

Chemistry Chapter 8 – Ionic Reactions in Solutions

1) An acidified solution of a sample of pure oxalic acid dihydrate,  $\text{H}_2\text{C}_2\text{O}_4 \cdot 2 \text{H}_2\text{O}$ , was used to standardize a solution of potassium permanganate. In this process, the oxalic acid was oxidized to carbon dioxide and the permanganate ion was reduced to manganese(II) ion. Use the following data to calculate the molarity (formality) of the permanganate solution:

Weight of flask and acid crystals	68.6153 grams
Weight of flask	68.4262 grams
Burette reading after adding $\text{KMnO}_4$ solution	31.52 mL
Burette reading before adding $\text{KMnO}_4$ solution	1.51 mL

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2. (a) What is the pH of a 2.0 molar solution of acetic acid.  $K_a$  of acetic acid =  $1.8 \times 10^{-5}$

(b) A buffer solution is prepared by adding 0.10 liter of 2.0 molar acetic acid solution to 0.1 liter of a 1.0 molar sodium hydroxide solution. Compute the hydrogen ion concentration of the buffer solution.

(c) Suppose that 0.010 liter of 0.50 molar hydrochloric acid is added to 0.040 liter of the buffer prepared in (b). Compute the hydrogen ion concentration of the resulting solution.

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3. An excess of  $\text{AgNO}_3$  solution was added to a solution that contained 1.65 grams of a chloride salt of an unknown element M. The  $\text{AgCl}$  precipitate was filtered, dried, and found to weigh 2.49 grams. Further experiments showed that the ion of element M had a +3 oxidation state in the chloride salt.

(a) Write a balanced equation for the reaction of  $\text{AgNO}_3$  with  $\text{MCl}_3$ .

(b) Calculate the percent of chlorine in  $\text{MCl}_3$ .

(c) Calculate the atomic weight of the unknown element M.

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4. Reactions requiring either an extremely strong acid or an extremely strong base are carried out in solvents other than water. Explain why this is necessary for both cases.

5. Briefly outline a laboratory procedure that can be used to determine the composition of an alloy of copper and silver. The alloy dissolves completely in concentrated nitric acid.

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6. 8) Account for the difference in solubility described in each of the following experimental observations:

(a)  $\text{BaCO}_3$ ,  $\text{BaSO}_3$ , and  $\text{BaSO}_4$  are only slightly soluble in water, but the first two dissolve in  $\text{HCl}$  solution whereas  $\text{BaSO}_4$  does not.

(b)  $\text{CuS}$  cannot be dissolved by warm, dilute  $\text{HCl}$  but it does dissolve in warm, dilute  $\text{HNO}_3$

(c)  $\text{AgCl}$ ,  $\text{Hg}_2\text{Cl}_2$ , and  $\text{PbCl}_2$  are only slightly soluble in water, but  $\text{AgCl}$  does dissolve in ammonia solution whereas the other two do not.

(d)  $\text{Fe}(\text{OH})_3$  and  $\text{Al}(\text{OH})_3$  are only slightly soluble in water, but  $\text{Al}(\text{OH})_3$  dissolves in concentrated  $\text{NaOH}$  whereas  $\text{Fe}(\text{OH})_3$  does not.

ANSWER: - Actually, I'm hoping that we get this one, cuz there's no calculations

a) 3 points

Recognition that dissolution is favored by shift to right of equilibrium,  
 $\text{BaX(s)} \rightleftharpoons \text{Ba}^{2+} + \text{X}^{2-}$

Recognition that conversions,  $\text{CO}_3^{2-} \rightarrow \text{HCO}_3^-$  and  $\text{SO}_3^{2-} \rightarrow \text{HSO}_3^-$  caused by  $\text{H}^+$ , result in greater solubility of the carbonate and sulfite

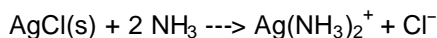
Recognition that  $\text{H}^+$  does not react appreciably with  $\text{SO}_4^{2-}$ , so as to increase the solubility of the sulfate, since  $\text{SO}_4^{2-}$  is a poor base (or equivalent)

---

b) 1 point

The sulfide ion is oxidized by warm nitric acid, shifting  $\text{CuS(s)} \rightleftharpoons \text{Cu}^{2+} + \text{S}^{2-}$  to the right.

c) 2 points

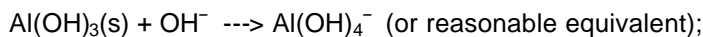


Recognition that  $\text{NH}_3$  complexes Ag

Recognition of  $\text{Ag(NH}_3)_2^+$  complex and that  $\text{Hg}_2^{2+}$  and  $\text{Pb}^{2+}$  do not form soluble ammine complexes

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d) 2 points



$\text{Fe(OH)}_3$  does not react in this way.

Recognition of amphoterism of  $\text{Al(OH)}_3$

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**9) An experiment is performed to determine the empirical formula of a copper iodide formed by direct combination of elements. A clean strip of copper metal is weighed accurately. It is suspended in a test tube containing iodine vapor generated by heating solid iodine, A white compound forms on the strip of copper, coating it uniformly. The strip with the adhering compound is weighed. Finally, the compound is washed completely from the surface of the metal and the clean strip is dried and reweighed.**

DATA TABLE	
Mass of clean copper strip	1.2789 grams
Mass of copper strip and compound	1.2874 grams
Mass of copper strip after washing	1.2748 grams

(a) State how you would use the data above to determine each of the following. (Calculations not required.)

- (1) The number of moles of iodine that reacted
- (2) The number of moles of copper that reacted

(b) Explain how you would determine the empirical formula for the copper iodide.

**(c) Explain how each of the following would affect the empirical formula that could be calculated.**

- (1) Some unreacted iodine condensed on the strip.**
- (2) A small amount of the white compound flaked off before weighing.**

ANSWER:

a) four points

1. mass of Cu strip and compound minus mass of original clean Cu strip = mass of iodine  
mass of iodine / atomic mass of iodine (126.9 g/mol) = moles iodine
2. mass of original clean Cu strip minus mass of strip after washing = mass of Cu  
mass of copper / atomic mass of copper (63.54 g/mol) = moles of Cu

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b) one point

The empirical formula is the ratio of mol. iodine to mol. copper ( or alternatively, mol. copper to mol. iodine)

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c) three points

1. Unreacted iodine would make the apparent mass of compound and the iodine too high. Thus, the I:Cu ratio in the empirical formula would be too high.
2. If some of the compound flaked off, the mass of compound (and the iodine) would be too low. Thus, the I:Cu ratio in the empirical formula would be too low.

In part (c) 2 points were awarded for a correct response on either part (1) or part (2) that included the prediction (1 pt.) and reason (1 pt.) An additional one point was awarded for the correct prediction and reason in the other part.

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18) Describe at least three methods of separating  $\text{Cu}^{2+}$  and  $\text{Al}^{3+}$  in aqueous solution and explain the chemical principles involved in each separation. Where appropriate, summarize your descriptions by means of equations.

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**18. Define oxidation and reduction in terms of electron transfer. Illustrate your answer by reference to the reactions which take place between the following pairs of substances in aqueous solutions:**

- (a) Iodine and hydrogen sulfide**
- (b) Bromine and iron(II) sulfate**

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Water is added to 4.267 grams of  $\text{UF}_6$ . The only products are 3.730 grams of a solid containing only uranium, oxygen and fluorine and 0.970 gram of a gas. The gas is 95.0% fluorine, and the remainder is hydrogen.

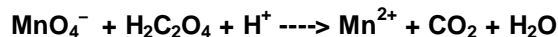
- (a) From these data, determine the empirical formula of the gas.
- (b) What fraction of the fluorine of the original compound is in the solid and what fraction in the gas after the reaction?
- (c) What is the formula of the solid product?

(d) Write a balanced equation for the reaction between  $\text{UF}_6$  and  $\text{H}_2\text{O}$ . Assume that the empirical formula of the gas is the true formula.

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A 1.2156-gram sample of a mixture of  $\text{CaCO}_3$  and  $\text{Na}_2\text{SO}_4$  was analyzed by dissolving the sample and completely precipitating the  $\text{Ca}^{2+}$  as  $\text{CaC}_2\text{O}_4$ . The  $\text{CaC}_2\text{O}_4$  was dissolved in sulfuric acid and the resulting  $\text{H}_2\text{C}_2\text{O}_4$  was titrated with a standard  $\text{KMnO}_4$  solution.

(a) On a page of your answer booklet, write the balanced equation for the titration reaction, shown balanced below.



Indicate which substance is the oxidizing agent and which substance is the reducing agent.

(b) The titration of the  $\text{H}_2\text{C}_2\text{O}_4$  obtained required 35.62 milliliters of 0.1092-molar  $\text{MnO}_4^-$  solution. Calculate the number of moles of  $\text{H}_2\text{C}_2\text{O}_4$  that reacted with the  $\text{MnO}_4^-$ .

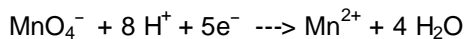
(c) Calculate the number of moles of  $\text{CaCO}_3$  in the original sample.

(d) Calculate the percentage by weight of  $\text{CaCO}_3$  in the original sample.

*I will ask on Monday if the AP question requires a calculator. If it does, THIS IS AN EXCELLENT QUESTION. MEMORIZE*

ANSWER:

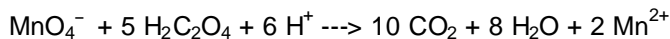
a) 5 points



or  $\text{MnO}_4^- (+7) \rightarrow \text{Mn}^{2+}$  (with an indication of 5  $\text{e}^-$  exchanged)



or  $\text{H}_2\text{C}_2\text{O}_4 (+3) \rightarrow 2 \text{CO}_2 (+4)$  (with an indication of 2  $\text{e}^-$  exchanged)



$\text{MnO}_4^-$  (or Mn) oxidizing agent

$\text{H}_2\text{C}_2\text{O}_4$  (or C) reducing agent

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b) 4 points

$$\text{moles MnO}_4^- = 0.03563 \text{ L} \times 0.1092 \text{ mol/L} = 3.890 \times 10^{-3} \text{ mol}$$

$$\text{moles H}_2\text{C}_2\text{O}_4 = 3.890 \times 10^{-3} \text{ mol MnO}_4^- \times (5 \text{ mol H}_2\text{C}_2\text{O}_4 / 2 \text{ mol MnO}_4^-) = 9.724 \times 10^{-3} \text{ mol}$$

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c) 2 points

# mol  $\text{CaCO}_3 = \# \text{ mol H}_2\text{C}_2\text{O}_4$

therefore  $9.72 \times 10^{-3}$  mol  $\text{CaCO}_3$  in sample

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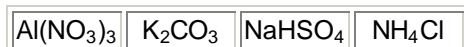
d) 4 points

$\text{g CaCO}_3 = 9.72 \times 10^{-3} \text{CaCO}_3 \times (100.1 \text{ g CaCO}_3 / \text{mol CaCO}_3) = 0.973 \text{ g CaCO}_3$

$\% \text{CaCO}_3 = (0.973 \text{ g CaCO}_3 / 1.2516 \text{ g sample}) \times 100 = 77.7\%$

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7)



(a) Predict whether a 0.10-molar solution of each of the salts above is acidic, neutral, or basic.

(b) For each of the solutions that is not neutral, write a balanced chemical equation for a reaction occurring with water that supports your prediction.

ANSWER:

7)

a) 4 points

a solution of  $\text{Al}(\text{NO}_3)_3$  is acidic

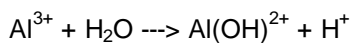
a solution of  $\text{K}_2\text{CO}_3$  is basic

a solution of  $\text{NaHSO}_4^-$  is acidic

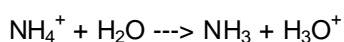
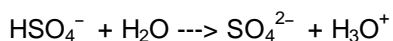
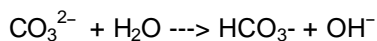
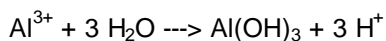
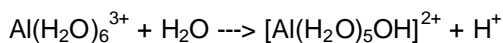
a solution of  $\text{NH}_4\text{Cl}$  is acidic

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b) 4 points



or



Equations needed to be balanced and ions must show correct charges. Molecular formulas were acceptable if acids and bases formed were marked strong and weak.

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(a) Describe what you would see if you added

1. a piece of zinc metal to a test tube that contains 6-molar hydrochloric acid;
2. a piece of copper metal to another test tube that contains 6-molar hydrochloric acid.

(b) Write balanced equations for any reactions that occur.

(c) Explain how you could use the table of standard reduction potentials on page 4 to predict the observed results.

(d) In a separate experiment, concentrated nitric acid is added to a test tube containing a piece of copper metal.

1. Describe what you would see.
2. Explain any differences between the results obtained in this experiment and those obtained with copper metal in part (a).

ANSWER:

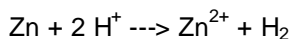
a) two points

1. Bubbling or dissolving of Zn

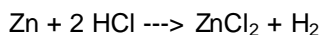
2. No reaction

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b) two points

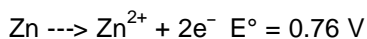
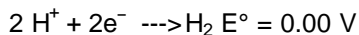


OR



Note: these last two answers are not off the official AP scoring standards, however the ChemTeam believes they are very representative of how the standards might look.

c) two points



The  $E^\circ$  of the overall reaction is positive, demonstrating that the forward reaction is spontaneous.

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d) two points

- 1) The Cu dissolves with the evolution of a brown gas.
- 2) The Cu was non-reactive toward the HCl and reacted spontaneously with the HNO<sub>3</sub>
- 

6) A 0.500-gram sample of a weak, nonvolatile acid, HA, was dissolved in sufficient water to make 50.0 milliliters of solution. The solution was then titrated with a standard NaOH solution. Predict how the calculated molar mass of HA would be affected (too high, too low, or not affected) by the following laboratory procedures. Explain each of your answers.

(a) After rinsing the buret with distilled water, the buret is filled with the standard NaOH solution; the weak acid HA is titrated to its equivalence point.

(b) Extra water is added to the 0.500-gram sample of HA.

(c) An indicator that changes color at pH 5 is used to signal the equivalence point.

(d) An air bubble passes unnoticed through the tip of the buret during the titration.

ANSWER:

(a) two points

Calculated  $M_m(\text{HA})$  too low

$M(\text{NaOH}) \Rightarrow V(\text{NaOH}) \Rightarrow n(\text{NaOH}) \Rightarrow n(\text{HA}) \Rightarrow M_m(\text{HA})$

$(M = n \div V)$  and  $(M_m = m \div n)$

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(b) two points

Calculated  $M_m(\text{HA})$  not affected

Any one of the following reasons. Water:

does not change  $n(\text{HA})$   
changes only  $M(\text{HA})$  -- sense of dilution  
is not a reactant

---

(c) two points

Calculated  $M_m(\text{HA})$  too high

**equivalence point**  $\Rightarrow n(\text{NaOH}) \Rightarrow n(\text{HA}) \Rightarrow M_m(\text{HA})$   
**(expected pH higher)**

Note: "no effect if NaOH standardized with same indicator" earns 2 points; no credit earned if pH=7 or neutral

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(d) two points

Calculated  $M_m(\text{HA})$  too low

$V(\text{NaOH}) \Rightarrow n(\text{NaOH}) \Rightarrow n(\text{HA}) \Rightarrow M_m(\text{HA})$

Note: point earned for  $V(\text{NaOH})$  only if:

- (i) no explanation point is earned in (a)
  - (ii) explanation in (a) also includes  $V(\text{NaOH})$
- 

2) a) A student studied the reactions which occur when magnesium metal is heated in air as follows. A sample of granulated magnesium weighing 20.0 grams was heated for some time in a current of air passing over a loosely covered crucible. The sample was then allowed to cool to room temperature. The powdery gray material was then boiled with excess water, and the ammonia gas given off was caught in  $2.00 \times 10^2$  milliliters of 1.00-molar sulfuric acid. Analysis showed that 83.3 milliliters of the acid had been neutralized by the ammonia to form  $(\text{NH}_4)_2\text{SO}_4$ . Calculate the weight of magnesium in the original sample which had been converted to magnesium nitride.

b) The residue remaining after the treatment with hot water, described in (a), was dissolved in excess 3.00-molar hydrochloric acid and was found to displace  $2.80 \times 10^3$  milliliters of hydrogen gas, measured over water at  $26^\circ\text{C}$  (v.p. = 25.2 mm Hg) and at a barometric pressure of 757 milliliters Hg. Calculate the weight of metallic magnesium remaining from the original sample after it had been heated in air.

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18. For the titration of 0.1-molar acetic acid with 0.1-molar sodium hydroxide, what factors need to be considered in selecting an indicator?

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**A 5.00 gram sample of a dry mixture of potassium hydroxide, potassium carbonate, and potassium chloride is reacted with 0.100 liter of 2.0 molar HCl solution.**

**(a) A 249 milliliter sample of dry  $\text{CO}_2$  gas, measured at  $22^\circ\text{C}$  and 740 torr, is obtained from the reaction. What is the percentage of potassium carbonate in the mixture?**

**(b) The excess HCl is found by titration to be chemically equivalent to 86.6 milliliters of 1.50 molar NaOH. Calculate the percentages of potassium hydroxide and of potassium chloride in the original mixture.**

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7)

**A 10.00 milliliter sample of  $\text{NH}_3$  solution is titrated with a standard HCl solution.**

- 1) An unknown volume of water is added to the HCl solution.
- 2) An unknown volume of water is added to the 10.00 milliliter sample of  $\text{NH}_3$  solution.
- 3) Phenolphthalein is used as the indicator.

**For each of these three steps taken during the titration:**

**(a) State whether it introduces an error into the titration results.**

**(b) For any of the steps that introduce(s) an error, state whether the titration result will be raised or lowered compared to the result obtained if the error had not been made.**

**(c) Explain why the result is high or low for each error that you detect.**

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3) The percentage by weight of nitric acid,  $\text{HNO}_3$ , in a sample of concentrated nitric acid is to be determined.

(a) Initially, a NaOH solution was standardized by titration with a sample of potassium hydrogen phthalate,  $\text{KHC}_8\text{H}_4\text{O}_4$ , a monoprotic acid often used as a primary standard. A sample of pure  $\text{KHC}_8\text{H}_4\text{O}_4$  weighing 1.518 grams was dissolved in water and titrated with NaOH solution. To reach the equivalence point, 26.90 milliliters of base required. Calculate the molarity of the NaOH solution. (Molecular weight:  $\text{KHC}_8\text{H}_4\text{O}_4 = 204.2$ )

(b) A 10.00-milliliter sample of the concentrated nitric acid was diluted with water to a total volume of 500.0 milliliters. Then 25.00 milliliters of the diluted acid solution was titrated with the standardized NaOH prepared in part (a). The equivalence point was reached after 28.35 milliliters of the base had been added. Calculate the molarity of the concentrated nitric acid.

(c) The density of the concentrated nitric used in this experiment was determined to be 1.42 grams per milliliter. Determine percentage by weight of  $\text{HNO}_3$  in the original sample of concentrated nitric acid.

**ANOTHER VERY VERY GOOD QUESTION!!!!**

ANSWER:

a) three points

$$1.518 \text{ g KHC}_8\text{H}_4\text{O}_4 \div (204.2 \text{ g / mole}) = 7.434 \times 10^{-3} \text{ moles KHP}$$

$$7.434 \times 10^{-3} \text{ mol KHP} = 7.434 \times 10^{-3} \text{ mol NaOH}$$

$$7.434 \text{ millimole NaOH} \div 26.90 \text{ mL NaOH} = 0.2764 \text{ M NaOH}$$

OR

$$7.434 \times 10^{-3} \text{ mol NaOH} / 0.02690 \text{ L} = 0.2764 \text{ M NaOH}$$

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b) three points

$$(28.35 \text{ mL NaOH} / 25.00 \text{ mL HNO}_3) \times (0.2764 \text{ moles NaOH} / \text{L}) \times (1 \text{ mole HNO}_3 / 1 \text{ mole NaOH}) = 0.3134 \text{ M HNO}_3 \text{ solution}$$

$$0.3134 \text{ M} \times 500 \text{ ml dilute soln} = x (10.00 \text{ mL conc. HNO}_3 \text{ soln})$$

$$x = (0.3134) (500) \div (10.00) = 15.67 \text{ M HNO}_3$$

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c) three points

$$\% = (\text{g HNO}_3 \div \text{g soln}) \times 100$$

to get numerator:

$$(15.67 \text{ HNO}_3 / \text{L soln}) \times (63.02 \text{ g HNO}_3 / \text{mole}) = 987.5 \text{ g HNO}_3 \text{ (in a L)}$$

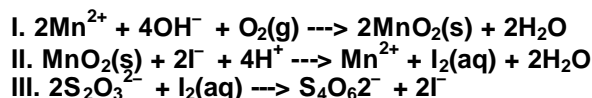
to get denominator:

$$(1.42 \text{ g soln} / \text{mL soln}) \times (1000 \text{ mL} / \text{L}) = 1,420 \text{ g HNO}_3 \text{ soln (in a L)}$$

then:

$$\% = (987.5 \text{ g HNO}_3 / 1,420 \text{ g soln}) \times 100 = 69.5\% \text{ HNO}_3$$

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The amount of oxygen,  $\text{O}_2$ , dissolved in water can be determined by titration. First,  $\text{MnSO}_4$  and  $\text{NaOH}$  are added to a sample of water to convert all of the dissolved  $\text{O}_2$  to  $\text{MnO}_2$ , as shown in equation I above. Then  $\text{H}_2\text{SO}_4$  and  $\text{KI}$  are added and the reaction represented by equation II proceeds. Finally, the  $\text{I}_2$  that is formed is titrated with standard sodium thiosulfate,  $\text{Na}_2\text{S}_2\text{O}_3$ , according to equation III.

(a) According to the equations above, how many moles of  $\text{S}_2\text{O}_3^{2-}$  are required for analyzing 1.00 mole of  $\text{O}_2$  dissolved in water?

(b) A student found that a 50.0 milliliter sample of water required 4.86 milliliters of 0.0112 molar  $\text{Na}_2\text{S}_2\text{O}_3$  to reach the equivalence point. Calculate the number of moles of  $\text{O}_2$  dissolved in this sample.

(c) How would the results in (b) be affected if some  $\text{I}_2$  were lost before the  $\text{S}_2\text{O}_3^{2-}$  was added? Explain.

(d) Name an appropriate indicator for the reaction shown in equation III and describe the change you would observe at the end point of the titration.

ANSWER

3) average = three points

a) one point

$$(1.00 \text{ mol O}_2) (2 \text{ mol MnO}_2 / 1 \text{ mol O}_2) (1 \text{ mol I}_2 / 1 \text{ mol MnO}_2) (2 \text{ mol S}_2\text{O}_3^{2-} / 1 \text{ mol I}_2) = 4 \text{ mol S}_2\text{O}_3^{2-}$$

Note: answer only is sufficient.

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b) two points

$$\text{mol S}_2\text{O}_3^{2-} = (0.00486) (0.0112) = 5.44 \times 10^{-5} \text{ mol S}_2\text{O}_3^{2-}$$

$$\text{mol O}_2 = 5.44 \times 10^{-5} / 4 = 1.36 \times 10^{-5} \text{ mol}$$

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c) one point

less  $\text{I}_2$  therefore less  $\text{S}_2\text{O}_3^{2-}$  required therefore lower amount of  $\text{O}_2$  (both direction and reason required)

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d) three points (one for M; one for correct use of R; one for correct T)

$$M_{\text{soln}} \text{ in (b)} = 1.36 \times 10^{-5} \text{ mol} / 0.050 \text{ L} = 2.72 \times 10^{-4} \text{ M}$$

$$V = (nRT) / P = [(2.72 \times 10^{-4}) (0.0821) (298)] / 1$$

=  $6.65 \times 10^{-3}$  L or 6.65 mL

---

e) two points

starch indicator

color disappears or blue disappears (violet or purple OK)

color change alone is not sufficient for 2nd pt

any other color with starch is not sufficient for 2nd pt

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**5) Discuss the roles of indicators in the titration of acids and bases. Explain the basis of their operation and the factors to be considered in selecting an appropriate indicator for a particular titration.**

ANSWER:

Points awarded are indicated at the end of each statement.

An indicator signals the endpoint of a titration by changing color. (2)

An indicator is a weak acid or a weak base where the acid form and the basic form of the indicator are different colors. (2)

An indicator usually changes color when the concentrations of the acid form and the basic form are about equal; that is, when the pH is near the pK for the indicator. (1)

When an indicator is selected, the pH at which the indicator changes color should bracket the pH of the titration solution at the equivalence point. (2)

For example, when a strong acid is titrated with a strong base, the pH at the equivalence point is 7, so an indicator that changes color at a pH of 7 should be used. (1)

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**3) The molecular weight of a monoprotic acid HX was to be determined. A sample of 15.126 grams of HX was dissolved in distilled water and the volume brought to exactly 250.00 milliliters in a volumetric flask. Several 50.00-milliliter portions of this solution were titrated against NaOH solution, requiring an average of 38.21 milliliters of NaOH.**

**The NaOH solution was standardized against oxalic acid dihydrate,  $\text{H}_2\text{C}_2\text{O}_4 \cdot 2 \text{H}_2\text{O}$  (molecular weight: 126.066  $\text{g mol}^{-1}$ ). The volume of NaOH solution required to neutralize 1.2596 grams of oxalic acid dihydrate was 41.24 milliliters.**

**(a) Calculate the molarity of the NaOH solution.**

**(b) Calculate the number of moles of HX in a 50.00-milliliter portion used for titration.**

**(c) Calculate the molecular weight of HX.**

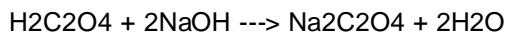
**(d) Discuss the effect on the calculated molecular weight of HX if the sample of oxalic acid dihydrate contained a nonacidic impurity.**

ANSWER:

a) three points

$$\text{mol H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O} = 1.2596 \text{ g H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O} / 126.066 \text{ g/mol}$$

$$= 9.9916 \times 10^{-3} \text{ mol}$$



$$\text{mol NaOH} = 9.9916 \times 10^{-3} \text{ mol H}_2\text{C}_2\text{O}_4 \times (2 \text{ mol NaOH} / 1 \text{ mol H}_2\text{C}_2\text{O}_4) = 1.9983 \times 10^{-2}$$

$$M_{\text{NaOH}} = 0.019983 \text{ mol} / 0.04124 \text{ L} = 0.4846$$

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b) two points

$$\text{mol HX} = \text{mol NaOH}$$

$$0.03821 \text{ L} \times 0.4846 \text{ mol/L} = 0.01852 \text{ mol HX}$$

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c) two points as follows: a logical argument plus a quantitative comparison = 2; a logical argument but no quantitative comparison = 1; one error in the quantitative comparison = 1

$$(0.01852 \text{ mol HX} / 50.00 \text{ mL}) \times 250.00 \text{ mL} = 0.09260 \text{ mol}$$

$$MW = (15.126 \text{ g} \times 0.09260 \text{ mol}) = 163.3 \text{ g/mol}$$

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d) two points

The calculated molecular weight is smaller than true value, because of the following:

- Measured g H<sub>2</sub>C<sub>2</sub>O<sub>4</sub> is larger than true value.
- Calculated mol H<sub>2</sub>C<sub>2</sub>O<sub>4</sub> is larger than true value.
- Calculated mol NaOH is larger than true value.
- Calculated M NaOH is larger than true value.
- Calculated mol HX is larger than true value.

Therefore,

$$MW = g_{\text{HX}} (\text{true value}) / \text{mol HX} (\text{calculated, and too large})$$

One point was subtracted for an arithmetic error(s) or for a serious significant figure error.

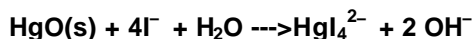
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**3) A solution of hydrochloric acid has a density of 1.15 grams per mL and is 30.% by weight HCl.**

**(a) What is the molarity of this solution of HCl?**

(b) What volume of this solution should be taken in order to prepare 5.0 liters of 0.20-molar hydrochloric acid by dilution with water?

(c) In order to obtain a precise concentration, the 0.20-molar hydrochloric acid is standardized against pure HgO (molecular weight = 216.59) by titrating the OH<sup>-</sup> produced according to the following quantitative reaction:



In a typical experiment, 0.7147 gram of HgO required 31.67 milliliters of the hydrochloric acid solution for titration. Based on these data, what is the molarity of the HCl solution expressed to four significant figures?

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6) Predict whether solutions of each of the following salts are acidic, basic, or neutral. Explain your prediction in each case.

(a) Al(NO<sub>3</sub>)<sub>3</sub>

(b) K<sub>2</sub>CO<sub>3</sub>

(c) NaBr

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1) Qualitative analysis revealed that an unknown contained the anions, NO<sub>3</sub><sup>-</sup>, Cl<sup>-</sup>, and ClO<sub>3</sub><sup>-</sup>. A student weighed a sample into a beaker and recorded the following data:

Weight of beaker and sample	68.962 grams
Weight of beaker	67.625 grams

The sample was dissolved and diluted to 200.0 milliliters. A 50.00-milliliter portion of the solution was treated with excess silver nitrate solution, and the precipitate was collected in a filter crucible. Weighings gave the following data:

Weight of crucible and AgCl	10.211 grams
Weight of crucible	10.102 grams

Another 50.00-milliliter portion was treated with a reducing agent before the addition of excess silver nitrate solution. The precipitate which formed was collected and weighed in another filter crucible.

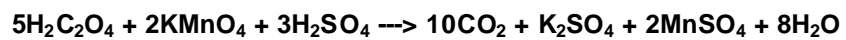
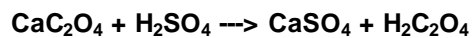
Weight of crucible and AgCl	10.247 grams
Weight of crucible	10.110 grams

- Determine the percentage of Cl<sup>-</sup> in the unknown.
  - Determine the percentage of ClO<sub>3</sub><sup>-</sup> in the unknown.
- 

1) Samples of a mixture of Ca<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub> and CaF<sub>2</sub> were analyzed to determine calcium and phosphorus.

a) Calcium was determined by conversion to CaC<sub>2</sub>O<sub>4</sub>, dissolution in H<sub>2</sub>SO<sub>4</sub>, and titration with KMnO<sub>4</sub>.

Reactions are as follows:



Data for one determination were as follows:

Weight of mixture 1.165 grams  
Volume of 0.105-molar  $\text{KMnO}_4$  47.18 milliliters

Calculate the per cent of calcium in the mixture.

b) Phosphorus was determined by conversion to  $\text{MgNH}_4\text{PO}_4$  and ignition to  $\text{Mg}_2\text{P}_2\text{O}_7$ .

Data for one determination were as follows:

Weight of mixture 1.684 grams  
Weight of  $\text{Mg}_2\text{P}_2\text{O}_7$  0.847 gram

Calculate the per cent of phosphorus in the mixture.

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