DATA SHEETS AND CALCULATIONS FOR ACIDS & BASES

Name

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Partner's Name

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Grade and Instructor Comments

Part 1: Experimental Measurement—Determining a Numerical Value for K_w

Experimental pH of 0.010 M NaOH = ____________

Part 1: Calculations——Determining a Numerical Value for K_w

What is your measured pH? _______________________

Based on the measured pH, what is the hydronium ion concentration? $[H_3O^+]$ = ______________ M

Knowing that in a 0.010 M NaOH solution, $[OH^-]$ = 0.010 M, calculate a value for K_w from your experimental value of the measured hydronium ion concentration and the known OH^- concentration of the 0.010 M NaOH.

$K_w = [H_3O^+][OH^-] =$

and $pK_w = -\log K_w =$

Compare your results with the data taken from the scientific literature:

T (˚C)	**K_w**	**pK_w**
20	0.68×10^{-14}	14.17
25	1.01×10^{-14}	14.00

K_w, as is the case for all equilibrium constants, varies with temperature. However, at a given temperature, K_w is a constant for an aqueous solution. This means that at 25 ˚C in any aqueous solution, regardless of solute, the value of K_w $\{=[H_3O^+][OH^-]\}$ is 1.01×10^{-14}.

What is the pH of pure water at 25 ˚C?

Part 2. Determination of K_a for the Ammonium Ion—Experimental Measurements

Data Table for Solutions of Ammonium Ion and Ammonia

Enter the experimental pH values you determine in the lab in the last column. Complete the open areas of the table.

** Be very careful to rinse your glass electrode thoroughly with water before an after making this measurement.*

Solution	$[NH_4Cl]$, M	$[NH_3]$, M	Solute Type (Acid, Base, or Acid + Conj. Base	Enter Your **Experimental pH**
A	0.10	0	Acid	
B	1.0	0		
C	0.050	0.050		
D	0.50	0.50		
E	0	0.10		

Part 2: Calculating K_a for the Ammonium Ion

a) Write the balanced, net ionic equation for the reaction of ammonium ion with water.

b) Write the equilibrium constant expression for K_a for aqueous NH_4^+

c) Enter your experimental information for $[NH_3]$ and $[NH_4^+]$ (from the previous page) into the table below. Use your measured pH values for each solution (A-E) to calculate $[H_3O^+]$ and enter these values in the table below. Finally, calculate K_a for the ammonium ion and enter the values in the table. *Show one representative calculation here.*

Average calculated K_a value = ______________________ **and average pK_a =** ____________________

Solution	$[H_3O^+]$, M	$[NH_3]$, M	$[NH_4^+]$, M	Calculated K_a for NH_4^+	Calculated pK_a for NH_4^+
A					
-B					
C					
D					
E				Not a required calculation	Not a required calculation

Part 3. Properties of NH_4^+/NH_3 Buffer Solutions—Experimental Measurements

- Your experimental readings from page 70 are entered in the column labeled "Initial pH."
- NOTE: make a pH measurement on pure water before adding acid or base.
- Data for the pH after the addition of excess acid or base is entered into the columns marked "pH on adding H^+" and "pH adding OH^-" as appropriate.
- Fill in the boxes marked "ΔpH" with calculated numbers.

Solution	$[NH_4^+]$, M	$[NH_3]$, M	ΔpH on adding H^+	pH on adding H^+	**Initial pH**	pH on adding OH^-	ΔpH adding OH^-
A	0.10	0					
-B	1.0	0					
C	0.050	0.050					
D	0.50	0.50					
E*	0	0.10					
-Pure water							

Part 3: Properties of NH_4^+/NH_3 Buffer Solutions—Questions and Calculations

Effect of Dilution on the pH of a Buffer:

If the solution is diluted more than 10-fold, which solution — 1.0 M NH_4Cl (solution B) or 0.50 M NH_4Cl + 0.50 M NH_3 (solution D) —does the pH change more? *(Base your answer on the data in the "Initial pH" column on page 76.)*

Explain, on the basis of the K_a expression, why dilution has less effect on the pH of a buffer solution than on the pH of a solution containing only the acid as a solute (here NH_4^+).

Effect of Added H_3O^+ and OH^- on a Buffer

Compare the values of ΔpH (the changes in pH) for solutions C and D (in the table on page 15) with those for solutions of the acid along (A and B) or conjugate base alone (E).

a) Which solutions show a buffering action?

b) Write balanced chemical equations for reactions that prevent larger changes in pH.

Part 4. Titration Curves

The change in indicator color in an acid-base titration is a signal that the equivalence point is very near. Here you test two indicators that change colors in two different pH ranges.

Indicator	Color in Acidic Solution	Color in Basic Solution
Bromcresol green		
Phenolphthalein		

~~See *Chemistry & Chemical Reactivity*, page 872, Figure 18.10 for indicator colors.~~

Titration Results: Option (a)—HCl + NaOH

The volumes of NaOH in the table are suggested values. Enter your actual volumes of NaOH used in the table (second and fifth columns). Enter experimental data in every cell in the table.

Deductions from the HCl Titration Curve

Suggested V_{NaOH}, mL	Actual V_{NaOH}, mL	Measured pH	Indicator Color	Suggested V_{NaOH}, mL	Actual V_{NaOH}, mL	Measured pH	Indicator Color
0				22			
3				23			
6				24			
8				24.5			
10				25			
12				25.5			
14				26			
16				28			
18				30			
20				32			

Be sure to attach to your report form a carefully drawn plot of pH versus volume of base added. Be sure your name appears on the plot.

a) Write a balanced, net ionic equation for the reaction that occurs during the titration.

b) How many equivalence points can you detect? Explain the connection between the number of equivalence points and the reaction occurring.

c) CLEARLY LABEL on your titration curve the formulas for the species present at:

i) before adding NaOH

(ii) after 15 mL of NaOH has been added

ii) at the equivalence point

d) What is the connection between the indicator colors and the equivalence point?

Titration Results: Option (b)—H_3PO_4 + NaOH

The volumes of NaOH in the table are suggested values. Enter your actual volumes of NaOH used in the table (second and fifth columns). Enter experimental data in every cell in the table.

Suggested V_{NaOH}, mL	Actual V_{NaOH}, mL	Measured pH	Indicator Color	Suggested V_{NaOH}, mL	Actual V_{NaOH}, mL	Measured pH	Indicator Color
0				19			
3				19.5			
6				20			
8				20.5			
9				21			
9.5				22			
10				23			
10.5				24			
11				26			
12				28			
13				30			
15				35			
17				40			

Be sure to attach to your report form a carefully drawn plot of pH versus volume of base added. Be sure your name appears on the plot.

Deductions from the Phosphoric Acid Titration Curve

a) Write balanced, net ionic equations for the three possible successive reactions that occur during the titration.

1.

2.

3.

b) How many equivalence points can you detect?

c) CLEARLY LABEL on your titration curve the formulas for the species present at:
   i) the equivalence points
   ii) between the equivalence points

d) For which reaction or reactions (in a above) did you NOT see an equivalence point?

e) Write the equilibrium constant expressions for K_1 and K_2 of H_3PO_4

f) Determine the pK values for H_3PO_4 from your curve.

pK_1 (H_3PO_4) = ______________ and so K_1 (H_3PO_4) = _______________

pK_2 (H_3PO_4) = ______________ and so K_2 (H_3PO_4) = _______________

g) Calculate the ratio of experimental K_1 and K_2 values:

K_1/K_2 = _________________________

h) What are the values of pH at the first and second equivalence point on your pH titration curve for phosphoric acid?

pH at 1st equivalence point_____________________________

pH at 2nd equivalence point_____________________________

Explain why bromcresol green and phenolphthalein are suitable indicators for determining the concentration of a phosphoric acid solution. (See Figure 18.10 on page 872 of *Chemistry & Chemical Reactivity*.)