

Exploring Creation with Chemistry 3<sup>rd</sup> Edition - Errata file instructions

Solutions and Tests Manual – 4<sup>TH</sup> PRINTING

This file contains the corrected pages for the 4<sup>th</sup> Printing Solutions and Tests Manual. (Updated July 2024). See last page for list of newest updates.

You can find which printing you have by going to the publications page in the front of your book. The printing for the Textbook and Solutions and Tests Manual may not be the same. Corrections for the Textbook are in a separate file.

The items highlighted in **bright green** are the corrections to the errors in the book. Make sure your Solutions Manual reads the same as the corrections on these pages.

If you have a test packet, you should double check the following test(s) with the corresponding page in the Solutions Manual, which are corrected in this file:

Test 10: #1 – page 191

Test 11: #11 – page 213

To make 150.0 g, then, we just multiply the amount of each component by 2.47:

$$\text{Mass of nitrogen} = 50.0 \text{ g} \times 2.47 = 124 \text{ g}$$

$$\text{Mass of hydrogen} = 10.7 \text{ g} \times 2.47 = 26.4 \text{ g}$$

You need 124 g of nitrogen and 26.4 g of hydrogen to make 150.0 g of ammonia. The **numbers** actually add to 150.4 because of the rounding done for significant figures.

8. In this question, we already have the recipe for making methane with no leftovers (12.0 g carbon + 4.00 g hydrogen). When the chemist adds 150.0 grams of carbon, he is increasing the amount by:

$$12.0 \text{ g} \times x = 150.0 \text{ g}$$

$$x = \frac{150.0 \text{ g}}{12.0 \text{ g}} = 12.5$$

If the chemist added 12.5 times as much carbon, he should also add 12.5 times as much hydrogen:

$$4.00 \text{ g} \times 12.5 = 50.0 \text{ g}$$

However, the chemist added 150.0 g of hydrogen. This means that 100.0 g of hydrogen will not react. 100.0 g of hydrogen will be left over once the methane is made.

9. The reaction starts with  $54.0 \text{ g} + 100.0 \text{ g} = 154.0 \text{ g}$  of matter; thus, there must be 154.0 g of matter after everything is finished. According to the question, these amounts of aluminum and sulfur made 150.3 g of aluminum sulfide along with leftover sulfur. Since all 154.0 g must be accounted for, the remaining mass must be in the sulfur:

$$\text{Mass of sulfur} = \text{Total Mass} - \text{Mass of product}$$

$$\text{Mass of sulfur} = 154.0 \text{ g} - 150.3 \text{ g} = 3.7 \text{ g}$$

By the law of mass conservation, then, 3.7 g of sulfur were left over. Since we started out with 100.0 g of sulfur and 3.7 g were left over, then only  $100.0 \text{ g} - 3.7 \text{ g} = 96.3 \text{ g}$  were used to make aluminum sulfide. The proper recipe for making 150.3 g of aluminum sulfide is to add 54.0 g of aluminum to 96.3 g of sulfur.

## SAMPLE CALCULATIONS FOR EXPERIMENT 7.1

Number of drops to get 10.0 mL: 251

mL per drop:

$$\frac{10.0 \text{ mL}}{251 \text{ drops}} = 0.0398 \frac{\text{mL}}{\text{drop}}$$

Diameter of the circle: 12.75 cm

The volume of 1 drop is 0.0398 mL. To get the mass in 1 drop, we just multiply that by the density of the solution, which was given:

$$0.0398 \text{ mL} \times \frac{1.00 \text{ g}}{\text{mL}} = 0.0398 \text{ g}$$

That's the mass of the solution, but only a fraction of that is sodium stearate:

$$\text{Mass of sodium stearate added} = 0.0398 \text{ g} \times 0.000125 = 4.98 \times 10^{-6} \text{ g}$$

Now we need to convert to moles. First, we need to know the mass of a sodium stearate molecule:

$$1 \times 23.0 \text{ amu} + 18 \times 12.0 \text{ amu} + 35 \times 1.01 \text{ amu} + 2 \times 16.0 \text{ amu} = 306.4 \text{ amu}$$

This means:

$$1 \text{ mole NaC}_{18}\text{H}_{35}\text{O}_2 = 306.4 \text{ g NaC}_{18}\text{H}_{35}\text{O}_2$$

We can use that to convert to moles:

$$\frac{4.98 \times 10^{-6} \text{ g NaC}_{18}\text{H}_{35}\text{O}_2}{1} \times \frac{1 \text{ mole NaC}_{18}\text{H}_{35}\text{O}_2}{306.4 \text{ g NaC}_{18}\text{H}_{35}\text{O}_2} = 1.63 \times 10^{-8} \text{ moles NaC}_{18}\text{H}_{35}\text{O}_2$$

Since a mole contains  $6.02 \times 10^{23}$  molecules, we can determine the number of molecules:

$$\text{Number of molecules} = (1.63 \times 10^{-8} \text{ moles}) \times (6.02 \times 10^{23} \frac{\text{molecules}}{\text{mole}}) = 9.81 \times 10^{15} \text{ molecules}$$

11. To find the number of molecules from a kilogram measurement, we will need to convert to grams, then to moles, and then convert to molecules. First, we need to find the formula for nickel (IV) bromide. The (IV) is the charge on nickel, so the ion is  $\text{Ni}^{4+}$ . Bromide is  $\text{Br}^-$ , so the formula will be  $\text{NiBr}_4$ . Now we need to convert the kilograms into grams.  $1.78 \text{ kg} = 1780 \text{ grams}$ . Next, we need to find out how many grams are in 1 mole.

$$\text{Mass of NiBr}_4 = 58.7 \text{ amu} + 4 \times 79.9 \text{ amu} = 378.3 \text{ amu}$$

This means:

$$378.3 \text{ grams NiBr}_4 = 1 \text{ mole NiBr}_4$$

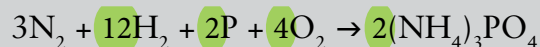
$$\frac{1780 \text{ g NiBr}_4}{1} \times \frac{1 \text{ mole NiBr}_4}{378.3 \text{ g NiBr}_4} = 4.71 \text{ moles NiBr}_4$$

Technically, this could be the final answer as the unit mole does tell us how many molecules there are. So if you stopped here then your answer is correct. However, you could also go one step farther by using Avogadro's number to convert to molecules.

$$\frac{4.71 \text{ moles NiBr}_4}{1} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mole}} = 2.84 \times 10^{24} \text{ molecules NiBr}_4$$

12. We start by determining the formula of the molecule. Ammonium is  $\text{NH}_4^+$ . Phosphate is  $\text{PO}_4^{3-}$ . We need to balance for charge so we need 3 ammonium ions with a charge of +1 to cancel the -3 charge from the phosphate. This gives us the formula for ammonium phosphate of  $(\text{NH}_4)_3\text{PO}_4$ .

Next write the balanced decomposition reaction.

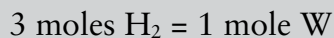


The balanced chemical reaction tells us that it takes 3 moles of  $\text{N}_2$  to produce 2 moles of  $(\text{NH}_4)_3\text{PO}_4$ . That allows us to convert from  $\text{N}_2$  to  $(\text{NH}_4)_3\text{PO}_4$ .

$$\frac{15 \text{ moles } (\text{NH}_4)_3\text{PO}_4}{1} \times \frac{3 \text{ moles } \text{N}_2}{2 \text{ moles } (\text{NH}_4)_3\text{PO}_4} = 23 \text{ moles } \text{N}_2$$

$$\frac{2.0 \times 10^6 \text{ g H}_2}{1} \times \frac{1 \text{ mole H}_2}{2.02 \text{ g H}_2} = 9.9 \times 10^5 \text{ moles H}_2$$

The chemical equation tells us:



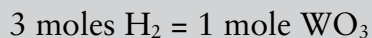
Thus, we can use it as a conversion relationship to convert moles of H<sub>2</sub> into moles of W:

$$\frac{9.9 \times 10^5 \text{ moles H}_2}{1} \times \frac{1 \text{ mole W}}{3 \text{ moles H}_2} = 3.3 \times 10^5 \text{ moles W}$$

Now we can convert to grams W:

$$\frac{3.3 \times 10^5 \text{ moles W}}{1} \times \frac{183.8 \text{ g W}}{1 \text{ mole W}} = 6.1 \times 10^7 \text{ g W}$$

3. We already have moles of H<sub>2</sub>, so we can use that to convert to moles of WO<sub>3</sub> using the chemical equation:



$$\frac{9.9 \times 10^5 \text{ moles H}_2}{1} \times \frac{1 \text{ mole WO}_3}{3 \text{ moles H}_2} = 3.3 \times 10^5 \text{ moles WO}_3$$

Now we can convert to grams WO<sub>3</sub>:

$$\text{Mass of WO}_3 = 1 \times 183.8 \text{ amu} + 3 \times 16.0 \text{ amu} = 231.8 \text{ amu}$$

$$1 \text{ mole WO}_3 = 231.8 \text{ g WO}_3$$

$$\frac{3.3 \times 10^5 \text{ moles WO}_3}{1} \times \frac{231.8 \text{ g WO}_3}{1 \text{ mole WO}_3} = 7.6 \times 10^7 \text{ g WO}_3$$

4. The first thing that we must do is convert to moles:

$$\frac{8.0 \times 10^3 \text{ g CCl}_4}{1} \times \frac{1 \text{ mole CCl}_4}{154 \text{ g CCl}_4} = 52 \text{ moles CCl}_4$$

$$\frac{1.0 \text{ g } \text{C}_2\text{H}_4\text{O}_2}{1} \times \frac{1 \text{ mole } \text{C}_2\text{H}_4\text{O}_2}{60.0 \text{ g } \text{C}_2\text{H}_4\text{O}_2} = 0.017 \text{ moles } \text{C}_2\text{H}_4\text{O}_2$$

Now we can determine the amount of base:

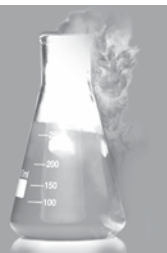
$$\frac{0.017 \text{ moles } \text{C}_2\text{H}_4\text{O}_2}{1} \times \frac{1 \text{ mole } \text{NH}_3}{1 \text{ mole } \text{C}_2\text{H}_4\text{O}_2} = 0.017 \text{ moles } \text{NH}_3$$

We used 10.0 mL of ammonia, so the concentration of the ammonia is:

$$\text{Concentration} = \frac{\# \text{ moles}}{\# \text{ liters}} = \frac{0.017 \text{ moles } \text{NH}_3}{0.010 \text{ L}} = \underline{1.7 \text{ M}}$$

**PLEASE NOTE:** Your answer may be quite different from ours, as the ammonia solution you used might have been stronger or more dilute than the one we used.

7. (4 pts) When  $\text{Mg}(\text{OH})_2$  dissolves in water, what ions form and how many of each ion are present?
- 1  $\text{Mg}^+$  ion, and  $\text{OH}^-$  ion
  - 2  $\text{Mg}$  ions, and 1  $\text{OH}$  ion
  - 2  $\text{Mg}^{2+}$  ions and 1  $\text{OH}^-$  ions
  - 1  $\text{Mg}^+$  ions and 2  $\text{OH}^{2-}$  ions
  - None of the above
8. (4 pts) What is an amphiprotic substance? Give an example of one.
9. (4 pts) The general equation for the reaction between an acid and a ionic base is:  
Base + Acid  $\rightarrow$
- Water
  - Ionic compound
  - Water + an ionic compound
  - Covalent compound
  - None of the above
10. (4 pts) Which of the following units would most likely be used as a concentration unit?
- mL/g
  - moles/g
  - %/moles
  - moles/mL
  - None of the above
11. (4 pts) Molarity is the moles of solute divided by the:
- Liters of solute
  - Liters of solvent
  - Liters of solution
  - Milliliters of solvent
  - None of the above
12. (8 pts) What would be the molarity of 25.6 g of  $\text{H}_2\text{SO}_4$  dissolved in enough water to make 250.0 mL of solution?
- 1.04 M
  - 15.3 M
  - 0.958 M
  - $1.04 \times 10^{-3}$  M
  - None of the above

SOLUTIONS TO THE TEST FOR  
**MODULE 9**

1. (4 pts) c. Acids turn red litmus paper blue.
2. (4 pts) b. H<sub>2</sub>O
3. (4 pts) d. H<sup>+</sup>
4. (4 pts) b. H<sub>2</sub>CO<sub>3</sub> only
5. (4 pts) False An indicator changes color as it goes from an acid to a base.
6. (4 pts) d. All of the above
7. (4 pts) e. None of the above 1 Mg<sup>2+</sup> ion and 2 OH<sup>-</sup> ions are present.
8. (4 pts) Amphiprotic substances can act as either acids or bases. Water is a common example.
9. (4 pts) c. Water + an ionic compound
10. (4 pts) d. moles/mL
11. (4 pts) c. Liters of solution
12. (8 pts) a. 1.04 M

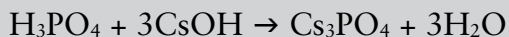
$$25.6 \text{ g H}_2\text{SO}_4 \times \frac{1 \text{ mol H}_2\text{SO}_4}{98.1 \text{ g}} = 0.261 \text{ moles H}_2\text{SO}_4$$



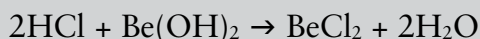
Convert volume to liters =  $250.0 \text{ mL} \times 1 \text{ L}/1000 \text{ mL} = 0.2500 \text{ L water}$

Concentration = # mol / # L =  $\frac{0.261 \text{ mol H}_2\text{SO}_4}{0.2500 \text{ L water}} = \underline{1.04 \text{ M}}$

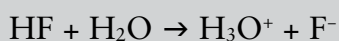
13. (4 pts: 2 pts for the correct substances, 2 pts for the equation being balanced)



14. (4 pts: 2 pts for the correct substances, 2 pts for the equation being balanced)



15. (4 pts: 2 pts for the correct substances, 2 pts for the equation being balanced)



16. (8 pts: 4 pts for both conversions, and 4 pts for the answer) **1.28 M**  
Concentration is number of moles divided by number of liters. Thus, first convert from grams to moles:

$$\frac{12.23 \text{ g-NH}_3}{1} \times \frac{1 \text{ mole NH}_3}{17.0 \text{ g-NH}_3} = 0.719 \text{ moles NH}_3$$

Now divide by the volume, making sure to convert 560.0 mL into 0.5600 L first:

$$\text{Concentration} = \frac{\# \text{ moles}}{\# \text{ liters}} = \frac{0.719 \text{ moles}}{0.5600 \text{ L}} = \underline{1.28 \text{ M}}$$

17. (4 pts) The chemist should take  $1.4 \times 10^2 \text{ mL}$  of the stock solution and dilute it to 450.0 mL. This is a dilution problem, so we use the dilution equation:

$$M_1V_1 = M_2V_2$$

$$(10.0 \text{ M}) \times V_1 = (3.0 \text{ M}) \times (450.0 \text{ mL})$$

$$V_1 = \frac{(3.0 \text{ M}) \times (450.0 \text{ mL})}{10.0 \text{ M}} = \underline{1.4 \times 10^2 \text{ mL}}$$

$$\frac{0.384 \text{ moles HCN}}{1} \times \frac{1 \text{ L}}{1.51 \text{ moles HCN}} = 0.254 \text{ L} = \underline{254 \text{ mL}}$$

4. We must first get our amount in moles.

$$\frac{500.0 \text{ g Cu}}{1} \times \frac{1 \text{ mole Cu}}{63.5 \text{ g Cu}} = 7.87 \text{ moles Cu}$$

Now that we have moles, we can do stoichiometry:

$$\frac{7.87 \text{ moles Cu}}{1} \times \frac{8 \text{ moles HNO}_3}{3 \text{ moles Cu}} = 21.0 \text{ moles HNO}_3$$

Of course, this is not quite the answer we need. We were asked to figure out how many mL of the HNO<sub>3</sub> solution is needed:

$$\frac{21.0 \text{ moles HNO}_3}{1} \times \frac{1 \text{ L}}{3.50 \text{ moles HNO}_3} = 6.00 \text{ L} = \underline{6.00 \times 10^3 \text{ mL}}$$

5. To calculate molality, we must have moles of solute and kg of solvent.

$$\frac{100.0 \text{ g Mg(NO}_3)_2}{1} \times \frac{1 \text{ mole Mg(NO}_3)_2}{148.3 \text{ g Mg(NO}_3)_2} = \underline{0.6743} \text{ moles Mg(NO}_3)_2$$

$$\frac{500.0 \text{ g}}{1} \times \frac{1 \text{ kg}}{1,000 \text{ g}} = 0.5000 \text{ kg}$$

Now that we have moles of solute and kg of solvent, we can use equation 10.1:

$$\text{molality} = \frac{\# \text{ moles solute}}{\# \text{ kg solvent}} = \frac{\underline{0.6743} \text{ moles Mg(NO}_3)_2}{0.5000 \text{ kg water}} = \underline{1.349 \text{ m}}$$

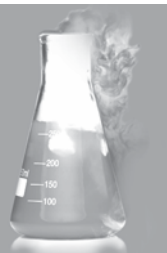
6. First, we need to see how many moles of CaCl<sub>2</sub> to add:

$$\text{molality} = \frac{\# \text{ moles solute}}{\# \text{ kg solvent}}$$

$$\frac{\text{moles CaCl}_2}{1.25 \text{ kg water}} = 2.0 \text{ m}$$

$$\text{moles CaCl}_2 = 2.5 \text{ moles}$$

# TEST FOR MODULE 10



1. (4 pts) When the maximum amount of solute that a solvent can hold is dissolved, chemists say that the solution is:
  - a. Supersaturated
  - b. Concentrated
  - c. Dissolving
  - d. Saturated
  - e. None of the above
2. (4 pts) If a solute's solubility in a liquid solvent is not affected by temperature or pressure, is the solute most likely a:
  - a. Solid
  - b. Liquid
  - c. Gas
  - d. All of the above
  - e. None of the above
3. (4 pts) If a solute's solubility in a liquid solvent decreases when the temperature increases, is the solute most likely a:
  - a. Solid
  - b. Liquid
  - c. Gas
  - d. All of the above
  - e. None of the above
4. (4 pts) A chemist is trying to dissolve a solid in water. If the chemist feels the beaker get hot while trying to make the solution, does the solute dissolve exothermically or endothermically?
  - a. Exothermically
  - b. Endothermically
5. (4 pts) What kind of solute usually increases in solubility with increasing temperature?
  - a. Solid
  - b. Liquid
  - c. Gas
  - d. All of the above
  - e. None of the above

11. (12 pts: 4 pts for each step) 670 mL  
This is just a stoichiometry problem:

$$\frac{350.0 \text{ g CuI}_2}{1} \times \frac{1 \text{ mole CuI}_2}{317.3 \text{ g CuI}_2} = 1.103 \text{ moles CuI}_2$$

$$\frac{1.103 \text{ moles CuI}_2}{1} \times \frac{2 \text{ moles NaI}}{1 \text{ mole CuI}_2} = 2.206 \text{ moles NaI}$$

$$\frac{2.206 \text{ moles NaI}}{1} \times \frac{1 \text{ L NaI}}{3.3 \text{ mole NaI}} = 0.67 \text{ L} = \underline{670 \text{ mL}}$$

12. (12 pts: 4 pts for each step) 0.67 g O<sub>2</sub>  
This is just a stoichiometry problem:

$$\frac{0.035 \text{ L H}_2\text{O}_2}{1} \times \frac{1.20 \text{ mole H}_2\text{O}_2}{1 \text{ L H}_2\text{O}_2} = 0.042 \text{ moles H}_2\text{O}_2$$

$$\frac{0.042 \text{ moles H}_2\text{O}_2}{1} \times \frac{1 \text{ mole O}_2}{2 \text{ moles H}_2\text{O}_2} = 0.021 \text{ moles O}_2$$

$$\frac{0.021 \text{ moles O}_2}{1} \times \frac{32.0 \text{ g O}_2}{1 \text{ mole O}_2} = \underline{0.67 \text{ g O}_2}$$

13. (4 pts) 2.66 m

$$m = \frac{\# \text{ moles solute}}{\# \text{ kg solvent}} = \frac{5.61 \text{ moles}}{2.11 \text{ kg}} = 2.66 \frac{\text{moles}}{\text{kg}} = \underline{2.66 \text{ m}}$$

14. (4 pts: 2 pts for # moles, and 2 pts for grams) 1.43 moles, 1.70 × 10<sup>2</sup> g KBr

$$m = \frac{\# \text{ moles solute}}{\# \text{ kg solvent}}$$

$$2.13 \text{ m} = \frac{\# \text{ moles}}{0.6700 \text{ kg}}$$

$$\# \text{ moles} = 1.43$$

$$\frac{1.43 \text{ moles KBr}}{1} \times \frac{119.0 \text{ g KBr}}{1 \text{ mole KBr}} = \underline{1.70 \times 10^2 \text{ g KBr}}$$

$$P_T = P_{\text{carbon dioxide}} + P_{\text{water vapor}}$$

$$790 \text{ torr} = P_{\text{carbon dioxide}} + 22.4 \text{ torr}$$

$$P_{\text{carbon dioxide}} = 790 \text{ torr} - 22.4 \text{ torr} = \underline{768 \text{ torr}}$$

Thus, only 768 torr of carbon dioxide was collected.

5. The mole fractions of SO<sub>2</sub>, NO, and SO<sub>3</sub> are 0.448, 0.502, and 0.051, respectively. Mole fraction is defined as the number of *moles* of component divided by the total number of moles. Right now, the problem gives us *grams*, not moles. Thus, we must first convert from grams to moles:

$$\frac{25.0 \text{ g SO}_2}{1} \times \frac{1 \text{ mole SO}_2}{64.1 \text{ g SO}_2} = 0.390 \text{ moles SO}_2$$

$$\frac{13.1 \text{ g NO}}{1} \times \frac{1 \text{ mole NO}}{30.0 \text{ g NO}} = 0.437 \text{ moles NO}$$

$$\frac{3.5 \text{ g SO}_3}{1} \times \frac{1 \text{ mole SO}_3}{80.1 \text{ g SO}_3} = 0.044 \text{ moles SO}_3$$

Now that we have the number of moles of each component, we can calculate the total number of moles in the mixture:

$$\text{Total number of moles} = 0.390 \text{ moles} + 0.437 \text{ moles} + 0.044 \text{ moles} = 0.871 \text{ moles}$$

Plugging that into equation 11.10:

$$X_{\text{SO}_2} = \frac{0.390 \text{ moles}}{0.871 \text{ moles}} = 0.448$$

$$X_{\text{NO}} = \frac{0.437 \text{ moles}}{0.871 \text{ moles}} = 0.502$$

$$X_{\text{SO}_3} = \frac{0.044 \text{ moles}}{0.871 \text{ moles}} = 0.051$$

The mole fractions of SO<sub>2</sub>, NO, and SO<sub>3</sub> are 0.448, 0.502, and 0.051, respectively.

6. The partial pressures of SO<sub>2</sub>, NO, and SO<sub>3</sub> are 0.54 atm, 0.60 atm, and 0.061 atm, respectively. Using the mole fractions we just obtained, this problem is an easy application of equation 11.11:

$$P_1 = X_1 \times P_T$$

$$P_{\text{SO}_2} = 0.448 \times 1.2 \text{ atm} = 0.54 \text{ atm}$$

$$P_{\text{NO}} = 0.502 \times 1.2 \text{ atm} = 0.60 \text{ atm}$$

$$P_{\text{SO}_3} = 0.051 \times 1.2 \text{ atm} = 0.061 \text{ atm}$$

The partial pressures of  $\text{SO}_2$ ,  $\text{NO}$ , and  $\text{SO}_3$  are 0.54 atm, 0.60 atm, and 0.061 atm, respectively.

7. In this problem, we are given the partial pressure of each gas. By Dalton's law, the total pressure is just the sum of the individual pressures:

$$P_T = 10.00 \text{ atm} + 4.00 \text{ atm} + 1.00 \text{ atm} = 15.00 \text{ atm}$$

By equation 11.11, then, we can calculate the mole fractions:

$$P_1 = X_1 \times P_T \quad X_1 = \frac{P_1}{P_T}$$

Plugging the numbers in for each gas:

$$X_{\text{N}_2} = \frac{10.00 \text{ atm}}{15.00 \text{ atm}} = 0.6667$$

$$X_{\text{O}_2} = \frac{4.00 \text{ atm}}{15.00 \text{ atm}} = 0.267$$

$$X_{\text{Ar}} = \frac{1.00 \text{ atm}}{15.00 \text{ atm}} = 0.0667$$

The mole fractions of  $\text{N}_2$ ,  $\text{O}_2$ , and  $\text{Ar}$  are 0.6667, 0.267, and 0.0667, respectively.

8. 11.1 liters In this problem, we are given pressure and temperature and the number of moles. We are then asked to calculate  $V$ . We can do this by rearranging the ideal gas law:

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

10.  $4.1 \times 10^3 \text{ L}$  In this stoichiometry problem, we are given the amount of limiting reactant and asked to calculate how much product will be made. We start by converting the amount of limiting reactant to moles:

$$\frac{2.89 \text{ kg}}{1} \times \frac{1000 \text{ g H}_2\text{O}_2}{1 \text{ kg}} \times \frac{1 \text{ mole H}_2\text{O}_2}{34.0 \text{ g H}_2\text{O}_2} = 85.0 \text{ moles H}_2\text{O}_2$$

We can then use stoichiometry to determine the number of moles of  $\text{H}_2\text{O}$  produced:

$$\frac{85.0 \text{ moles H}_2\text{O}_2}{1} \times \frac{8 \text{ moles H}_2\text{O}}{7 \text{ moles H}_2\text{O}_2} = 97.1 \text{ moles H}_2\text{O}$$

Now we need to use the ideal gas law, realizing that we must convert the temperature to Kelvin:

$$PV = nRT$$

$$V = \frac{nRT}{P} = \frac{97.1 \text{ moles} \times 0.0821 \frac{\text{L} \times \text{atm}}{\text{mole} \times \text{K}} \times 773.2 \text{ K}}{1.5 \text{ atm}} = 4.1 \times 10^3 \text{ L}$$

The volume produced is  $4.1 \times 10^3 \text{ L}$ .

## SAMPLE CALCULATIONS FOR EXPERIMENT 11.1

**Atmospheric Pressure:**

$$\frac{30.12 \text{ in Hg}}{1} \times \frac{2.54 \text{ cm}}{1.00 \text{ in}} \times \frac{10 \text{ mm}}{1 \text{ cm}} \times \frac{1.000 \text{ atm}}{760.0 \text{ mm Hg}} = 1.01 \text{ atm}$$

**Initial Mass of Butane Lighter:** 21.5 g

**Water Temperature:**  $23.1 \text{ }^\circ\text{C} + 273.15 = 296.3 \text{ K}$

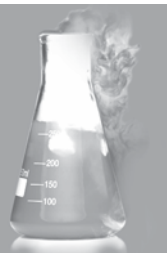
**Volume of Butane:** 200 mL = 0.200 L

**Final Mass of Butane Lighter:** 21.0 g

6. (4 pts) Following an established trend in the data even though no data are available for that region is called:
- Theorizing
  - Extrapolation
  - Estimating
  - Hypothesizing
  - None of the above
7. (4 pts) Which of the following are properties of an ideal gas?
- The molecules that make up the gas take up most of the volume.
  - The molecules are fairly close to one another.
  - The molecules are always losing energy with one another when they collide.
  - All of the above
  - None of the above
8. (4 pts) Which of the following conditions is the closest to STP?
- 273 K and 273 atm
  - 1.00 K and 1.00 atm
  - 1.00 K and 273 atm
  - 273 K and 1.00 atm
  - None of the above
9. (4 pts) Which of the following statements is true about vapor pressure?
- The vapor pressure for any liquid increases as the temperature increases.
  - The vapor pressure of water is always constant.
  - Vapor pressure is the pressure exerted by a gas on its container.
  - Vapor pressure only occurs when the liquid is boiling.
  - None of the above
10. (4 pts) What is the temperature at which the vapor pressure of a liquid is equal to normal atmospheric pressure?
- Standard temperature
  - Standard pressure
  - Melting point
  - Boiling point
  - None of the above
11. (4 pts) Which of the following is equal to the ideal gas constant?
- $R = 0.0821 \text{ L atm/mole } ^\circ\text{C}$
  - $R = 0.0821 \text{ cm}^3 \text{ atm/mole K}$
  - $R = 0.0821 \text{ L kPa /mole K}$
  - $R = 0.0821 \text{ L atm/mole K}$
  - None of the above



# SOLUTIONS TO THE TEST FOR MODULE II



1. (4 pts) d. Pressure
2. (4 pts) d. All of the above
3. (4 pts) d. All of the above
4. (4 pts) a. As long as pressure is constant, temperature and volume are linearly proportional to each other.
5. (4 pts) e. All of the above
6. (4 pts) b. Extrapolation
7. (4 pts) e. None of the above
8. (4 pts) d. 273 K and 1.00 atm
9. (4 pts) a. The vapor pressure for any liquid increases as the temperature increases.
10. (4 pts) d. Boiling point
11. (4 pts) d.  $R = 0.0821 \text{ L atm/mole K}$
12. (4 pts)  $7.0 \times 10^2 \text{ K}$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Now look at the significant figures in this problem. First, we subtracted 15 from 45. For that, we had to use the rule of addition and subtraction. Since both 45 and 15 have their last significant figure in the ones place, the answer must have its last significant figure in the ones place. Thus, the only way we could report that answer to indicate that the resulting zero in the ones place is significant is to use scientific notation. That's why the change in temperature is reported at  $3.0 \times 10^1$  °C. After that, the equation uses only multiplication, so at that point, we must count significant figures. The lowest number of significant figures in the equation is 2 ( $3.0 \times 10^1$  °C), so the answer can have only 2.

3.  $11.1 \frac{\text{J}}{\text{g}^\circ\text{C}}$  In this problem, we are given  $\Delta T$ , mass, and the heat absorbed, and we must calculate specific heat. So, first we rearrange equation 12.1 to solve for specific heat, and then we plug our numbers in:

$$c = \frac{q}{m \times \Delta T}$$

$$c = \frac{50.0 \text{ kJ}}{(124.1 \text{ g})(36.3 \text{ }^\circ\text{C})} = 0.0111 \frac{\text{kJ}}{\text{g}^\circ\text{C}}$$

You might wonder why we didn't convert kJ into J. In this problem, we were given no restrictions on units. We didn't need any units to cancel out, and the problem didn't specify what units to give specific heat in; thus, we didn't need to convert the energy unit. This unit is a perfectly acceptable unit for specific heat. If you did convert from kJ to J, your answer should be  $11.1 \frac{\text{J}}{\text{g}^\circ\text{C}}$ .

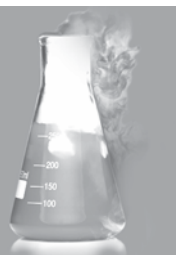
4.  $2.0 \times 10^1$  °C In order to get the copper's new temperature, we need to solve for  $\Delta T$  in equation 12.1. We can do this because we have the mass and heat given in the problem and the specific heat from table 12.1. Remember, though, since the copper lost heat, its  $q$  is negative! So first we rearrange the equation to solve for  $\Delta T$ :

$$\Delta T = \frac{q}{m \times c}$$

$$\Delta T = \frac{-456.7 \text{ J}}{(245 \text{ g}) \times \left(0.3851 \frac{\text{J}}{\text{g}^\circ\text{C}}\right)} = -4.84 \text{ }^\circ\text{C}$$

Now that we have  $\Delta T$ , we can rearrange equation 12.2 to solve for final temperature:

# SOLUTIONS TO QUARTER 3 TEST FOR MOD. 9-12



- (2 pts) a.  $C_2H_6O$   
Acids donate  $H^+$  ions. In this reaction,  $C_2H_6O$  becomes  $C_2H_5O^-$ . The only way that can happen is if it gives up an  $H^+$ . Thus,  $C_2H_6O$  is the acid.
- (2 pts) a. 2 ions of  $Al^{3+}$  and 3 ions of  $SO_4^{2-}$   
Looking at the chemical formula, you should immediately see the sulfate ion ( $SO_4^{2-}$ ). The positive ion is the aluminum ion. Aluminum's position on the periodic table tells you it's a  $3^+$  ion. Thus, the molecule splits into 2  $Al^{3+}$  ions and 3  $SO_4^{2-}$  ions. \_\_
- (4 pts: Give 2 pts partial credit for the unbalanced equation, and 2 pts for the balanced equation; or 4 pts for final answer.)  $2HNO_3 + Mg(OH)_2 \rightarrow 2H_2O + Mg(NO_3)_2$   
Magnesium hydroxide is  $Mg(OH)_2$ . Acids and bases usually react to give a salt and water. The salt is made up of the positive ion from the base ( $Mg^{2+}$ ) and the negative ion left over when the acid gets rid of its  $H^+$  ions. In this case, that will be  $NO_3^-$ . These 2 ions form  $Mg(NO_3)_2$ . The unbalanced equation, then, is:  

$$HNO_3 + Mg(OH)_2 \rightarrow 2H_2O + Mg(NO_3)_2$$

Now all we have to do is balance it:

$$2HNO_3 + Mg(OH)_2 \rightarrow 2H_2O + Mg(NO_3)_2$$
- (2 pts)  $HNO_3 + PH_3 \rightarrow PH_4^+ + NO_3^-$   
In this case, the base does not contain an hydroxide ion. Thus, a salt and water are not formed in this problem. Here, we just rely on the definition of acids and bases.  $PH_3$  will want to gain an  $H^+$  to become  $PH_4^+$ , and the nitric acid will want to give up its  $H^+$  ion to become  $NO_3^-$ .
- (3 pts) d. Decreased temperature and increased pressure
- (2 pts) a.  $Al(NO_3)_3$   
The freezing point depression depends on the solvent (which is water in each case),

$$\text{Concentration} = \frac{\# \text{ moles solute}}{\# \text{ liters solution}} = \frac{18.7 \text{ moles KOH}}{1.50 \text{ L}} = 12.5 \text{ M}$$

Now that we know the concentration of the stock solution, this is just a dilution problem:

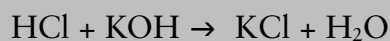
$$M_1V_1 = M_2V_2$$

So rearranging the equation gives:

$$V_1 = \frac{M_2V_2}{M_1} = \frac{0.10 \text{ M} \times 200.0 \text{ mL}}{12.5 \text{ M}} = 1.6 \text{ mL}$$

18. (6 pts–2 pts for getting moles of HCl, 2 pts for converting to moles of KOH, and 2 pts for getting the concentration of KOH) 0.74 M

Remember, titrations are just stoichiometry problems, so first we have to come up with a balanced chemical equation:



Since the endpoint was reached, we know that exactly enough acid was added to eat up all of the base. First, then, we calculate how many moles of acid were added:

$$\frac{0.0147 \text{ L}}{1} \times \frac{2.5 \text{ moles HCl}}{1 \text{ L}} = 0.037 \text{ moles HCl}$$

We can now use the chemical equation to determine how many moles of base were present:

$$\frac{0.037 \text{ moles HCl}}{1} \times \frac{1 \text{ mole KOH}}{1 \text{ mole HCl}} = 0.037 \text{ moles KOH}$$

Now that we have the number of moles of base present, we simply divide by the volume of acid to get concentration:

$$\text{Concentration} = \frac{\# \text{ moles solute}}{\# \text{ liters solution}} = \frac{0.037 \text{ moles KOH}}{0.050 \text{ L}} = \underline{0.74 \text{ M of KOH}}$$

19. (6 pts) c. 2.0 m

Freezing point depression is determined by the equation  $\Delta T = -iK_f m$ . We are already given 2 of the 4 variables in the equation ( $K_f, \Delta T$ ), and we can calculate a

third (i). Molality is the only unknown, so we can solve for it. Since  $\text{Na}_2\text{CO}_3$  is ionic, it will split up into ions (2 sodium ions and 1 carbonate ion), so  $i = 3$ :

$\Delta T = -iK_f m$  rearranges to solve for  $m$  (molal):

$$m = \frac{\Delta T}{-iK_f} = \frac{-11\text{ }^\circ\text{C}}{-3 \times 1.86\text{ }^\circ\text{C/molal}} = \underline{2.0\text{ molal of Na}_2\text{CO}_3}$$

20. (6 pts: 3 pts for conversion to moles, 3 pts for conversion to grams)

110 g  $\text{Na}_2\text{CO}_3$

Molality is defined as the number of moles of solute per kg of solvent. Thus, 2.0 m is 2.0 moles of  $\text{Na}_2\text{CO}_3$  per 1.00 kg of solvent. We don't have 1.00 kg; we have 0.5000 kg, so we first need to calculate how many moles are required:

$$\frac{2.0\text{ moles of Na}_2\text{CO}_3}{1.0\text{ kg water}} \times 0.5000\text{ kg water} = 1.0\text{ mole Na}_2\text{CO}_3$$

That's how many moles are required, but we need to know grams. That's an easy conversion:

$$\frac{1.0\text{ moles Na}_2\text{CO}_3}{1} \times \frac{106\text{ g Na}_2\text{CO}_3}{1\text{ mole Na}_2\text{CO}_3} = \underline{110\text{ g Na}_2\text{CO}_3}$$

**Note: 106 rounded to 2 significant figures equals 110.**

21. (6 pts: 3 pts for the molality, and 3 pts for the boiling point) 101.6 °C

To calculate boiling points, we must use equation 11.3. To do that, however, we must know  $i$  and  $m$ . To calculate  $m$ :

$$\frac{100.0\text{ g NaCl}}{1} \times \frac{1\text{ mole NaCl}}{58.5\text{ g NaCl}} = 1.71\text{ moles NaCl}$$

$$\text{molality} = \frac{\# \text{ moles solute}}{\# \text{ kg solvent}} = \frac{1.71\text{ moles NaCl}}{1.1000\text{ kg water}} = 1.55\text{ molal}$$

Since sodium chloride is an ionic compound, it dissolves by splitting up into 2 ions; thus,  $i = 2$ .

$$\Delta T = i K_b m = 2 \times 0.512\text{ }^\circ\text{C/m} \times 1.55\text{ m} = 1.59\text{ }^\circ\text{C}$$

This means that the boiling point of the solution is  $1.59\text{ }^\circ\text{C}$  *higher* than that of pure water. The boiling point of pure water is  $100.0\text{ }^\circ\text{C}$ , so the boiling point of this solution is 101.6 °C.

In this stoichiometry problem, we are given the amount of limiting reactant and asked to calculate how much product will be made. We start by converting the amount of limiting reactant to moles:

$$\frac{110.0 \text{ g H}_2\text{O}_2}{1} \times \frac{1 \text{ mole H}_2\text{O}_2}{34.0 \text{ g H}_2\text{O}_2} = 3.24 \text{ moles H}_2\text{O}_2$$

We can then use stoichiometry to determine the number of moles of H<sub>2</sub>O produced:

$$\frac{3.24 \text{ moles H}_2\text{O}_2}{1} \times \frac{8 \text{ moles H}_2\text{O}}{7 \text{ moles H}_2\text{O}_2} = 3.70 \text{ moles H}_2\text{O}$$

Now we need to use the ideal gas law solving for volume:

$$K = 341 \text{ }^\circ\text{C} + 273 = 614 \text{ K}$$

$$V = \frac{nRT}{P} = \frac{3.70 \text{ moles} \times 0.0821 \text{ L atm/mole K} \times 614 \text{ K}}{2.1 \text{ atm}} = 89 \text{ L}$$

Since the temperature increased, that means heat was released by the reaction, so when calculating  $\Delta H$ ,  $q$  must be made negative.

One molecule of NaOH has a mass of 40.0 amu, which means 1 mole NaOH = 40.0 g NaOH:

$$\text{Moles NaOH} = \frac{3.54 \text{ g NaOH}}{1} \times \frac{1 \text{ mole NaOH}}{40.0 \text{ g NaOH}} = 0.0885 \text{ moles NaOH}$$

$$\Delta H = \frac{-1,500 \text{ J}}{0.0885 \text{ moles}} = -17,000 \frac{\text{J}}{\text{mole}}$$

**NOTE:** Your answer might be quite different from ours. That's fine. It depends a lot on the brand of lye that you used.

Thus, the final rate equation is:

$$R = (230 \frac{1}{M^2 s}) \times [NO]^2 [Cl_2]$$

4. The rate equation will take on the form:

$$R = k[C_3H_6Br_2]^x [I^-]^y$$

To determine  $x$  and  $y$ , we look at trials where the concentration of one reactant stayed the same and the concentration of the other reactant changed. In trials 1 and 2, the concentration of  $I^-$  remained the same but the concentration of  $C_3H_6Br_2$  doubled. When that happened, the rate doubled. This means that  $x = 1$ . In trials 1 and 3, the  $C_3H_6Br_2$  concentration remained constant, but the  $I^-$  concentration doubled. When that happened, the rate doubled. This means  $y = 1$ . The rate equation, then, looks like:

$$R = k[C_3H_6Br_2][I^-]$$

To solve for  $k$ , we can use the data from any trial and plug it into our rate equation. We can then solve for  $k$ :

$$R = k[C_3H_6Br_2][I^-]$$

$$0.234 \frac{M}{s} = k \times (0.100 M) \times (0.200 M)$$

$$k = \frac{0.234 \frac{M}{s}}{(0.100 M)(0.200 M)} = 11.7 \frac{1}{M \times s}$$

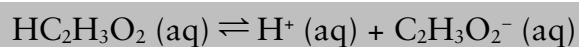
The overall rate equation, then is

$$R = (11.7 \frac{1}{M s}) \times [C_3H_6Br_2][I^-]$$

5. Since chemical reaction rate doubles for every  $10^\circ C$  increment, then to increase the rate of the reaction by a factor of 16, we just need to raise the temperature by 4 ten-degree increments. That way, we will multiply the old rate by  $2 \times 2 \times 2 \times 2$ , which equals 16. To increase the reaction rate by a factor of 16, we just raise the temperature by 40 degrees. Therefore, the new temperature should be  $25^\circ C + 40^\circ C = 65^\circ C$ .



- products. Thus, more  $\text{H}_2\text{CO}_3$  and  $\text{CaO}$  will be made and there will be less  $\text{CaCO}_3$ . Remember, solids and liquids are not sources of stress because their *concentrations* do not change. However, if we stress the equilibrium by varying the amount of another substance in the equation, the *amounts* of liquids and solids can change.
7. a. The reaction is exothermic, which means energy is a product. If the temperature is raised, the equilibrium shifts away from the side with the energy, so the  $\text{H}_2$  and  $\text{F}_2$  concentration will go down while the concentration of  $\text{HF}$  will go up.
    - b. When temperature is lowered, the reaction shifts toward the side with energy. Thus, the concentration of  $\text{HF}$  will lower and the concentrations of  $\text{H}_2$  and  $\text{F}_2$  will increase.
  8. a. This is an endothermic reaction, so energy is a reactant. When the temperature is raised, the reaction shifts away from the side with the energy, so the concentration of  $\text{H}_2$  and  $\text{N}_2$  will increase and the concentration of  $\text{NH}_3$  will decrease.
    - b. When pressure is raised, the reaction shifts away from the side with the most gas molecules. There are 4 gas molecules on the **products** side and only 2 on the **reactants** side. Thus, the reaction will shift away from reactants, causing the concentrations of  $\text{N}_2$  and  $\text{H}_2$  to decrease while the concentration of  $\text{NH}_3$  to rise.
    - c. When pressure is lowered, the reaction shifts toward the side with more gas molecules, making the concentrations of  $\text{N}_2$  and  $\text{H}_2$  increase and  $\text{NH}_3$  decrease.
  9. The ionization constant is simply the equilibrium constant for the acid ionization reaction. In order to determine the ionization reaction, you simply take the acid in its aqueous phase and remove an  $\text{H}^+$ . When we remove an  $\text{H}^+$  ion from  $\text{HC}_2\text{H}_3\text{O}_2$ , we are left with  $\text{C}_2\text{H}_3\text{O}_2^-$ . In the end, then, the aqueous acid is the reactant, and the  $\text{H}^+$  ion and  $\text{C}_2\text{H}_3\text{O}_2^-$  (both in aqueous phase) will be the products:



The equilibrium constant for this reaction is the ionization constant,  $K_a$ :

$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

## *Exploring Creation with Chemistry, 3<sup>rd</sup> Edition – Errata File*

This file contains the corrections for the 4<sup>th</sup> Printing: April 2018, of the **Solutions and Tests Manual**. The printing for the Solutions and Tests Manual may not be the same as for the Textbook and Student Notebook. Corrections for the Textbook and Student Notebook are in separate files. (Updated July 2024)

Additional corrections added 8/16/2023

### Test Corrections for the Test Pages:

Module 8 Test – see page 145 below

Module 9 Test question #7 answer c. Change  $\text{Mg}^{+2}$  to  $\text{Mg}^{2+}$  and answer d. change  $\text{OH}^{-2}$  to  $\text{OH}^{2-}$ .

Quarterly Test 4 – see page 324 below

### Corrections for the Solutions Manual:

Page 108 – Solution to Module 6 Extra Practice Problem #8 – On the products side, change  $2\text{NaBr (s)}$  to  $2\text{NaBr (aq)}$ .

Page 130 – Module 7 Test Answer for #10 – in the 2<sup>nd</sup> gray box, change 10 g Li to **10.0** g Li in the first numerator.

Page 142 – Solution to Module 8 Extra Practice Problem #8 – in the 2<sup>nd</sup> line, change 100.0 g to **200.0 g** two times.

Page 145 – Module 8 Test Question #11 – change 6.4 g of O to 6.4 g **O<sub>2</sub>**. Change this on the Test Pages, page 34 as well.

Page 256 – Solution to Module 13 Practice Problem #1 – in the 2<sup>nd</sup> paragraph of explanation, change the last sentence to: Equation **13.2** becomes:

Page 324 – Quarterly Test 4 – Question #24 – On the products side of the given equation, change  $2\text{VO}^{2+}(\text{aq})$  to  **$2\text{VO}^{2+}(\text{aq})$** . The plus sign should be a superscript. Change this on the Test Pages, page 83 as well.

Additional Edit added 7/26/2024

Page 15 – Test #1, answer to Question #6, gray box: change 300 mg to **30.0** mg

Page 85 – In the answer to Module 5 Practice Problem #7 at the top of the page, underline the word '**bent**' along with the rest of the underlined answer.

Page 174 – Module 9 Extra Practice Problem #11, 3<sup>rd</sup> gray box, final answer: add a 0 to the end of the number so it reads “**0.000880**” which matches the next gray box.

Page 200 – In the answer to Module 11 Practice Problem #3, last gray box, add **P<sub>2</sub> =** before the answer.

Page 251 – Quarterly Test 3, question # 21 – in the written explanation, first sentence, change ‘...use equation 11.3.’ to ‘...use equation **10.3.**’