NATIONAL MATH + SCIENCE INITIATIVE

AP CHEMISTRY

Laboratory Based Questions

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The Lab Based Questions

Writing for a "5"

What I Absolutely Have to Know to Survive the AP Exam

If asked to do the following, it might indicate the question deals with laboratory questions:
Design an experiment; list measurements needed; show setup of calculations needed, use sample data to do calculations; interpret or draw graphs; explain the affect of error on the calculated value; use qualitative observations…

Parts of a Lab Question

You may not have done that exact experiment but you can use the knowledge gained doing other experiments to help account for certain observations.

You may be asked to describe how to do an experiment or to design an experiment.

- Materials: If they give you a list of equipment, don't think you have to use all of it.
- Procedure: Be sure to include important techniques like rinsing the buret with the solution before a titration or heating to constant mass.
- Data needed: The data needed are values that can be measured like initial and final temperatures. Writing all the mathematical equations needed to do the calculations will help you determine what data is needed.
- Calculations: A calculation is using what was measured like temperature change. Show the set up of the mathematical equations required for the calculations. Use sample data when appropriate.
- Graphs: Be sure you label the axes and other important points on your graph.
- Error Analysis: State whether the quantity will be too high, too low, or no change. Use equations to help you determine what change will occur and to support your answer.

Common Lab Procedure: Calorimetry

Calorimetry is used to determine the heat released or absorbed in a chemical reaction. A calorimeter can determine the heat of a solution reaction at constant pressure.

Techniques:

- Use a double Styrofoam cup with a plastic top and hole for the thermometer
- Determine the change in temperature accurately
- Measure solution volumes precisely
- Start with a dry calorimeter

Information to know about calorimetry:

- Heat capacity (C) = the amount of heat needed to raise the temperature of an object by one degree Celsius or Kelvin, J/°C or J/K.
- The heat capacity of 1 mol of a substance is called its molar heat capacity (Joules per mole per degree) J/mol·°C or J/mol·K.
- Specific heat, c, also known as specific heat capacity, is defined as the amount of heat necessary to raise the temperature of 1.00 g of a substance by one degree. Units are (joules per gram per degree), J/g·°C or J/g·K. You often use the specific heat capacity in analyzing gathered data then convert to molar heat capacity.

Assumptions often made during calorimetry: (Be able to answer error analysis questions about each assumption below)

- The density of dilute solutions is the same for water. D = 1.0 g/mL
- The specific heat of the solutions is the same as that for water. c = 4.184 J/g°C
- The solutions react in their stochiometric amounts.
- There is no loss of heat to the surroundings.

Equations:

- $q = mc\Delta T$ (c = 4.184 J/g°C)
- $q = \Delta H = mc\Delta T$, at constant pressure
- Use the density of the solution to convert from volume to mass for $q = mc\Delta T$

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Common Lab Procedure: Titration

A titration is a laboratory procedure for quantitative analysis. In a titration two reagents are mixed, one with a known concentration & known volume [or a solid with a known mass] and one with an unknown concentration. There is some way to indicate when the two reagents have reacted completely (typically an indicator), and at the end of the titration the unknown solution's concentration can be calculated since you have accurately determined the volume of that solution required to complete the reaction.

Terms to know:

- Titrant – A solution of known concentration; it is often standardized
- Standardized solution – A solution in which the exact concentration is known.
- Indicator - A weak acid or base used in a titration to indicate the endpoint has been reached.
- Equivalence point – moles of acid = moles of base; point at which enough titrant has been added to completely react with the solution being analyzed.
- End point – the point at which the indicator changes color; important to pick an indicator with a pKa very close to the pH at the equivalence point.

Techniques:

Preparing the Buret:

1. Rinse a clean buret with distilled water and then the titrant (the solution that will be added to the flask).
2. Allow the titrant to drain through the buret so that the tip gets rinsed with titrant as well.
3. Discard the rinse solution. Fill it with the titrant. Remove air bubbles from the tip of the buret by draining several milliliters of titrant.

Preparing the Sample:

4. Pipet the desired volume of the solution to be analyzed into an Erlenmeyer flask. Record the exact volume. If the sample is a solid, weigh the desired mass, add the solid to an Erlenmeyer flask, and dissolve it in distilled water (the amount of water does matter since it doesn't change the moles of the solid). Be sure to record the exact mass of sample used.
5. Add the titrant to the flask until the equivalence point is reached. Calculate the volume of titrant added.
6. Change in color of a chemical indicator is usually used to signal the endpoint of the titration. The endpoint for this titration is reached when you reach a pale color that persists for several seconds.

Measured Data Required:

moles titrant = **moles** of substance analyzed @ equivalence point

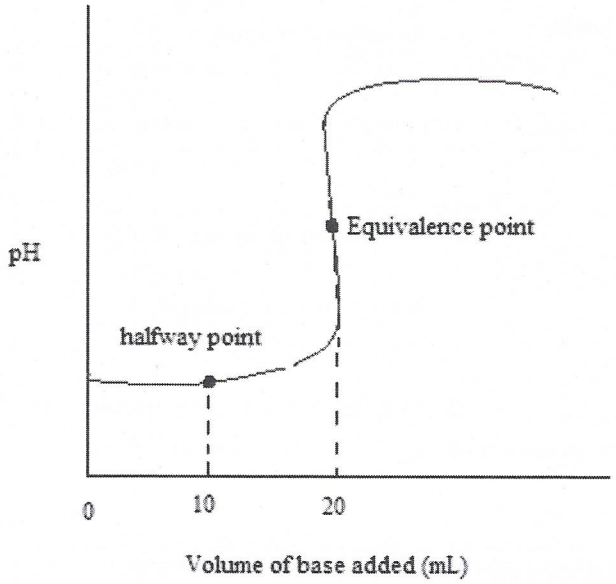
- mass of DRY substance analyzed OR accurately measured volume of solution analyzed [measure with a pipet OR buret]
- *initial* volume of titrant (substance of known molarity) and *final* volume of titrant (required to reach end point)
- Molarity of titrant

Calculation Hints:

- Substance analyzed is solution/liquid:
  - $M_1V_1 = M_2V_2$ @ equivalence point
  - Volume of titrant used to reach end point [difference between final and initial volumes]
  - $M_{titrant}V_{of\ titrant\ added}$ = moles of titrant = moles of unknown
  - Molarity of unknown = Molarity of unknown = $\frac{\text{moles of unknown}}{\text{Liters of unknown}}$

- Substance analyzed is solid:
  - Same as process as above
  - $M_{titrant} V_{of\ titrant\ added}$ = moles of titrant = moles of unknown
  - Molecular weight of the unknown $= \dfrac{\text{mass of solid dissolved}}{\text{moles of unknown}}$

Graphs:



Common Lab Procedure: Gravimetric Analysis

One method for determining the amount of a given substance in solution is to form a precipitate that includes the substance. The precipitate is then filtered and dried. This process is called gravimetric analysis.

Techniques/Procedure:

1. Weigh sample
2. Form precipitate
3. Filter precipitate (A buchner funnel and aspirator can be used)
4. Dry precipitate (Be sure to dry to constant mass)
5. Weigh precipitate

For example if we wanted to determine the amount of chloride ions present in a given solid, we would weigh the solid sample, dissolve the sample in water, add an excess of silver nitrate solution to form the precipitate silver chloride. This precipitant would be filtered, and dried to constant mass. From the mass of silver chloride formed, we can determine the moles of silver chloride and the moles of chloride ion in the original sample.

- Substance analyzed is solid:
  - Same as process as above
  - $M_{\text{titrant}} V_{\text{of titrant added}} = \textit{moles of titrant} = \textit{moles of unknown}$
  - Molecular weight of the unknown $= \dfrac{\text{mass of solid dissolved}}{\text{moles of unknown}}$

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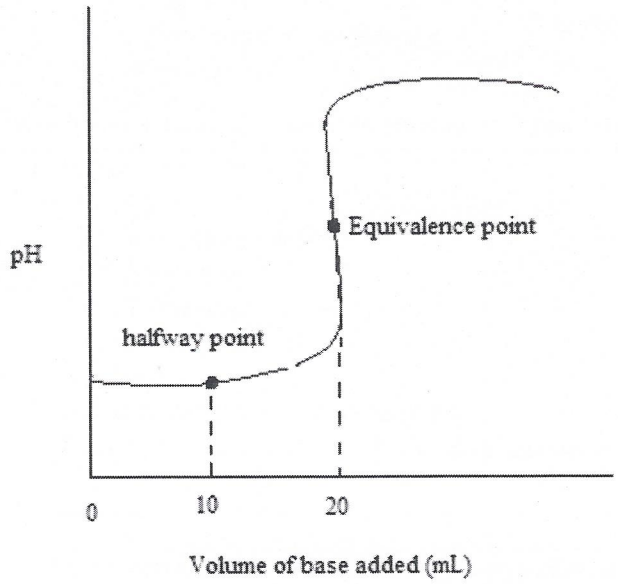

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Common Lab Procedure: Determining Molar Mass

Organize answer around calculations—paying special attention to what quantities are **measured** versus **calculated**!

$$\text{molar mass} = \frac{\text{mass of sample in grams}}{\text{moles of sample}}$$

Titration Data

- **moles** titrant = **moles** of substance analyzed @ equivalence point
- **Measured DATA** required:
  - mass of substance
  - *initial* volume of titrant (substance of known molarity) and *final* volume of titrant (required to reach end point)
  - Molarity of titrant
- **Calculations** required:
  - Molarity of titrant
  - Substance analyzed is solid:
  - $M_{titrant}V_{titrant}$ = moles titrant
  - $\text{molar mass} = \frac{\text{g of solid analzyed}}{\text{moles of titrant used}}$

Vaporization of a Volatile Liquid

- $PV = nRT$ used to determine moles
- **Measured DATA** Required:
  - Pressure = atmospheric pressure unless collected over water
  - *initial* mass of flask
  - *final* mass of flask
  - Temperature of boiling water—don't assume 100°C
  - Volume of gas = fill flask with water and measure the volume of water in a graduated cylinder
- **Constants** needed:
  - If collected over water, the water vapor pressure at the experimental temperature.
- **Calculations** Required:
  - If the gas was collected over water $P_{vapor} = P_{atmospheric/baraometric} - P_{water\ vapor\ at\ certain\ temperature}$
  - mass of sample = *final* mass of flask [includes vapors] – *initial* mass of flask

Common Lab Procedure: Colorimetric or Spectrophotometric Analysis

Colorimetric analysis is a quantitative analysis of a solution using color based on Beer's Law. Colorimetric analysis can be used to determine the concentration of an unknown solution, the rate constant of a reaction, the order of a reaction, etc.

Beer's Law is an expression than can be used to determine how much light passes through the solution. It also shows that concentration and absorbance are directly related.

$$A = \varepsilon bc$$

- A = absorbance (measured with a colorimeter)
- ε = molar absorptivity (how much light will be absorbed by 1 cm of a 1 M solution of the chemical)
- b = path length of the cuvette in cm
- c = concentration in molarity

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Lab Based Questions Cheat Sheet

Relationships

Calorimetry $q = \Delta H = mc\Delta T$ at constant pressure • c = specific heat • ΔT = change is temperature • m = mass • q = heat	Beer's Law $A = \varepsilon bc$ • A = absorbance • ε = molar absorptivity • b = path length of the cuvette in cm • c = concentration in molarity
Ideal Gas Law $PV = nRT$ • V = volume in liters • n = moles • T = temperature in Kelvin • P = pressure If collected using water displacement: $P_{vapor} = P_{atmospheric/baraometric} - P_{water\ vapor\ at\ certain\ temperature}$	Other important relationships: @ Equivalence point $M_1V_1 = M_2V_2$ $\text{molar mass} = \frac{\text{mass of sample in grams}}{\text{moles of sample}}$ $\%\text{error} = \frac{\text{accepted - experimental}}{\text{accepted}} \times 100$

Connections

Electrochemistry: draw diagram of galvanic or electrolytic cell, qualitative observations that can be made at the cathode and anode, etc.	Stoichiometry: empirical formula of a compound, percent of an element in a compound, etc.
Thermodynamics: heat of reaction, molar heat capacity, etc.	Acids and Bases: standardizing a solution, drawing a titration curve, etc.
Qualitative Analysis: identifying a compound based on observations and tests	

Potential Pitfalls

Be aware there are quantities you **measure** [such an initial temperature and final temperature or initial pressure and final pressure, etc.] and terms you **calculate** *using what you measured* such as Δ T or Δ P

Be sure to include important steps in the procedure like "heat to constant mass," "rinse the buret with distilled water and then the ***solution*** before titrating", "dissolve the solid in about 100 mL of water and then add water to the 500 mL mark on the volumetric flask," etc.

Writing the mathematical equations may help determine how an error affects the results. Use the equations to justify your error analysis.

If given a laboratory situation you have not specifically done, use the observations and the concepts you learned in your labs throughout the year to reason your way through the lab question.

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NMSI SUPER PROBLEM

A sample of 6 *M* hydrochloric acid (about 5 mL) is placed at the bottom of the test tube and then carefully filled with distilled water so not to disturb the HCl. A piece of magnesium is placed on top of the distilled water and then the test tube is inverted into a beaker with distilled water. The magnesium is allowed to react completely. The gas is collected in the test tube at room temperature. Before the test tube with the gas is removed, the water levels inside and outside the test tube are the same. The chemicals and lab equipment used are listed below.

6 *M* hydrochloric acid	Barometer	Strip of magnesium metal	600 mL beaker
Distilled water	Balance	Test tube	
Table of water vapor pressures	Graduated cylinder	Thermometer	

(a) Write the balanced net ionic reaction that occurs when hydrochloric acid reacts with magnesium.

(b) List the *measurements* needed to calculate the molar volume of the gas produced?

(c) Show the setup for the calculations needed to determine:

(i) the moles of gas produced

(ii) the volume of the gas produced

(iii) the molar volume of the gas (at STP)

(d) What test can be done to prove which gas is produced?

(e) What is the purpose of making the water level inside the test tube equal to the water level in the beaker?

(f) If the vapor pressure of water is not used in the calculation, how will the molar volume of the gas at STP be affected? (higher, lower or the same) Explain.

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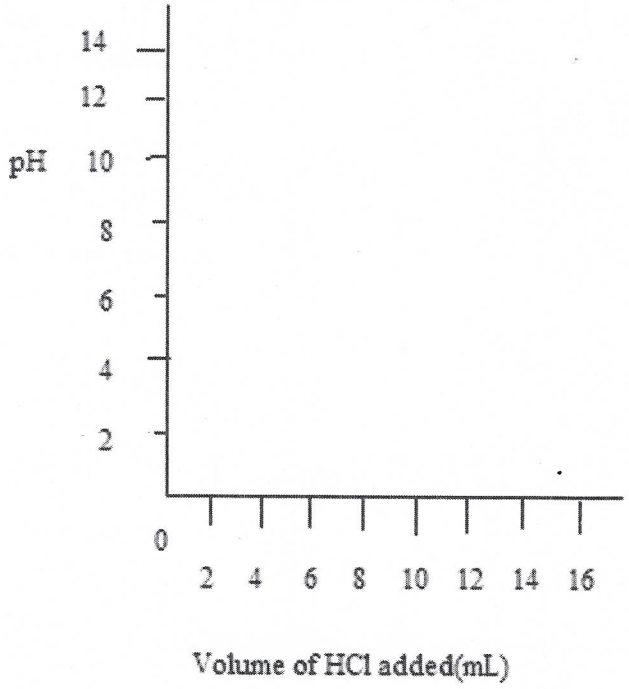
Another sample of hydrochloric acid was used to titrate a solution of calcium hydroxide.

(g) Describe the steps needed to make 50.0 mL of a 0.50 *M* solution from the 6 *M* solution of hydrochloric acid, using a dropper, 5.0 mL pipet, 50.0 mL volumetric flask, and distilled water. Be sure to mathematically justify your process.

(h) Write the balanced net ionic equation for the dissociation of $Ca(OH)_2(s)$ in aqueous solution, and write the equilibrium-constant expression for the dissolving $Ca(OH)_2$.

A 15.0 mL unknown solution of $Ca(OH)_2$ is titrated with a standardized 0.50 *M* solution of hydrochloric acid. It takes exactly 10.5 mL of hydrochloric acid to reach the equivalence point where the pH is 8.50.

(i) Sketch the titration curve that shows the pH change as the volume of hydrochloric acid added increases from 0 to 16.0 mL. Be sure to label the equivalence point of the titration.



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(j) Calculate the concentration of $[OH]^-$ from the titration.