

# **GENERAL CHEMISTRY TOPICAL:**

## **Stoichiometry Test 1**

Time: 23 Minutes\*  
Number of Questions: 18

\* The timing restrictions for the science topical tests are optional. If you are using this test for the sole purpose of content reinforcement, you may want to disregard the time limit.

**DIRECTIONS:** Most of the questions in the following test are organized into groups, with a descriptive passage preceding each group of questions. Study the passage, then select the single best answer to each question in the group. Some of the questions are not based on a descriptive passage; you must also select the best answer to these questions. If you are unsure of the best answer, eliminate the choices that you know are incorrect, then select an answer from the choices that remain. Indicate your selection by blackening the corresponding circle on your answer sheet. A periodic table is provided below for your use with the questions.

**PERIODIC TABLE OF THE ELEMENTS**

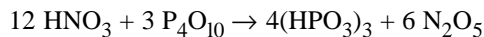
1 <b>H</b> 1.0																	2 <b>He</b> 4.0
3 <b>Li</b> 6.9	4 <b>Be</b> 9.0											5 <b>B</b> 10.8	6 <b>C</b> 12.0	7 <b>N</b> 14.0	8 <b>O</b> 16.0	9 <b>F</b> 19.0	10 <b>Ne</b> 20.2
11 <b>Na</b> 23.0	12 <b>Mg</b> 24.3											13 <b>Al</b> 27.0	14 <b>Si</b> 28.1	15 <b>P</b> 31.0	16 <b>S</b> 32.1	17 <b>Cl</b> 35.5	18 <b>Ar</b> 39.9
19 <b>K</b> 39.1	20 <b>Ca</b> 40.1	21 <b>Sc</b> 45.0	22 <b>Ti</b> 47.9	23 <b>V</b> 50.9	24 <b>Cr</b> 52.0	25 <b>Mn</b> 54.9	26 <b>Fe</b> 55.8	27 <b>Co</b> 58.9	28 <b>Ni</b> 58.7	29 <b>Cu</b> 63.5	30 <b>Zn</b> 65.4	31 <b>Ga</b> 69.7	32 <b>Ge</b> 72.6	33 <b>As</b> 74.9	34 <b>Se</b> 79.0	35 <b>Br</b> 79.9	36 <b>Kr</b> 83.8
37 <b>Rb</b> 85.5	38 <b>Sr</b> 87.6	39 <b>Y</b> 88.9	40 <b>Zr</b> 91.2	41 <b>Nb</b> 92.9	42 <b>Mo</b> 95.9	43 <b>Tc</b> (98)	44 <b>Ru</b> 101.1	45 <b>Rh</b> 102.9	46 <b>Pd</b> 106.4	47 <b>Ag</b> 107.9	48 <b>Cd</b> 112.4	49 <b>In</b> 114.8	50 <b>Sn</b> 118.7	51 <b>Sb</b> 121.8	52 <b>Te</b> 127.6	53 <b>I</b> 126.9	54 <b>Xe</b> 131.3
55 <b>Cs</b> 132.9	56 <b>Ba</b> 137.3	57 <b>La</b> * 138.9	72 <b>Hf</b> 178.5	73 <b>Ta</b> 180.9	74 <b>W</b> 183.9	75 <b>Re</b> 186.2	76 <b>Os</b> 190.2	77 <b>Ir</b> 192.2	78 <b>Pt</b> 195.1	79 <b>Au</b> 197.0	80 <b>Hg</b> 200.6	81 <b>Tl</b> 204.4	82 <b>Pb</b> 207.2	83 <b>Bi</b> 209.0	84 <b>Po</b> (209)	85 <b>At</b> (210)	86 <b>Rn</b> (222)
87 <b>Fr</b> (223)	88 <b>Ra</b> 226.0	89 <b>Ac</b> † 227.0	104 <b>Unq</b> (261)	105 <b>Unp</b> (262)	106 <b>Unh</b> (263)	107 <b>Uns</b> (262)	108 <b>Uno</b> (265)	109 <b>Une</b> (267)									

*	58 <b>Ce</b> 140.1	59 <b>Pr</b> 140.9	60 <b>Nd</b> 144.2	61 <b>Pm</b> (145)	62 <b>Sm</b> 150.4	63 <b>Eu</b> 152.0	64 <b>Gd</b> 157.3	65 <b>Tb</b> 158.9	66 <b>Dy</b> 162.5	67 <b>Ho</b> 164.9	68 <b>Er</b> 167.3	69 <b>Tm</b> 168.9	70 <b>Yb</b> 173.0	71 <b>Lu</b> 175.0
†	90 <b>Th</b> 232.0	91 <b>Pa</b> (231)	92 <b>U</b> 238.0	93 <b>Np</b> (237)	94 <b>Pu</b> (244)	95 <b>Am</b> (243)	96 <b>Cm</b> (247)	97 <b>Bk</b> (247)	98 <b>Cf</b> (251)	99 <b>Es</b> (252)	100 <b>Fm</b> (257)	101 <b>Md</b> (258)	102 <b>No</b> (259)	103 <b>Lr</b> (260)

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**Passage I (Questions 1–7)**

Tetraphosphorus decaoxide,  $P_4O_{10}$ , a powerful dehydrating agent that can be used in the preparation of nitric anhydride,  $N_2O_5$ . In this reaction,  $P_4O_{10}$  is used to remove water from concentrated nitric acid as shown in Reaction 1.



Reaction 1

Using this reaction, a chemist investigated which concentrations of reactants would give the highest yield of products. Three experiments were performed in which varying amounts of tetraphosphorus decaoxide were used; the results of these experiments are summarized in Table 1.

Table 1

Reactants Used and Products Produced				
Experiment	$\text{HNO}_3$	$\text{P}_4\text{O}_{10}$	$(\text{HPO}_3)_3$	$\text{N}_2\text{O}_5$
#1	126 g	excess	40 g	54 g
#2	126 g	200 g	100 g	4 g
#3	126 g	71 g	50 g	10 g

[Note: The molecular weight of  $\text{HNO}_3$  is 63.0 g; of  $\text{P}_4\text{O}_{10}$  is 284.0 g; of  $(\text{HPO}_3)_3$  is 240.0 g; and of  $\text{N}_2\text{O}_5$  is 108.0 g.]

It is interesting to note that nitric anhydride,  $\text{N}_2\text{O}_5$ , in the solid state exists as an ionic compound: it is composed of the nitronium ion,  $\text{NO}_2^+$ , and the nitrate ion,  $\text{NO}_3^-$ . In the gas phase it exists in the molecular form shown by its formula in Reaction 1. Nitric anhydride sublimates at  $32.4^\circ\text{C}$ , and due to its explosive nature as a solid, it must be handled with great care.

- Nitric anhydride is thermally unstable and decomposes into nitrogen dioxide gas and oxygen gas. If all the  $\text{N}_2\text{O}_5$  produced in Experiment 1 decomposed in this way, how much oxygen gas would be produced?
  - 8 g
  - 16 g
  - 24 g
  - 32 g
- What is the theoretical yield of  $(\text{HPO}_3)_3$  in Experiment #1?
  - 66 g
  - 160 g
  - 240 g
  - 480 g
- Which is the limiting reagent in Experiment #3?
  - $\text{HNO}_3$
  - $\text{P}_4\text{O}_{10}$
  - $(\text{HPO}_3)_3$
  - $\text{N}_2\text{O}_5$
- What is the percent yield of  $\text{N}_2\text{O}_5$  in Experiment #1?
  - 10
  - 25
  - 50
  - 100
- In Experiment #3, how many moles of nitric acid remain after the reaction is complete?
  - 0.0
  - 0.5
  - 1.0
  - 1.5
- Which of the following statements is true about stoichiometry calculations?
  - One should get the same mass of phosphorus-containing product as the mass of phosphorus-containing starting material used.
  - The limiting reagent is the one present in the smallest amount by mass.
  - The limiting reagent is the reagent present in the smallest molar amount.
  - The amount of the limiting reagent used will determine the amount of products obtained.

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7. In Experiment #1, how much tetraphosphorus decaoxide should have been used if the chemist wanted exactly enough of it to react with the nitric acid without any excess?
- A. 142 g
  - B. 284 g
  - C. 568 g
  - D. 852 g
- 

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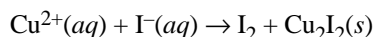
**Passage II (Questions 8–13)**

The transition metals are of central importance to aqueous redox chemistry because they often have a number of stable oxidation states, whereas the main group metals, such as potassium and calcium, usually have only one stable oxidation state. Copper(II) is generally stable in aqueous solutions and many copper(II) compounds are familiar. However, the copper(I) cation is not stable in aqueous solution because it disproportionates according to the following reaction:



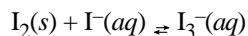
## Reaction 1

Initially unaware of this fact, a student planned to prepare copper(I) iodide by treating copper(II) with a mild reducing agent in the presence of iodide. The student was surprised after adding a soluble copper(II) salt to the solution of potassium iodide when a precipitate formed immediately. The precipitate was filtered from the solution and subsequent analysis revealed it to be  $\text{Cu}_2\text{I}_2$ . This suggested the reaction proceeded in the manner given by Reaction 2.



## Reaction 2

Molecular iodine is only slightly soluble in water, to the extent of 0.3 g/L of water at room temperature. The molecular iodine, which would normally precipitate out of solution, is solubilized here by the presence of iodide, which was present in excess. Reaction 3 shows the favorable equilibrium that accounts for this phenomenon.



## Reaction 3

8. What is the most likely reason for the apparent stability of  $\text{Cu}_2\text{I}_2$ ?
- The copper is actually in the +2 oxidation state.
  - $\text{Cu}_2\text{I}_2$  contains copper in both the 0 and +2 oxidation states.
  - The potassium from the KI keeps it from being oxidized.
  - It does not react because it is insoluble.

9. In Reaction 1, copper(I) acts as a(n):

- oxidizing agent.
- reducing agent.
- complexing agent.

- I only
- II only
- I and II only
- I, II, and III

10. In Reaction 2, in addition to acting as a precipitating reagent, iodide acts as a(n):

- oxidizing agent
- reducing agent.
- complexing agent.

- II only
- III only
- I and II only
- I, II, and III

11. When Reaction 2 is balanced, if the stoichiometric coefficient for  $\text{Cu}_2\text{I}_2$  is 1, the stoichiometric coefficient for iodide as:

- 1.
- 2.
- 3.
- 4.

12. If the student began with 83.2 g of copper(II) sulfate pentahydrate,  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ , how much  $\text{Cu}_2\text{I}_2$  could be made?

[Note: The molecular weights of  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  and  $\text{Cu}_2\text{I}_2$  are 249.5 g and 380.8 g, respectively.]

- 31.8 g
- 63.5 g
- 127 g
- 190 g

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13. What would be expected to happen if 0.1 mol of  $I_2$  is added to a 100mL solution containing 1.0 mol of precipitated AgI?
- A. Both the AgI and  $I_2$  would completely dissolve.
  - B. All the  $I_2$  would dissolve but only a portion of the AgI would dissolve.
  - C. All the AgI would dissolve but only a portion of the  $I_2$  would dissolve.
  - D. Neither the AgI nor the  $I_2$  would dissolve.
- 

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Question 14 through 18 are **NOT**  
based on a descriptive passage.

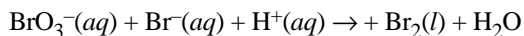
14. What is the molecular formula of a compound with the empirical formula  $C_3H_6O_2$  and a mass of 148 amu?

- A.  $C_6H_{12}O_4$
- B.  $C_2H_6O_2$
- C.  $C_9H_{18}O_6$
- D.  $C_2H_3O$

15. An oxide of arsenic contains 65.2% arsenic by weight. What is its simplest formula?

- A. AsO
- B.  $As_2O_3$
- C.  $AsO_2$
- D.  $As_2O_5$

16. In the following unbalanced reaction:



the ratio of bromate to bromide is:

- A. 1:5.
- B. 1:3.
- C. 1:2.
- D. 1:1.

17. What is the mass of nitrogen in a 50.0 g sample of sodium nitrite?

- A. 20.2 g
- B. 16.4 g
- C. 10.1 g
- D. 8.23 g

18. How many atoms are in a 365 g sample of  $SF_6$  gas?

- A.  $1.51 \times 10^{22}$
- B.  $1.06 \times 10^{22}$
- C.  $1.51 \times 10^{23}$
- D.  $1.06 \times 10^{24}$

END OF TEST

**ANSWER KEY:**

- |    |          |     |          |     |          |     |          |
|----|----------|-----|----------|-----|----------|-----|----------|
| 1. | <b>A</b> | 6.  | <b>D</b> | 11. | <b>D</b> | 16. | <b>A</b> |
| 2. | <b>B</b> | 7.  | <b>A</b> | 12. | <b>B</b> | 17. | <b>C</b> |
| 3. | <b>B</b> | 8.  | <b>D</b> | 13. | <b>B</b> | 18. | <b>D</b> |
| 4. | <b>C</b> | 9.  | <b>C</b> | 14. | <b>A</b> |     |          |
| 5. | <b>C</b> | 10. | <b>A</b> | 15. | <b>D</b> |     |          |

## STOICHIOMETRY TEST 1 EXPLANATIONS

## Passage I (Questions 1–7)

The first passage concerns the study of a reaction between nitric acid and tetraphosphorus decaoxide. Three "runs" of this reaction were performed with varying amounts of  $\text{P}_4\text{O}_{10}$ . That the reaction is a dehydration is interesting but turns out to be of no use in solving any of the problems. You should become used to the idea that not every fact given to you in a passage is essential or even relevant to solving the problems following them.

**1. The correct choice for question 1 is A.** The first step in solving this question is to write and balance the equation for the decomposition reaction. Nitric anhydride decomposes into nitrogen dioxide and oxygen. Because there are two atoms of nitrogen in nitric anhydride and only one nitrogen atom in nitrogen dioxide, and since there is no other nitrogen-containing compound, each mole of nitric anhydride must produce two moles of nitrogen dioxide. After putting the stoichiometric coefficient of two in front of the nitrogen dioxide, we see that to balance the oxygen atoms, we have to put a coefficient of  $1/2$  in front of the  $\text{O}_2$ . Many people don't like fractional coefficients so we may double them all if it makes you feel better. Because the coefficients tell us the ratio of the products and reactant to each other, multiplying them all by a constant does not change the ratio. Therefore, saying that one mole of nitric anhydride decomposes to give two moles of nitrogen dioxide and half a mole of oxygen gas is equivalent to saying that two moles of nitric anhydride decomposes to give four moles of nitrogen dioxide and one mole of oxygen gas. The important thing is that only half as many moles of oxygen gas are produced as moles of nitric anhydride decompose. In experiment one we see that 54 grams of nitrogen dioxide are made. Dividing 54 grams by the molecular weight of nitric anhydride, 108 grams per mole, gives the number of moles of nitric anhydride,  $1/2$ . Because only half the number of moles of oxygen gas would result from the decomposition, one-quarter of a mole of oxygen is produced. Multiplying  $1/4$  mole by the molecular weight of oxygen gas, 32 grams per mole (remember: oxygen gas is diatomic), gives 8 grams of oxygen produced. The correct choice is A.

**2. The correct choice for question 2 is B.** Because we know that the tetraphosphorus decaoxide is in excess, the nitric acid must be the limiting reagent, determining how much triphosphoric acid is produced. Using the molecular weight of nitric acid, you know that 126 grams of nitric acid is 2 moles of nitric acid. The balanced equation of Reaction 1 tells me that for every 12 moles of nitric acid used I will produce 4 moles of triphosphoric acid. The mole ratio between nitric acid and triphosphoric acid is therefore 3 to 1. 2 moles of nitric acid should produce  $2/3$  of a mole of triphosphoric acid. The molecular weight of triphosphoric acid is given in the table as 240 grams per mole. Multiplying this molecular weight by  $2/3$  of a mole, we find that we should get 160 grams of triphosphoric acid. Although this is the amount we would expect to get, called the theoretical yield, you can see from Table 1 that only 40 grams were actually made, and the actual yield is lower. Though this fact is interesting, it is not necessary to solve the problem. The amount of triphosphoric acid theoretically possible from experiment 1 is 160 grams, which corresponds to choice B.

**3. The correct answer for question 3 is choice B.** The limiting reagent is the reactant that is completely used up during the reaction, limiting the amount of product that can be made, though additional co-reactant remains. Looking at the balanced equation we see that the ratio in the number of moles of nitric acid to tetraphosphorus decaoxide necessary for the reaction is 12 to 3. In other words, 4 moles of nitric acid are needed for each mole of tetraphosphorus decaoxide. Experiment 3 began with 126 grams of nitric acid. If we divide this mass by the molecular weight of nitric acid, 63 grams per mole, we find that this is 2 moles. Because we only need  $1/4$  of the number of moles of tetraphosphorus decaoxide that we need of nitric acid, here we need a total of only  $1/2$  of a mole of tetraphosphorus decaoxide. The table shows that tetraphosphorus decaoxide has a molecular weight of 284 grams per mole. If we multiply this by the  $1/2$  of a mole we need, we find we need 142 grams of tetraphosphorus decaoxide. Since we have less than this, 71 grams, tetraphosphorus decaoxide must be the limiting reagent. If we had exactly the amount of tetraphosphorus decaoxide necessary for the reaction, 142 grams, neither reactant would be the limiting reagent and both reactants would be completely used up with nothing remaining. However, in experiment 3 there is only 71 grams of tetraphosphorus decaoxide, limiting the extent of the reaction. Tetraphosphorus decaoxide is the limiting reagent. Again, the correct choice is B.

**4. The correct choice for question 4 is C.** The percent yield of product is the actual yield divided by the theoretical yield times 100%. If the actual yield, the amount actually produced, equals the theoretical yield, the amount we would predict from the balanced equation and the amounts of reactant we began with, then the product yield would be 100%. The percent yield can therefore vary from 0 to 100%. Looking at the balanced equation, Reaction 1, we see that the mole ratio of nitric acid to nitric anhydride is 12 to 6 or, more simply, 2 to 1. Because I know that 126 grams is 2 moles of nitric acid, I would expect to make 1 mole of nitric anhydride based on Reaction 1. Nitric anhydride has a given molecular weight of 108 grams per mole, so I would expect to get 108 grams of it from Experiment 1. Looking at Table 1, I see that we only get 54 grams of nitric anhydride. Dividing my actual yield, 54 grams, by my theoretical yield, 108 grams, and multiplying by 100%, I find my percent yield is 50%. This means that I have gotten half the amount of nitric anhydride I would have expected given the amount of nitric acid I had started with. The correct choice is therefore C.

**5. The answer for problem 5 is C.** In determining the reagent in excess we must compare the mole ratios of the reactants weighed out for Experiment 3 to the mole ratios in Reaction 1. The balanced equation of Reaction 1 tells me that I need 12 moles of nitric acid for every 3 moles of tetraphosphorus decaoxide. Another way to say this would be that the mole ratio of nitric acid to tetraphosphorus decaoxide is 12 to 3 or, more simply, 4 to 1. I know that 126 grams of nitric acid is 2 moles. Given the balanced equation I know that I should have 1/2 mole of tetraphosphorus decaoxide to react completely with it; if I have less than 1/2 mole of tetraphosphorus decaoxide, the nitric acid is in excess, if I have more than 1/2 mole, the tetraphosphorus decaoxide is in excess. In this problem we are told that there is at least enough nitric acid to react with all the tetraphosphorus decaoxide, and perhaps there is an excess. But this does not help us much, we still have to determine how much nitric acid reacted. The molecular weight of tetraphosphorus decaoxide is given as 284 grams per mole. Looking at Experiment 3 in Table 1, I see that we have 71 grams of tetraphosphorus decaoxide; this is 1/4 mole. This should react with 1 mole of nitric acid and leave 1 mole unreacted. The answer is therefore C.

**6. The answer for question 6 is D.** The problem requires knowledge of the definition of a limiting reagent. The limiting reagent is not necessarily present in the smallest amount or in the smallest molar amount. In fact, Passage I itself provides an example of the falsity of the latter statement. Reaction 1 shows that we need 12 moles of nitric acid to react with 3 moles of tetraphosphorus decaoxide. If we had only 10 moles of nitric acid and 3 moles of tetraphosphorus decaoxide, nitric acid would be the limiting reagent though there were more moles of it to begin with. Choice A is equally untrue, the phosphorus-containing reactant need not produce the same mass of phosphorus-containing product for some atoms of the reactant could end up in other non-phosphorus-containing products. The limiting reagent is the starting material that we have less of than we need to react with its co-reactant. Because it is completely consumed, the reaction stops before the co-reactant is completely used up. Therefore, the limiting reagent determines the amount of all products that are made by the reaction. Choice D is the correct answer.

**7. The correct answer for question 7 is A.** In Experiment 1, as in all the experiments, the quantity of nitric acid is 2 moles. From the balanced equation in Reaction 1 we know we need 1/4 as many moles of tetraphosphorus decaoxide as nitric acid; that would be 1/2 mole. Again, the molecular weight of tetraphosphorus decaoxide is given in the table as 284 grams per mole; multiplying this by 1/2 mole gives 142 grams. This is the maximum amount of tetraphosphorus decaoxide that could be consumed by 2 moles of nitric acid in this reaction. The answer is therefore A.

That brings us to the end of Passage I. Your ability to balance equations and to convert grams to moles and back again, where needed, was the skill most useful in successfully navigating this passage. While these skills were on the forefront of these problems, you should remember that they are very rarely absent from chemistry problems in some form. Passage II continues the emphasis on this area.

#### Passage II (Questions 8–13)

The second passage introduces the oxidation-reduction or redox properties of the transition metals, in this case using the example of copper. Like all passages, the bits of information the passage provides may or may not be useful in solving the problems that follow. Initially we are told that the copper(II) cation, sometimes called the cupric ion, is stable in aqueous solution and the copper(I) cation, also known as the cuprous ion, is not. The copper(I) cation in fact reacts with another copper(I) cation to form copper metal and copper(II) ion. Following this fact we are apparently presented with a paradox: a student in mixing a copper(II) salt with potassium iodide recovers a precipitate that has the formula  $\text{Cu}_2\text{I}_2$ , which would suggest a copper(I) salt was made in aqueous solution. Lastly we are told of a reaction that takes place incidentally to the copper reaction, a favorable equilibrium between iodine, the element, and iodine, an anion, to form the soluble triiodide ion. Given these three facts and knowledge of stoichiometry and basic chemistry you should be able to solve the problems that follow.

**8. The answer to question 8 is D.** We have been told that copper(I) ion is unstable in aqueous solution. That last point turns out to be critical, the fact that the ion is unstable in *solution*. It does not say that copper(I) is unstable always, just in aqueous solution. That would seem to explain the stability of the copper(I) iodide. The copper(I) iodide is insoluble, which we know because it precipitated out of solution; thus, most of the copper(I) is present in the solid, not in aqueous solution where it can disproportionate. This is not to say that there is no copper(I) in solution. There is copper(I) in solution but it is minuscule compared to the amount that was precipitated. As time goes on, the copper(I) in solution disproportionates and the solubility equilibrium for copper(I) iodide would shift to dissolve the solid and produce more copper(I) in solution. However, at any one time the amount of copper(I) in solution is small and so the dissolution of copper(I) iodide is very slow and actually stops when the precipitate is filtered and removed from water. The other choices have various problems. Choice A would seem to require the iodide ion to be in the minus two oxidation state to balance out the charge of the coppers, since the molecule is uncharged overall. This should strike you as a very unusual oxidation state for iodine. You should realize that iodine in the minus two oxidation state would have 9 electrons in its valence shell rather than the favored octet of the more familiar iodide ion. Iodine in the  $-2$  oxidation state does not, in fact, exist. Choice B may fool you initially. There are many examples of compounds, usually oxides, containing elements in two oxidation states. One example is  $\text{Fe}_3\text{O}_4$ , where iron is in both its  $+2$  and  $+3$  oxidation states. The IUPAC name for this compound would be

iron(II) diiron(III) oxide. But a closer look at choice B shows that this would be an error. Copper in the zero oxidation state is in the form of the metal and would be unlikely to be a constituent of a compound. It would also seem unusual that iodide would selectively reduce only half of the available copper(II) to copper(0). Choice C is also wrong. The potassium ion is generally a spectator ion in solution because it is impossible to oxidize or reduce it in aqueous solution. There is no reason to believe that it would prevent any reaction, let alone oxidation. Choice D is therefore correct.

**9. The correct choice for question 9 is C.** In answering the question correctly you must have a good idea of what oxidizing and reducing agents are. Oxidation is the loss of electrons and reduction is the gain of electrons. Oxidation and reduction are therefore complimentary phenomena because they are reverse processes. Where there is an oxidation there must be a corresponding reduction, because one species transfers electrons and the other accepts them. An oxidizing agent is a species that causes oxidation or equivalently, gains electrons. An oxidizing agent is therefore reduced. A reducing agent is a species causes reduction or loses electrons. A reducing agent is thereby oxidized. In Reaction 1, two copper(I) ions react together, one gaining an electron to become copper in the zero oxidation state, and the other copper thereby losing an electron to become copper(II). Thus, one of the copper(I) cations behaves as an oxidizing agent and another copper(I) cation behaves as a reducing agent. Therefore, choices C and D would seem to be possible. To determine the correct answer we must know what a complexing agent is. A complexing agent is a molecule or ion that acts as a Lewis base, donating its electron pairs to a central metal cation, generally making it more soluble. Common complexing agents are many polar solvents, such as water and ammonia, and simple and complex ions such as cyanide and EDTA. The complexing reaction is usually very favorable and probably accounts in large part for the solubility of many ionic compounds in water. Copper(I) is not behaving like a complexing agent here, only as an oxidizing and reducing agent. Choice C is therefore correct.

**10. The answer for question 10 is A.** This is similar in style to the previous question. In Reaction 2 we can see that at least some iodide is being converted to molecular, or elemental, iodine. The oxidation state of iodine as iodide is minus one, the oxidation state of iodine as the element is zero. Thus the iodide is losing its electrons and is being oxidized; therefore iodide is acting as a reducing agent in this reaction. This would eliminate choice B because iodide is acting at least as a reducing agent which is not expressed in choice B. However there is no indication that iodide is acting as an oxidizing agent, in fact iodide can never act as an oxidizing agent. Iodide cannot accept additional electrons because it has a complete octet of electrons, which very stable, and is in its lowest oxidation state, minus 1. Thus the remaining choices are eliminated without even coming to the question of whether iodide behaves as a complexing agent. As noted in the prior question a complexing agent acts as a Lewis base to a metal cation, binding to the metal. Though iodide has free electron pairs and can act as a Lewis base and complexing agent to metals, there is no indication that it is doing so. In fact, there is every indication that it is not, as complexing agents generally increase the solubility of a metal cation. Therefore the correct answer for question 10 is choice A.

**11. The correct answer for question 11 is choice D.** The correct stoichiometric coefficients relate the ratios of all the reactants and products involved in a reaction. We call this balancing the equation. It is important to realize that though the ratios for a particular reaction are unique, the coefficients may be different. In other words, to say that 1 mole of A reacts with 2 moles of B to give 1 mole of C is the same as saying 2 moles of A react with 4 moles of B to give 2 moles of C. Though the coefficients would be different, the ratio between A, B, and C would be constant. Though most people have a preference for the smallest whole number integer coefficients that are possible for a given reaction, this is not mandatory and you will sometimes see equations that do not follow this convention and even have fractional coefficients. In this case we are told that the coefficient in front of the copper(I) iodide is 1; this fixes the other coefficients in this reaction and tells us to find the moles of the other species that would result in 1 mole of copper(I) iodide. Because copper(I) iodide has two copper atoms in its molecular formula, we actually have 2 moles of copper ion in this product. As copper is found in no other product, all the copper from the reactants must be in the copper(I) iodide. Looking to the left side of the reaction equation, we see that only one reactant has copper, namely the copper(II) ion. Because we have 2 moles of copper ion in the product we must have 2 moles of copper provided by the reactant. We therefore put a 2 in front of the copper(II) ion to indicate this. Now we turn to the iodide. Iodide both reduces the copper(II) ion and reacts with the copper(I) ion to form the product precipitate, copper(I) iodide. Because we have 1 mole of copper(I) iodide and each contains 2 iodide ions, we need 2 iodide ions to form the copper iodide product. But we also need iodide to reduce the copper(II) ion to copper(I) ion. The iodide, in acting as a reducing agent here, is oxidized from the  $-1$  oxidation state of iodide to the  $0$  oxidation state of molecular iodine, transferring 1 electron to the copper(II) ion. Since we need to convert 2 moles of copper(II) ion to copper(I) ion, each conversion requiring 1 electron, we need 2 moles of electrons which are provided by 2 moles of iodide. Therefore, we need a total of 4 moles of iodide: 2 moles to reduce the copper(II) ion to copper(I) ion and 2 moles to react with the copper(I) ion to form copper(I) iodide. These 4 moles of iodide must appear in some form in the products. 2 of the 4 moles are accounted for in the 1 mole of copper(I) iodide, the other 2 moles must be in the iodine product. Because molecular iodine is diatomic, that is, it contains 2 atoms of iodine, we only need 1 mole of elemental iodine to account for the remaining 2 moles of iodide. The stoichiometric coefficients in the balanced equation are therefore, reading left to right, 2, 4, 1, and 1. Were it not for the fact that we are told the stoichiometric coefficient in front of the copper(I) iodide is one, they might just have easily have been written 1, 2,  $1/2$ , and  $1/2$ , as the ratio between all the reactant and product species would be preserved. In any case, the coefficient in front of the iodide is 4 and the correct choice is D.

**12. The correct answer for question 12 is choice B.** This question deals with straightforward reaction stoichiometry. Since the molecular weight of copper(II) sulfate pentahydrate is given as 249.5 grams per mole, 83.2 grams of it is about 1/3 of a mole. 1/3 of a mole of copper(II) sulfate pentahydrate would provide 1/3 of a mole of copper(II). This amount of copper(II) will make 1/6 of a mole of copper(I) iodide, since each mole of that product requires 2 moles of copper. The molecular weight of copper(I) iodide is 380.8 grams per mole. Multiplying this by 1/6 gives 63.5 grams. The correct choice is therefore B.

**13. The answer to question 13 is B.** A few different skills are necessary here. One is a good command of equilibria. This passage introduces Reaction 3 as the explanation for the solubility of iodine in iodide solutions. Of particular importance, we are told this reaction is very favorable. Silver iodide, you should know, is very insoluble and in fact we are told it is precipitated. In normal circumstances, the concentration of iodide in solution is very slight but it is in dynamic equilibrium with the solid silver iodide. If the iodide in solution is removed in any way, the equilibrium is disturbed, and more silver iodide dissolves. When iodine is added to the silver iodide solution it converts the small amount of iodide in solution to triiodide and silver iodide dissolves to replenish the iodide removed until equilibrium is reestablished. A system in equilibrium responds to a stress on that system by acting in such a way to minimize the effect of that stress. Here, iodine stresses the silver iodide system by removing iodide from solution by converting it into triiodide ion. Thus choice D may be eliminated, because at least some iodine and silver iodide would dissolve. Note that we have 1 mole of silver iodide. In contrast, there is only 1/10 of a mole of iodine, and so the iodide can only make a maximum of 1/10 of a mole of triiodide before it is all used up. Having been used up, the iodine is completely dissolved, but only 1/10 of a mole of silver iodide could have dissolved, leaving 9/10 of a mole undissolved. Choice B is therefore correct.

That brings us to the end of Passage II. The important areas covered included many aspects of stoichiometry, such as balancing redox and other reaction equations and the interconversion of moles and grams by using molecular weights. In addition, background knowledge of oxidation and reduction, chemical equilibrium, and nomenclature was involved. However, like most passages in the MCAT, this passage requires to use your knowledge of chemistry to solve situations which you may never have seen before by providing enough background information to allow you to answer the questions that are related to it.

#### Discrete Questions

**14. The answer to question 14 is A.** The empirical formula of a compound is the formula that shows the smallest whole number relationship between the elements of a molecule. The empirical formula for this compound,  $C_3H_6O_2$ , tells us that for every 3 carbon atoms in the molecule there are 6 hydrogen atoms and 2 oxygen atoms. This does not mean that a molecule actually contains 3 carbons, 6 hydrogens, and 2 oxygens (though it might), instead it tells us the ratio between them. Therefore, the molecular formula could be choice C,  $C_9H_{18}O_6$ , where the carbon, hydrogen, and oxygen ratio is 3 to 6 to 2 as in the empirical formula, and might also be choice A where the ratio is correct, but it cannot be choices B or D, where the ratio between the elements is different from the empirical formula. Knowing that the mass of the molecule is 148 amu (or atomic mass units) distinguishes between choices A and C, which are equally likely before this piece of information. Amu stands for atomic mass units and means that the compound has a molecular weight of 148 grams per mole. Choice A has a molecular weight of 148 grams per mole; choice C has a molecular weight of 222 grams per mole. Choice A must therefore be the correct answer.

**15. The answer for question 15 is D.** Perhaps the easiest way to solve such a problem is to imagine a particular sample mass of the compound. We shall choose a 100 grams for convenience, though any mass will arrive at the correct answer. Because the oxide of arsenic contains only arsenic and oxygen, a 100 gram sample would contain 65.2 grams of arsenic and the remainder, 34.8 grams, must be oxygen. To find the ratio between these two elements in the compound we divide the mass of arsenic by the atomic weight of arsenic, 74.7 grams per mole, and the mass of oxygen by the atomic weight of oxygen, 16 grams per mole. This will give the mole ratio between arsenic and oxygen in the compound. To convert to a more easily useful ratio, we divide both by the lowest number of the two, here arsenic. This gives us a ratio of 1 mole arsenic to 2.5 moles oxygen, which is better stated by doubling both numbers and getting 2 moles arsenic to 5 moles of oxygen. This would correspond to the formula in the correct answer, choice D.

**16. The answer to question 16 is A.** This is a redox reaction. We can tell this because the oxidation state of species change during the reaction. The oxidation state of bromine in the product, the element bromine, is 0. The oxidation state of the bromine in the bromide reactant is  $-1$  and the oxidation state of the bromine in the bromate is  $+5$ . The fact that the bromine of one reactant reduces the bromine of the other should not bother you, it is immaterial. The important thing in redox reactions is the number of electrons. Since bromide is acting as a reducing agent, going from the  $-1$  oxidation state to the 0 oxidation state, it transfers one electron per bromide. Bromate, however, is going from the  $+5$  oxidation state to the 0 oxidation state, requiring 5 electrons. Bromate gets these electrons from bromide, requiring 5 of them. The ratio of bromate to bromide is therefore 1 to 5. The correct answer is choice A.

**17. The answer to question 17 is choice C.** The first step needed to solve the problem shows how essential it is to know your nomenclature. Even if you knew how to solve this problem, not knowing the formula prevents you from applying that knowledge. The formula for sodium nitrite is  $\text{NaNO}_2$ . The nitrite ion is a common anion with a charge of  $-1$ . The molecular weight of sodium nitrite is 69 grams per mole. As there is 1 nitrogen atom per formula unit we can find weight fraction of sodium nitrite that is nitrogen by dividing the atomic weight of nitrogen, 14 grams per mole, by the molecular weight of sodium nitrite, 69 grams per mole. If we multiply this fraction,  $14/69$ , by 50 grams we get 10.1 grams, or choice C. Notice that you did not have to determine the number of moles of sodium nitrite to solve this problem, though you could have. The calculations are actually identical: one could have instead found the number of moles of sodium nitrite in a 50 gram sample and, realizing that each mole of sodium nitrite has one atom of nitrogen, multiply that number by the atomic weight of nitrogen. This illustrates an important aspect of solving chemistry problems: there are often many ways of looking at a problem and each is merely a change in perspective. Whether you view this problem in either of these two ways, you are thinking about the problem correctly. Beware the solution methods that do not require thought and are mechanical; they are the methods that fail where the problem does not fall within a type you are used to. Again, the correct answer for problem 17 is C.

**18. The answer for problem 18 is D.** This is a relatively simple calculation but really tests your ability to read a problem carefully. The molecular weight of sulfur hexafluoride is 146 grams per mole. If I divide the mass given, 36.5 grams, by the molecular weight, I find that we have  $1/4$  of a mole of sulfur hexafluoride molecules. A mole is Avogadro's Number of particles, in other words, there are  $6.022 \times 10^{23}$  particles per mole. If we multiply the  $1/4$  of a mole of sulfur hexafluoride by Avogadro's Number we find the number of molecules of sulfur hexafluoride,  $1.51 \times 10^{23}$ . We might be tempted to circle choice C now, but this is wrong:  $1.51 \times 10^{23}$  is the number of *molecules* of sulfur hexafluoride in the sample, not the number of atoms. There are seven atoms in each molecule, so we must now multiply the number of molecules by the number of atoms per molecule, seven. This trick is easily seen if you make sure all your units cancel when you use an equation. The answer is therefore D,  $1.06 \times 10^{24}$  atoms.