past equilibrium' because more people are leaving to go back to work or home than are sitting down. Only when the number of customers in the restaurant is constant do we consider the restaurant to be 'at equilibrium.' Similarly, only when the chemical concentrations are constant do we consider a reaction to be at equilibrium.

Le Chatelier's principle can be summarized as 'a reaction at equilibrium will always try to offset any change made to it to return to the original equilibrium state.'

11. The least amount of energy that is required for molecules to undergo a reaction and form products is called

    (A) Ionization energy
    (B) Enthalpy of formation
    (C) Activation energy
    (D) Reduction potential
    (E) Gibbs free energy

***Solution:** This question is directly referring to activation energy, choice (C).

Remarks: Chemical reactions always require an input of energy in order to cause the reactants to change and form products. This energy is called the **activation energy**, and is used to put the molecules into the **transition state**.

Note: The other answers are all concepts that follow periodic trends or refer to thermochemistry. See problem 35 for further information about ionization energy, problem 74 for more on enthalpy, problem 142 for reduction potential, and problem 106 for Gibbs free energy.

Definitions: The **transition state** is a structure that the reactants take on before they become products. The transition state resembles a temporary or in-between form for the reactants and is unstable. This unstable state results in the creation of the products. **Activation energy** is the energy that is used to force the reactants to adopt the transition state on their way to becoming products.

$$\_Na_{(s)} + \_H_2O_{(l)} \rightarrow \_NaOH_{(aq)} + \_H_{2(g)}$$

12. When 2 moles of sodium react with an excess of $H_2O_{(l)}$, how much $NaOH_{(aq)}$ is produced? (Equation is <u>not balanced)</u>

(A) 0.5 moles
(B) 1 mole
(C) 2 moles
(D) 4 moles
(E) 6 moles

***Solution:** The **coefficients** of the balanced equation are 2, 2, 2, 1. Two moles of sodium will produce two moles of $NaOH_{(aq)}$. Therefore, the answer is choice (C).

Remarks: (1) The law of conservation of mass says that matter can be neither created nor destroyed. Therefore, all of the atoms on the reactants' (left) side must be present on the products' (right) side. This problem requires you to balance the numbers of atoms for each element on both sides.

(2) Subscripts with letters indicate the state of matter of a compound:

$_{(s)}$ = solid, $_{(aq)}$ = aqueous or dissolved, $_{(l)}$ = liquid, and $_{(g)}$ = gas

(3) The best place to start a problem like this is by making a small table that will help you to keep track of the elements. This first table sets up the initial conditions with all coefficients being 1.

$$1\ Na_{(s)} + 1\ H_2O_{(l)} \rightarrow 1\ NaOH_{(aq)} + 1\ H_{2(g)}$$

	Reactants	Products
Na	1	1
H	2	3
O	1	1

The first thing we see is that the equation is not balanced for hydrogen. We cannot have any less hydrogen atoms on the right hand side, so we must increase the amount of hydrogen on the reactants' side. The left hand side can only contribute hydrogen through H_2O, so it is only able to add two at a time. If we raise the number of water molecules on the left by two, we will have four hydrogen atoms. In order to balance this on the right hand side, it is only possible if we increase the number of NaOH to 2 as well (If we were to make $H_{2(g)}$ have a coefficient of 2, this would bring us to 5 hydrogen atoms on the right and this would not be correct). Based on both these changes, the oxygen atoms and sodium atoms will change as well. The chart and the updated equation should now look like this:

$$Na_{(s)} + \overset{2}{\cancel{1}}\,H_2O_{(l)} \rightarrow \overset{2}{\cancel{1}}\,NaOH_{(aq)} + H_{2(g)}$$

	Reactants	Products
Na	1	~~1~~ 2
H	~~2~~ 4	~~3~~ 4
O	~~1~~ 2	~~1~~ 2

At this point, we only need to deal with sodium (Na). Since sodium is by itself in the chemical equation, we can change its coefficient without upsetting any of the other numbers which brings us to the balanced equation:

$$2\ Na_{(s)} + 2\ H_2O_{(l)} \rightarrow 2\ NaOH_{(aq)} + H_{2(g)}$$

Note: (1) Many students find balancing equations tricky because it initially appears that the least common multiple will help to solve the problem. However, the 3 hydrogen atoms on the right in this problem are actually split between two different molecules. This means that using the least common multiple will probably not work.

(2) Whenever a coefficient or charge in a chemical equation is 1, you do not need to write the 1. This was done after the first step with hydrogen in this problem. This provides a convenient, simpler notation without losing any information.

Definitions: The **coefficients** in a chemical equation indicate how many of each molecule are either used or produced depending on whether the molecule is on the left (reactants) or right (products) side of the equation.

13. A 27.0 g sample of a hydrated salt is heated until all of the water is driven off. The final mass of the solid is 18.0 g. What is the mass percent of water in the original sample?

    (A) 150.0%
    (B) 66.7%
    (C) 70.0%
    (D) 50.0%
    (E) 33.3%

***Solution:** Water made up 27.0 – 18.0 = 9.0 grams of the original sample. Therefore, the mass percent of water in the original sample was $\frac{9}{27}$ or 33.3%, choice (E).

Remarks: (1) Certain ionic compounds include water molecules as part of their structure without creating covalent or ionic bonds with these molecules. These compounds are called **hydrates**.

(2) In this problem "driven off" refers to the process of evaporation. Hydrates can have the water molecules of their structure removed by evaporation due to heating. When a hydrate is heated, its mass will always decrease due to this evaporation. This is a reversible process. If left alone, a dried hydrate will reabsorb water molecules as it cools down until its original structure, including water molecules, is restored.

(3) Hydrates can have different numbers of water molecules associated with their structure. The number of water molecules is always indicated by the prefix of the word 'hydrate.' Below are three examples of hydrates with their associated names.

(4) Notice that each of the ionic compounds below follow normal ionic naming rules with the 'hydrate' term at the end. Ionic compounds are normally named with the cation first and the anion ending in the suffix '-ide'. For examples with transition metals see problem 102.

Copper (II) sulfate pentahydrate: $CuSO_4 \cdot 5\ H_2O$

Nickel (II) chloride hexahydrate: $NiCl_2 \cdot 6\ H_2O$

Cobalt (II) chloride hexahydrate: $CoCl_2 \cdot 6\ H_2O$

Definitions: A **hydrate** is a solid compound which has water molecules incorporated into its crystal structure. Hydrates are special because the water molecules are not held in place permanently and can be removed by evaporation due to heating.

$$C_7H_{16}O_{4(l)} + 9\ O_{2(g)} \rightarrow 7\ CO_{2(g)} + 8\ H_2O_{(l)}$$

14. If 4.5 moles of $O_{2(g)}$ reacts completely according to the balanced equation above, which of the following must be true?

    (A) 4 moles of $CO_{2(g)}$ are formed
    (B) 3 moles of $H_2O_{(l)}$ are used
    (C) 0.5 moles of $C_7H_{16}O_{4(l)}$ are formed
    (D) 1 mole of $C_7H_{16}O_{4(l)}$ are used
    (E) 4 moles of $H_2O_{(l)}$ are formed

***Solution:** The **stoichiometry** of the reaction indicates that choice (E) must be the amount of product that is formed.

Remarks: (1) This question requires an understanding of how to use the coefficients of a balanced chemical equation to predict amounts of reactants consumed or products formed. The molecules on the left side of the equation are called reactants and those molecules are consumed. The molecules on the right side of the equation are called products and those molecules are produced. This level of understanding eliminates choices (B) and (C) which each suggest that products are consumed and reactants are formed.

(2) By using 4.5 moles of $O_{2(g)}$, this problem is indicating that you should compare 4.5 to the coefficient of $O_{2(g)}$, which is 9. This comparison shows that the reaction occurred at a ratio corresponding to half of the coefficient. Neither (A) nor (D) are one half of the given coefficients, leaving choice (E) as the answer.

Observe that half of the normal moles of water $(\frac{1}{2} \cdot 8 = 4)$ are produced.

Definitions: Stoichiometry is the word describing the relationship between the coefficients of products and reactants of a reaction. It is generally used to refer to the whole number coefficients of a balanced chemical equation. While chemical equations are normally discussed in terms of the simplest number of moles, the stoichiometry of a reaction holds true for individual molecules or multiples of molecules.

For example, if 4 legs + 1 seat → 1 chair, then 400 legs + 100 seats → 100 chairs. Similarly, if CO_2 + H_2O → H_2CO_3, then 1 molecule of CO_2 + 1 molecule of H_2O will produce 1 molecule of H_2CO_3 and 1 mole of CO_2 + 1 mole of H_2O will produce 1 mole of H_2CO_3.

15. Which of the following has a molar mass greater than that of CO_2?

I. H_3O^+
II. HCl
III. Kr

(A) I only
(B) II only
(C) III only
(D) II and III only
(E) I, II, and III

***Solution:** By tallying the atomic weight of the atoms in each choice, only Kr has a larger molar mass. Therefore, choice (C) is the correct answer.

Remarks: (1) The **molar mass** of a compound is the combined mass of all of the atoms within the formula of the compound. The reason it is called "molar" is because it is the mass of a large number of atoms. Molar masses are given in grams. For more on moles see problem 9.

The **atomic weight** is the number on the periodic table below each element symbol. Atomic weight, as its name suggests, refers to the weight of the average atom. The number does not have any units shown and it is up to you to understand what units to use. When talking about an individual molecule's mass, the mass you tally will be called the **molecular mass**. Individual atoms and molecules are very small and are measured in atomic mass units or amu.

Generally, you will work with the molar mass, which is the weight of many, many molecules. The molar mass is simply the molecular mass in grams.

The reference point in this question is CO2, which has one carbon atom (C) and two oxygen atoms (O). The atomic weight for each element is found in the periodic table, and to find the molecular weight we add the atomic weights together. One mole of carbon atoms has a mass of 12.011 g, and two moles of oxygen atoms have a mass of 32.00 g

(16.00 g × 2). The total mass, and therefore molar mass, is 12.011 + 32 = 44.011 grams.

(2) The goal now is to find a compound that exceeds 44.011 grams in its mass. The fastest way to do this will be to take ballpark estimates of each atomic mass. In the event that one of the choices comes very close to 44.011 g, we can then compute exact values.

(3) The mass of H_3O^+ is composed of three hydrogen atoms (H) and one oxygen atom. Each hydrogen atom contributes a mass of approximately 1 g, and each oxygen atom contributes a mass of 16 g. The calculation becomes 3 × 1 g + 16 g = 19 grams, which is much less than 44.011 g and can be discarded. HCl has a molar mass consisting of 1 hydrogen (1 g) and 1 Chlorine (Cl, 35 g). HCl's molar mass is therefore 36 grams which is less than 44.011 grams. Kr consists of one atom of krypton which has a molar mass of 84 g. This is greater than 44.011g.

Definitions: The **atomic weight** of an atom is the larger number shown for each element on the periodic table. It is the average of all isotopes of the atom. See problem 25 for more on isotopes. Atomic weight is measured in **atomic mass units**, or **amu**. Atomic mass units are mainly derived from the numbers of protons and neutrons of an atom. Protons and neutrons both have a mass of 1 amu, while electrons have a mass of 1/1,836 amu. Because the mass contribution of electrons is so small, it is generally approximated as zero.

Working with individual atoms or molecules is the exception to the rule in chemistry as they are very small and difficult to work with. Therefore, atomic mass units are rarely asked for.

A much more useful number is the **molar mass** of a molecule or compound. The molar mass is the number of grams found in one mole of a compound or molecule. One of the 'hidden beauties' of the periodic table is that the number of amu for an atom of any element is also the number of grams in a mole. So 12.011 amu is the atomic weight of carbon and 12.011 grams is the molar mass of carbon. This allows an easy calculation of the **molecular mass** of a compound as they are the same number, but the molar mass has the units amu.

16. A certain compound has a carbon to oxygen ratio of 6 to 3. What is the empirical formula of this compound?

    (A) $6C + 3O$
    (B) CO_3
    (C) C_2O_3
    (D) C_2O
    (E) $2C + O$

***Solution:** The empirical formula for this compound would be two carbon atoms to one oxygen atom, choice (D).

Remark: The **empirical formula** of a compound is found by reducing the ratio of the elements to their smallest multiples form. In this problem, 6 to 3 reduces to become 2 to 1. When writing a chemical formula, the number of atoms always becomes a subscript.

Definitions: An **empirical formula** is the simplest whole number ratio of atoms for each element in a compound. The empirical formula may or may not also be the **molecular formula**. The molecular formula of a compound is the actual number of each atom. For example, glucose has a molecular formula of $C_6H_{12}O_6$ and an empirical formula of CH_2O.

Note: Ionic compounds always have a molecular formula that is also an empirical formula.

LEVEL 1: STATES AND REACTIONS OF MATTER

17. The statement that 'particles in a substance move slower and slower until some are able to lock into fixed positions from their previously free moving state,' is best described by which of the following?

    (A) An equilibrium state between gas and liquid phases.
    (B) A general method for determining temperature.
    (C) Relationship between pressure and volume at constant temperature.
    (D) Relationship between specific heat and temperature change.
    (E) Liquid phase changing into the solid phase.

***Solution:** The liquid state of matter is defined by a fluid movement of molecules or atoms past each other, while a solid is defined by the minimal movement of molecules. The action being described here is best associated with the transition from liquid to solid, choice (E).

Remarks: (1) **Temperature** correlates with a molecule's speed. The substance here is slowing down, and therefore the temperature is decreasing. This suggests either the action of **condensation** or **freezing.**

(2) The term *fixed* in the question indicates a solid state of matter, while *free moving* could indicate liquid or gas. The only answer choice that involves a solid versus a liquid or gas is choice (E).

Definitions: Temperature is defined as the average kinetic energy of a substance, and is related to the mass and speed of the molecules. See problem 18 for a quantitative look at temperature.

Condensation is the changing of a substance's state from a gas to a liquid, while **freezing** is the change of state from a liquid to a solid. Both involve transitions to slower moving phases of matter and generally occur as a substance's temperature is lowered.

18.

I		II
The average kinetic energy of molecules in a liquid decreases as the temperature of the liquid decreases	BECAUSE	the average speed of the molecules in the liquid decreases as the temperature decreases.

***Solution:** Temperature is the term for the average kinetic energy, and therefore statement I is **True**. Kinetic energy is determined by the formula $\frac{1}{2}\text{ mass} * (\text{velocity})^2$, and therefore statement II is **True**. Statement II is a correct explanation of statement I, so **CE** is also the case.

Remarks: (1) Because temperature measures the average **kinetic energy** of a substance, we need to look further at kinetic energy (KE). KE is represented by how much the molecules are moving or vibrating, which is a way of describing speed. There is no mention of a gain or loss of mass in this question. Therefore, the only way for the liquid to lower its temperature is by slowing down the velocity, or speed, of its particles.

(2) Advanced questions about temperature may ask about the speed of different molecules when they are at the same temperature. In this case, the lightest molecule will always be the fastest one. This is because mass is inversely proportional to velocity when kinetic energy is held constant.

Definitions: Kinetic energy is the energy of 'movement' represented by the equation KE = $\frac{1}{2}$ mass $*$ (velocity)^2, or KE = $\frac{1}{2}mv^2$.

19. When a rigid container filled with a gas is cooled, which of the following is true?

    (A) The pressure of the gas inside is decreased.
    (B) The pressure of the gas inside is increased.
    (C) The volume of the gas inside is increased.
    (D) The volume of the gas inside is decreased.
    (E) The average speed of the molecules inside is increased.

*** Solution:** Temperature varies directly with pressure. Therefore, as the container is cooled, the pressure of the gas is decreased, choice (A).

Remarks: (1) A rigid container will not change its **volume**. A property of gases is that they will always expand or contract to fill their container. Therefore, choices (C) and (D) are not correct. As discussed in the previous problem, a temperature decrease should result in the average speed of the molecules decreasing, ruling out choice (E).

(2) **Solution by plugging in:** If you are not confident in your gas property knowledge, you can refer back to an important property of gases, which is that they obey the **combined gas law**. The combined gas law is $\frac{(P_1)(V_1)}{T_1} = \frac{(P_2)(V_2)}{T_2}$. P_1 represents the **pressure** before the change and P_2 represents the new pressure, V_1 represents the volume before the change and V_2 represents the new volume, likewise T_1 represents the temperature before the change and T_2 represents the new temperature. When given non-specific changes such as in this question, you can substitute very simple values to figure out the answer.

For example, the volume does not change for the container, so let's say that V_1 and V_2 are both 1 L. The temperature goes down, so let's say the initial temperature T_1 is 2° and the final temperature T_2 is 1°. We can set the initial pressure P_1 to 1 atm, but we do not know what happens to the pressure at the end, so we leave it as the variable P_2.

$$\frac{(P_1)(V_1)}{T_1} = \frac{(P_2)(V_2)}{T_2} \rightarrow \frac{(1\text{ atm})(1\text{ L})}{2^{\text{o}}} = \frac{(P_2)(1\text{ L})}{1^{\text{o}}} \rightarrow 1\text{ atm} = 2(P_2)$$

$$\frac{1}{2}\text{ atm} = P_2$$

Since the initial pressure was 1 atm we can confirm that the pressure decreases in response to cooling.

Definitions: Volume is the amount of space that something occupies. The standard unit of volume is liters (L). **Pressure** is defined as the number of molecule collisions, or force, over an area and is measured in kiloPascals (kPa), atmospheres (atm), or millimeters of mercury (mm Hg), also known as torr. For more on these unit values, see problem 20 remark (3).

The **combined gas law** is a formula that describes the mathematical relationships between temperature, pressure, and volume for a set quantity of gas. When one of these is held constant, the relationships are simplified as follows: Pressure and volume are inversely proportional; as one increases the other must decrease. Both pressure and volume are directly proportional to temperature. When temperature increases, volume increases if pressure is constant. When temperature increases, pressure increases if volume is constant. Conversely, when either volume or pressure increase, as the other is held constant, temperature increases as well. All of that is a mouthful compared with $\frac{(P_1)(V_1)}{T_1} = \frac{(P_2)(V_2)}{T_2}$. Memorizing and then plugging simple numbers into this equation will reproduce each of the relationships just described.

Here is a quick lesson in **direct variation:**

The following are all equivalent ways of saying the same thing:

(i) y varies directly as x
(ii) y is directly proportional to x
(iii) $y = kx$ for some constant k
(iv) $\frac{y}{x}$ is constant

For example, in the equation $P = kT$, P varies directly as T. Here is an example of a partial table of values for this equation.

T	1	2	3	4
P	6	12	18	24

Note that we can tell that this table represents a direct relationship between T and P because $\frac{6}{1} = \frac{12}{2} = \frac{18}{3} = \frac{24}{4}$. Here the **constant** k is 6.

Here is a quick lesson in **inverse variation:**

The following are all equivalent ways of saying the same thing:

(i) y varies inversely as x
(ii) y is inversely proportional to x
(iii) $y = \frac{k}{x}$ for some constant k
(iv) xy is constant

The following is a consequence of (i), (ii) (iii) or (iv).

(v) The graph of $y = f(x)$ is a hyperbola.

For example, in the equation $PV = k$, P varies inversely as V. Here is an example of a partial table of values for this equation.

P	1	2	3	4
V	22.4	11.2	7.33	5.6

Note that we can tell that this table represents an inverse relationship between P and V because.

$$22.4 = (1)(22.4) = (2)(11.2) = (3)(7.33) = (4)(5.6).$$

Here the **constant** k is 22.4.

20. Mount Washington is the highest point in the northeastern US at a height of 1,917 meters. The average pressure on top of Mount Washington is 602 mm Hg. According to the table below, at what temperature would methanol boil at this altitude?

Temperature (°C)	Pressure (mm Hg)
-44.0	1
-16.2	10
21.2	100
49.9	400
64.7	760

(A) -44.0° C
(B) between -44.0° C and 21.2° C
(C) between 21.2° C and 49.9° C
(D) between 49.9° C and 64.7° C
(E) 64.7° C

***Solution:** Methanol will boil when its vapor pressure is 602 mm Hg which occurs between 49.9° C and 64.7° C, choice (D).

Remarks: (1) When a liquid's **vapor pressure** is equal to the atmospheric pressure, the gaseous molecules are able to escape as bubbles. The spontaneous production of bubbles that escape a liquid is what we call **boiling**. Since vapor pressure depends on temperature, liquids will boil when they are heated to or above specific temperatures, but not when they are below those temperatures.

(2) The **boiling point** of a liquid is when it's vapor pressure is equal to atmospheric pressure. '760 mm Hg' should look familiar. This is the value of atmospheric pressure at sea level and is a value you should know. Since this problem asks about a height above sea level, you should know that atmospheric pressure decreases as height increases and that 760 mm Hg is not the target pressure in this problem.

(3) Since pressure is commonly measured in three different units, it is helpful to know the atmospheric pressure for the other units at sea level. At sea level the pressure is: 760 torr, 1 atm (atmospheres), or 101.3 kPa (kilopascals).

Definitions: All liquids evaporate to a greater or lesser degree. **Vapor pressure** is the amount of pressure that the evaporated gas produces. When the pressure of this evaporated gas is equal to the atmospheric pressure, we observe the phenomenon called boiling. The temperature when boiling first occurs is the **boiling point** of that liquid. An important boiling point to know is that of water, which is 100 $^{\circ}$ C.

21. A gas is kept in a rigid container. Which of the empirical laws does the following data support? (k is a constant)

Pressure (mm Hg)	300	400	700	1000
Temperature (ºC)	250	333	583	833

(A) $P = kT$ at constant V
(B) $T = \frac{V}{k}$ at constant P
(C) $P_{tot} = P1 + P2 + P3 \ldots$
(D) $P = k^{*}$ moles
(E) $P = \frac{k}{V}$ at constant T

***Solution:** This data supports a direct relationship between P and T, choice (A).

Remarks: (1) The inclusion of a rigid container lets us know that volume is not changing, so we can say that volume is constant.

(2) In this data we see that as the pressure increases, so does the temperature. The only answer that relates pressure and temperature changing while showing a constant volume is (A).

(3) The other answers point to **Boyle's, Charles', Gay-Lussac's,** and **Avogadro's laws** for gases, as well as **Dalton's law of partial pressures.**

(4) Be careful not to confuse the empirical gas laws with the term 'empirical formula' from problem 16. **Empirical** simply refers to the ability to observe something with an experiment as opposed to theory.

(5) See the discussion following problem 19 for more information on direct relationships (direct variation).

Definitions: Boyle's law shows the inverse relationship between only pressure and volume ($P \propto \frac{1}{V}$), as in choice (E)**.** The symbol ∝ means that each side is proportional to the other side, or that each side will change as shown to maintain a constant relationship. In this case the relationship is inverse since one of the terms (V) is in a denominator.

Charles' law shows the direct relationship between only volume and temperature ($V \propto T$), as in choice (B).

Gay-Lussac's law shows the direct relationship between only pressure and temperature ($P \propto T$), as in choice (A).

Avogadro's law shows that the volume of a gas is directly related to the moles ($V \propto n$), as in choice (D).

Dalton's law of partial pressures states that the total pressure of a mixture of gas is equal to the partial pressure of each gas that makes up the mixture, as in choice (C). This means that in a container having 10 atm of pressure of CO_2 and 5 atm of pressure of NO_2, the total pressure will be 15 atm.

22. For one mole of the following gases, which is the least dense at STP?

    (A) NO_2
    (B) O_2
    (C) CH_4
    (D) H_2
    (E) He

***Solution:** The gas with the least mass will be the least dense, answer choice (D).

Remarks: (1) One of the unique qualities of gases, as opposed to solids or liquids, is that the volume of a gas is nearly constant for the same number of particles in each gas (this is a consequence of Avogadro's Law). One mole of a gas behaving as an **ideal gas** will occupy 22.4 L at STP, regardless of the gas' molecular formula.

(2) Since gases can be considered to have the same volume, this makes finding their relative densities easy. $\text{Density} = \frac{\text{mass}}{\text{volume}}$, and it is given that all the gases have the same volume, determining the least dense gas is reduced to just finding the gas with the least mass. We find the mass of each gas by finding the molar mass for each molecule shown. For a review of determining molar mass see problem 15. H_2 has a mass of 2 grams, which makes it the least dense. For completeness, the other gases' masses are 46 g (NO_2), 32 g (O_2), 16 g (CH_4), 4 g (He).

(3) STP refers to standard temperature and pressure, which is 1 atm and 273° K. Do not confuse STP with *standard state,* which is used for thermochemistry calculations and is 298° K, 1 atm.

Definitions: An **ideal gas** is a gas that demonstrates two 'ideal behaviors.' The first behavior is that its particles will occupy the same volume as any other gas, which is 22.4 L per mole at STP. The second behavior is that the gas will exert the same pressure at the same temperature as any other gas. Observing these ideal behaviors always require *high temperature* and *low pressure* (we say that a gas behaves most ideally at high temperatures and low pressures). STP is an example of high *T* and low *P* conditions.

23. At 100°C, what is the vapor pressure of water?

(A) 1 mm Hg
(B) 100 mm Hg
(C) 76 mm Hg
(D) 760 mm Hg
(E) 1000 mm Hg

*** Solution:** The Celsius scale is based on water boiling at 100° C. Therefore, choice (D) is the correct answer.

Remarks: (1) Boiling happens when the vapor pressure of a liquid is equal to the atmospheric pressure (see problem 20 for more information). Since water's **normal boiling point** is 100° C, its vapor pressure must be equal to the atmospheric pressure at this temperature, which is 760 mm Hg.

(2) Water freezes at 0° C and boils at 100° C.

(3) The Celsius scale was calibrated for 100 divisions between water's freezing and boiling point. This is why you might also hear the term centigrade used to refer to these degrees. Centi- is the prefix meaning $1/100^{th}$.

Definition: The **normal boiling point** of a liquid is the temperature at which its vapor pressure is equal to 1 atm.

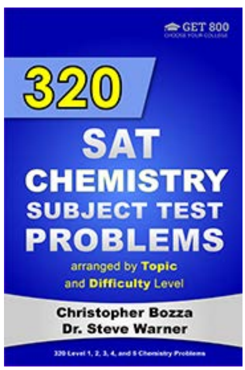
24. Liquids can be described as having:

    (A) indefinite shape and definite volume
    (B) indefinite shape and indefinite volume
    (C) definite shape and definite volume
    (D) definite shape and indefinite volume
    (E) indefinite shape and an indefinite density

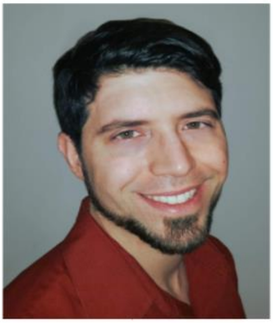
***Solution:** Liquids are defined as having an indefinite shape and a definite volume, choice (A).

Remarks: (1) Liquids can be poured into differently shaped containers showing that they have an indefinite shape. You could have a square cup and a round cup of the same volume and a liquid poured from one to the other would fit into each cup. This is as opposed to trying to fit a square peg into a round hole.

(2) A liquid will occupy the same volume in any container, and therefore liquids have definite volumes. Solids are defined as having a definite shape and definite volume. Gases are defined as having an indefinite shape and an indefinite volume. The reason that a gas has an indefinite volume is that a gas will expand to fill whatever container it is put in. It follows that the volume of a gas can change.



About the Author



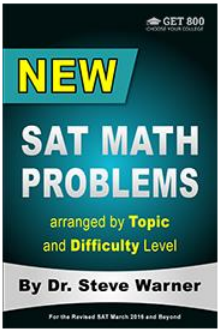
Christopher Bozza has spent more than five years as a tutor finding and inventing tips, tricks, and strategies to help his students rise up and over an average result and get ahead of their peers. He hones his tutoring practices in the highly competitive academic environment in and around NYC, where some of the most prestigious and challenging high schools and colleges can be found. He focuses on the subjects of chemistry, biology, physics, organic chemistry, biochemistry, molecular biology, microbiology, and genetics. Prior to helping students conquer science, he earned his bachelor's degree from Rutgers University in biotechnology, specializing in bioinformatics, and his master's degree from Cornell University in Plant Biology where he studied the cytology and molecular biology of chromosome pairing in maize.

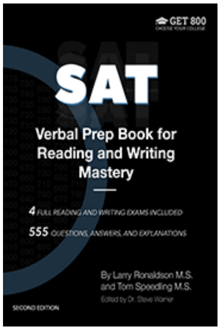
About the Editor

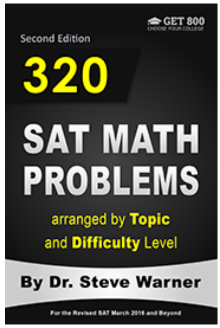
Dr. Steve Warner earned his Ph.D. at Rutgers University in Mathematics, and he currently works as an Associate Professor at Hofstra University. Dr. Warner has more than 20 years of experience in general math tutoring and more than 15 years of experience in SAT math tutoring. He has tutored students both individually and in group settings and has published several math prep books for the SAT, ACT and AP Calculus exams.

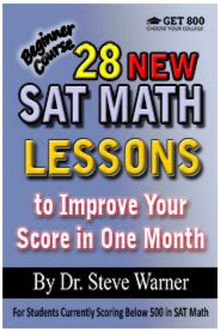
Dr. Steve Warner can be reached at steve@SATPrepGet800.com.

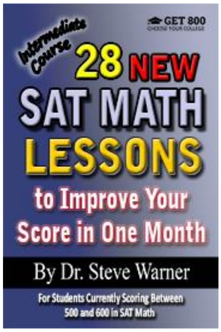
BOOKS BY DR. STEVE WARNER

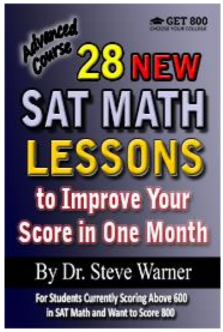


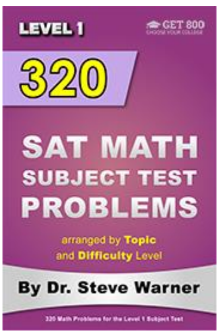


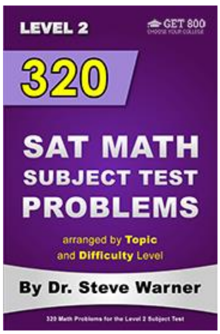


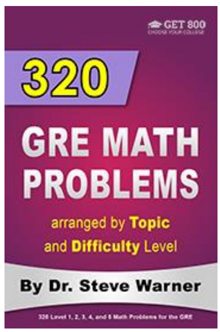












GET 800
CHOOSE YOUR COLLEGE

Put your first-choice college well within your reach with 320 SAT Chemistry Subject Test Problems Arranged by Topic and Difficulty Level. The problems in this book were carefully chosen by a scientist with more than 5 years of SAT chemistry tutoring experience. This book is laid out in such a way that any student can immediately find the problems he or she needs to improve in a quick and efficient manner.

Using this book you will learn to solve SAT chemistry subject test problems in clever and efficient ways that will have you spending less time on each problem, and answering difficult questions with ease. You will feel confident that you are applying a trusted system to one of the most important tests you will ever take.

The main part of the book consists of chemistry problems arranged by topic and difficulty level, many with multiple solutions; the quickest way to solve each of these problems is included – a feature ideal for those students going for a perfect 800 or near perfect score!

Here's to your success on the SAT, in college, and in life.