

## EXPERIMENT

# Titration Curve

## Goal

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The goal of this laboratory is to graph a titration curve and to use this curve to standardize an acid solution while determining the acidity constant of the acid.

## Background

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Titration curves were used in previous experiments to analyze the concentration of a solution, the so-called *analyte*. Using the *titrant*, a reactant of known concentration, it is possible to determine the concentration of the analyte if the volumes are measured carefully. The titrant is added from a buret until the solution changes color due to the presence of an indicator. The indicator denotes the moment when the *equivalence point* is achieved. The equivalence point is the moment at which equivalent molar amounts of both reactants have been added.

In an acid-base reaction, the acid reacts with the base, forming salt and water. This process is called *neutralization*; the protons released by the acid are neutralized by the hydroxide ions formed in the presence of the base. For strong acids and strong bases that dissociate completely in water, the equivalence point occurs at pH=7 and, therefore, an indicator with a color transition at pH=7 is used.

However, not all acids have the same strength. Weak acids do not release as many protons as strong acids do. The strength of a weak acid, and the consequent number of protons released, depend on its dissociation constant,  $K_a$ . Because of the partial dissociation of these species, the equivalence point in a weak acid titration does not occur at pH=7. The moment at which equivalent amounts of a weak acid and a strong base (titrant) have been added will take place at pH>7 (due to the non-complete dissociation). Determining the equivalence point is necessary to know the correct indicator to be used for titrations of weak acids.

### The titration curve

During the neutralization process of an acid, the pH of the solution starts at a low value - lower than 7 since it is an acid - and increases with the addition of the base. If one plots the pH versus the volume of base added, the result is not a straight line, but an S-shape curve. This shape shows that the pH changes slowly at the beginning and the end of the process, but there is a region in the graph where the pH increases rapidly with small additions of the base. When that abrupt pH change takes place depends on the acid and the base used.

For the titration of a weak acid (analyte) using a strong base (titrant), the endpoint will be achieved at a pH higher than 7. For the titration of a weak base (analyte) using a strong acid (titrant), the endpoint will be achieved at a pH lower than 7. The titration curve for every base or acid is unique and it is useful for determining the dissociation constant ( $K_a$  or  $K_b$ ) of the analyte.

### The acid dissociation constant

The acid dissociation constant for a weak acid is a ratio of the amounts of products and reactants at the equilibrium position. The ratio is constant for a fixed temperature. The concentration of  $H^+$  will always be among the products for the dissociation of an acid. Recalling that the concentration of  $H^+$  defines the pH, the latter can be solved as follows:

$$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$$

$$K_a = \frac{[H^+][A^-]}{[HA]} \quad [H^+] = K_a \frac{[HA]}{[A^-]} \quad ; \quad pH = pK_a + \log \frac{[A^-]}{[HA]}$$

This expression is known as the *Henderson-Hasselbalch equation*, where  $[HA]$  is the concentration of the weak acid, and  $[A^-]$  is the concentration of its conjugated base. It is possible to visualize the  $pK_a$  from the titration curve by using the Henderson-Hasselbalch equation.

First, notice that at the equivalence point, the same amounts of acid and base have reacted and there will be no acid;  $[HA] = 0$ . But *halfway* to that point, when half of the acid has been converted to the same amount of its conjugated base, the concentrations of the two will be equal,  $[A^-] = [HA]$ . The fraction in the previous expression becomes equal to 1 and, since  $\log 1 = 0$ , then  $pH = pK_a$ .

$$pH = pK_a + \log \frac{[A^-]}{[HA]} \quad ; \quad pH = pK_a + \log 1 \quad ; \quad pH = pK_a$$

The equivalence point would correspond to the inflection point of the titration curve. By reading the volume of base added at that point, it is possible to calculate the volume needed to reach the halfway point ( $V_{eq}/2$ ). The pH when that volume of base was added is taken from the titration curve and corresponds to the weak acid's  $pK_a$ .

The location of the equivalence point can be estimated by fitting two straight lines to the straight sections of the curve, as depicted in Figure 1, and a third line equidistant to the previous two. The point at which the third line and the titration curve intersect is the equivalence point.

## The experiment

The experimental part consists of the titration only. The data will be used later to plot the curve. From the plotted curve the equivalence point and the acid  $pK_a$  should be calculated.

The weak acid studied in this experiment is potassium hydrogen phthalate,  $KHC_8H_4O_4$  (KHP). The base used as a titrant is sodium hydroxide (NaOH), a strong base. It is very important to obtain precise measurements of the volumes and masses used during this experiment.

The analyte (KHP) will be dissolved in a beaker. The titrant (NaOH) will be added from a buret. A pH meter will be used to monitor the pH and magnetic stirring will be used during the measurement to homogenize the solution pH.

### Example

A solution of hydrogen cyanide (HCN), a weak acid, is titrated using NaOH, a strong base. The pH is monitored during the titration using a pH-meter. The following plot is obtained:

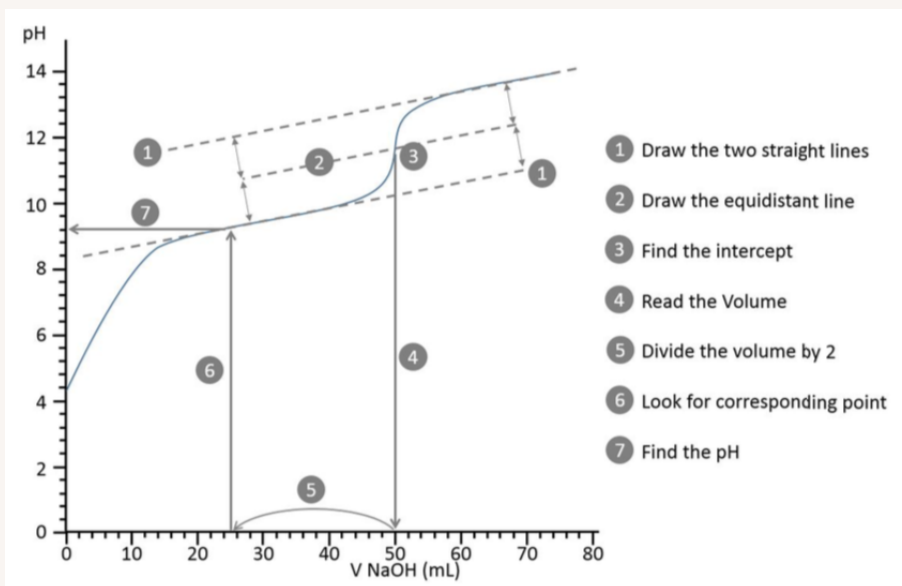


Figure 1: Titration curve of a HCN solution using NaOH. The numbers indicates the steps to follow as indicated in the answer.

Use the graphical method to identify the equivalence point for the titration, the volume of base added at the equivalence point, the midpoint, and the HCN  $K_a$ .

**Answer:** The equivalence point is obtained from the graph. (1) Draw the lines for the 2 straight sectors. (2) Draw





the equidistant line. (3) Identify the equivalence point where the equidistant line intercepts the curve. (4) Draw a vertical line to find the volume; 50.0 mL. (5) The half-way point would be:  $50.0 \text{ mL} / 2 = 25.0 \text{ mL}$ . (6) Draw a vertical line up to the curve. (7) From the curve draw an horizontal line to find the pH; 9.21. The  $pK_a = 9.21$  (7), and therefore  $K_a = 10^{-9.21} = 6.2 \times 10^{-10}$

## Procedure

### Preparing the setup

- ☐ *Step 1:* – Obtain a 50 mL buret, rinse it with distilled water, and verify that it operates correctly (you should be able to add drops but also larger quantities).
- ☐ *Step 2:* – Obtain a hot plate with magnetic stirring, a magnetic stirring rod, a stand, two 150 mL beakers, and a pH meter.
- ☐ *Step 3:* – Arrange the setup so that the beaker is on the hot plate, the pH meter is to its left, with the probe inserted in one of the beakers, and the tip of the buret is also slightly inserted into the beaker. The number markings on the buret should be facing you, and the valve should be accessible and operative.
- ☐ *Step 4:* – Rinse the buret with a small amount of 0.1 M NaOH solution. Discard the rinsing liquid. Fill the buret with fresh NaOH solution. Your instructor will either give you an unknown or instruct you to use the already prepared NaOH solution. Discard the first 3-4 mL (the volume of solution in the tip of the buret used for the rinsing).

#### Good Lab Practice

-  Burets are read from top to bottom.
-  The volume added is calculated by difference. It is not necessary to start from 0 mL.
-  Use a funnel to fill the buret but remove it always before the first reading. Drops from the funnel will affect the measurements.
-  Do **never** refill the buret with the Erlenmeyer/beaker of the analyte below it. A spill would ruin the titration.

### Preparing the KHP solution

- ☐ *Step 1:* – Put one of the beakers on the scale and set it to zero (T/O). The beaker must be dry. Remove the beaker from the scale to add a small amount of KHP. You need an amount between 0.45 and 0.55 grams. Record the value with all the decimals.
- ☐ *Step 2:* – Obtain 50.0 mL of distilled water in a graduated cylinder.
- ☐ *Step 3:* – Carefully add the 50.0 mL of distilled water from the graduated cylinder into the beaker containing the KHP.
- ☐ *Step 4:* – Prepare a second solution on the second beaker.

### Measuring

- ☐ *Step 1:* – Add the magnetic stirrer to one of the beakers, put the beaker on the hotplate, and set a gentle stirring speed. Do not start the heat on the plate!
- ☐ *Step 2:* – Clean the pH-meter probe before immersing it in the KHP solution.
- ☐ *Step 3:* – Either you are given an unknown or you will use the NaOH already prepared from the lab. Read and record the initial pH before adding any NaOH.

- ☐ *Step 4:* – Start adding the titrant (the base). After every addition (see next step) record the **buret reading** (not the volume added) and the corresponding pH measurement.
- ☐ *Step 5:* – NaOH should be added initially in 1 mL increments, but at a certain moment, additions should be reduced to 0.2 mL portions. Mind that the pH will change very slowly at the beginning and the end of the titration. However, around the equivalence point, the pH will change extremely fast with small additions of the titrant.
- ☐ *Step 6:* – Read the buret and record the volume with all the digits given by the buret plus the estimated digit.
- ☐ *Step 7:* – Add the titrant in 1 mL portions making sure the buret readings are precise.
- ☐ *Step 8:* – After every addition, estimate how much the pH changed from the previous pH measurement. If the pH has changed by less than 0.3 pH units, add 1 mL of titrant in the next step. If the pH has changed by 0.3 pH-units or more, begin adding the titrant in 0.2 mL portions.
- ☐ *Step 9:* – When adding 0.2 mL portions, if the pH change starts to slow down and increases by less than 0.3 pH-units between measurements **for at least 3 measurements in a row**, then you can start adding 1 mL portions again.
- ☐ *Step 10:* – Stop the titration when the pH exceeds 11.5.
- ☐ *Step 11:* – Do a second titration trial. Clean and reuse the magnetic stirred; in the results section, plot by hand the titration plot.

## Calculations

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- ① This is the molarity of the base solution, written on the label.
- ① This is the mass of your sample.
- ② These are the moles on your sample:

$$n_{\text{KHP}} = \frac{\textcircled{1}}{204.22}$$

- ③ These are the molarity on your acid sample:

$$c_a = \frac{\textcircled{2}}{50} \times 10^3$$

- ④ This is the burette reading in mL.
- ⑤ This is the PH as given by the OH meter.
- ⑥ This is the cumulative volume of base added.

**STUDENT INFO**

Name:

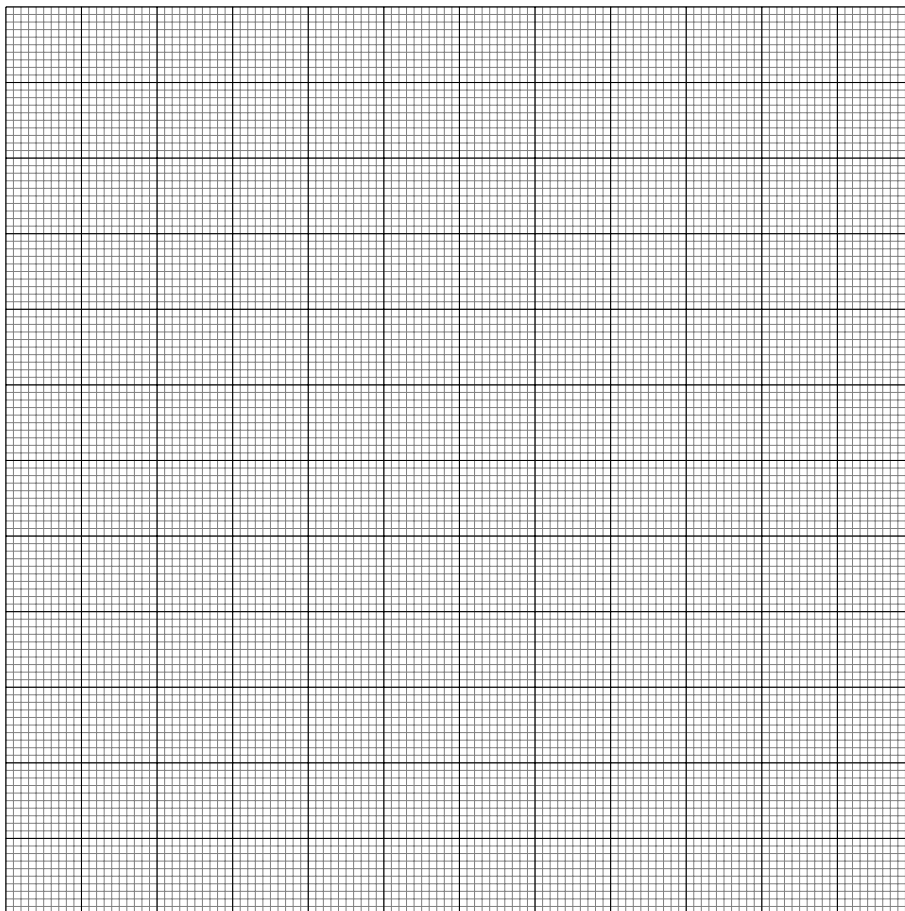
Date:

**Pre-lab Questions****Titration Curve**

1. The data below represents the titration of a weak base with a strong acid. Use the data below to answer the following questions:

$V_{HCl}^{Added}$ mL	PH	$V_{HCl}^{Added}$ mL	PH	$V_{HCl}^{Added}$ mL	PH	$V_{HCl}^{Added}$ mL	PH
0	11.12	45.01	8.3	49.97	6.1	50.32	3.5
0.71	10.9	45.95	8.2	49.98	5.9	50.4	3.4
1.48	10.7	47.89	7.9	49.99	5.7	50.63	3.2
1.96	10.6	48.31	7.8	50.00	5.4	50.8	3.1
4.03	10.3	48.65	7.7	50.01	5	51.01	3
6.18	10.1	49.31	7.4	50.02	4.7	51.27	2.9
11.06	9.8	49.45	7.3	50.03	4.5	51.61	2.8
13.18	9.7	49.56	7.2	50.04	4.4	52.04	2.7
18.12	9.5	49.72	7.0	50.06	4.2	52.58	2.6
20.85	9.4	49.78	6.9	50.08	4.1	53.27	2.5
26.58	9.2	49.83	6.8	50.1	4	54.15	2.4
29.41	9.1	49.89	6.6	50.13	3.9	55.28	2.3
37.02	8.8	49.91	6.5	50.16	3.8	56.73	2.2
40.94	8.6	49.93	6.4	50.2	3.7	58.63	2.1
42.53	8.5	49.95	6.3	50.25	3.6	61.11	2

- 1.1. Plot PH (vertical axis) vs. volume of titrant added (horizontal axis), identifying the initial, mid, and endpoints.



1.2. Fill the table below based on the plot above:

Experimental Data						
	Initial Point		Midpoint		Endpoint	
	$V_{\text{HCl}}^{\text{Added}}$	PH	$V_{\text{HCl}}^{\text{Added}}$	PH	$V_{\text{HCl}}^{\text{Added}}$	PH
	0					

1.3. Use the midpoint to compute the basicity constant of the titrate given that at that point  $\text{PH} = \text{PK}_b$ . Write down the  $K_b$  value under the graph.

2. Calculate the PH of a 0.5M solution of an acid with  $K_a = 1.4 \times 10^{-5}$ .

STUDENT INFO	
Name:	Date:

Date:

# Titration Curve

①  $c_{\text{NaOH}} \text{ (M)} =$  \_\_\_\_\_

[illegible][illegible]









Experimental Data						
	Initial Point		Midpoint		Endpoint	
	$V_{\text{NaOH}}^{\text{Added}}$	PH	$V_{\text{NaOH}}^{\text{Added}}$	PH	$V_{\text{NaOH}}^{\text{Added}}$	PH
Trial 1	0					

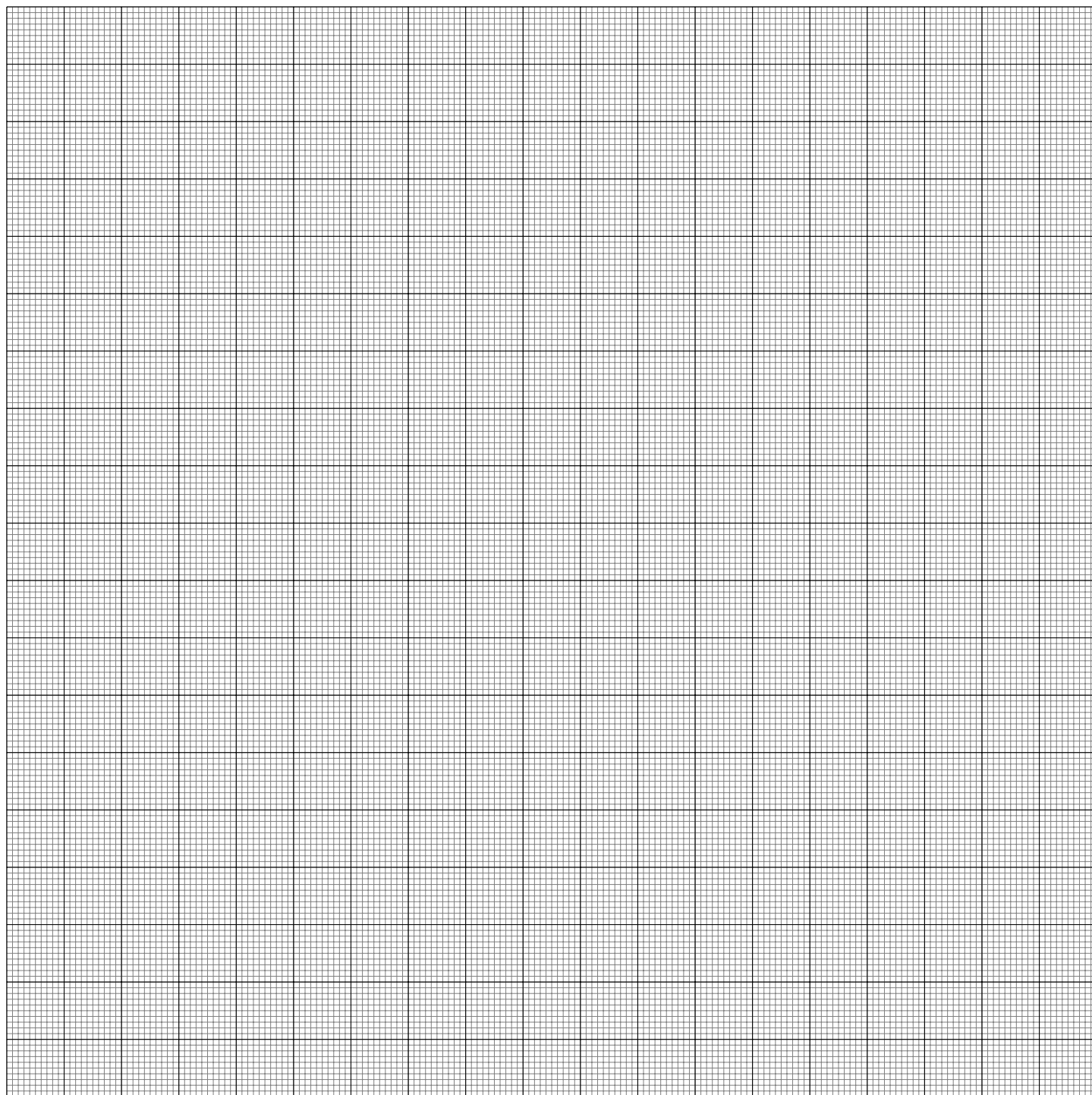


Figure 2: PH (Y axis) vs.  $V_{\text{NaOH}}^{\text{Added}}$  (X axis)

Experimental Data						
	Initial Point		Midpoint		Endpoint	
	$V_{NaOH}^{Added}$	PH	$V_{NaOH}^{Added}$	PH	$V_{NaOH}^{Added}$	PH
Trial 2	0					

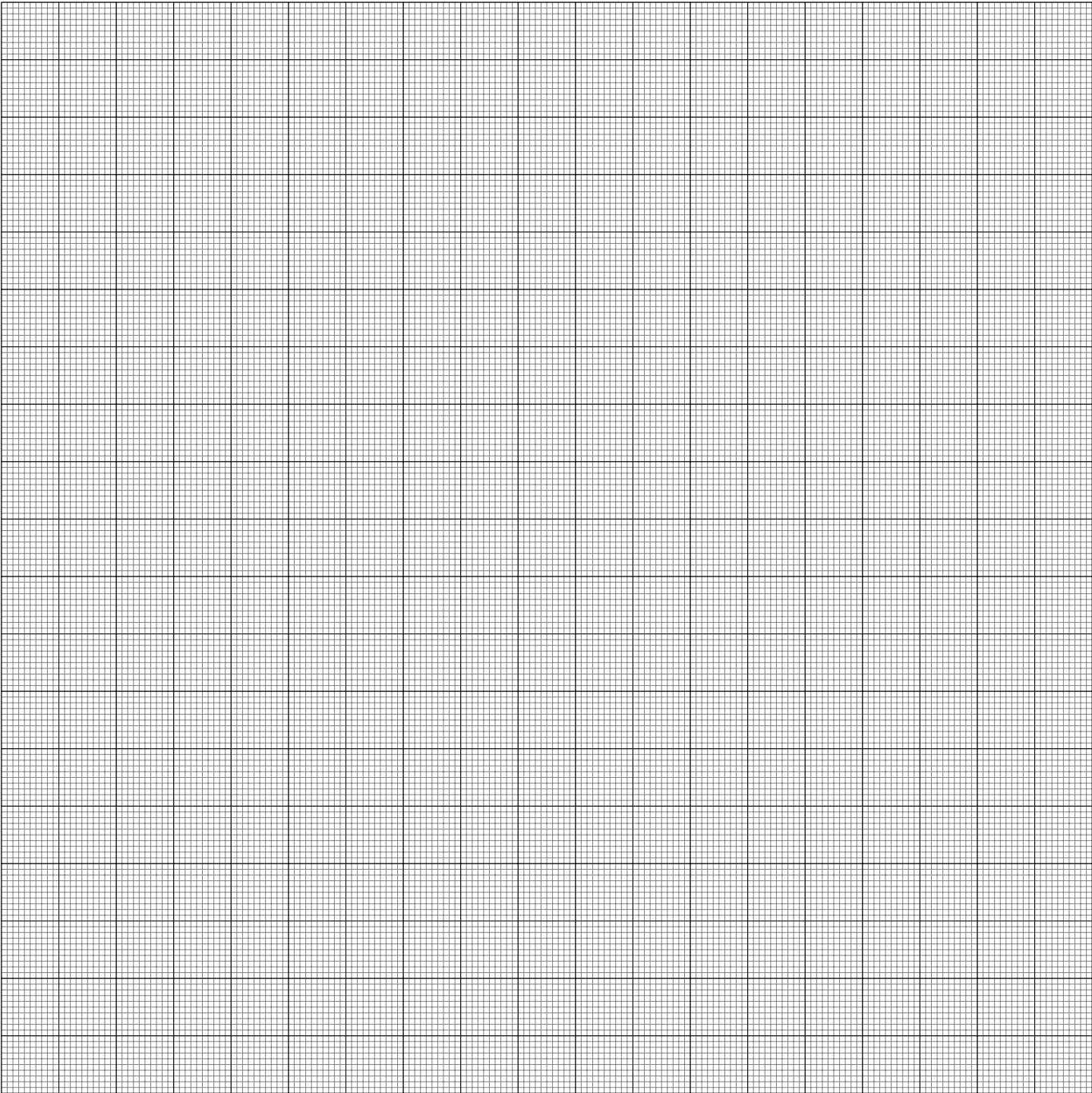


Figure 3: PH (Y axis) vs.  $V_{NaOH}^{Added}$  (X axis)

**STUDENT INFO**

Name:

Date:

**Post-lab Questions****Titration Curve**

1. Preparing a NaOH solution, starting from a solid NaOH sample, is difficult because NaOH is very hygroscopic (absorbs water from the air). This affects the mass measured and therefore the real concentration expected. The NaOH solution used was estimated to be 0.10 M but it might not be correct. Using the information from the titration (mass of the phthalate), and knowing that the acid and the base react 1:1, calculate the real concentration from each trial, and the mean value.

	Trial 1	Trial 2
KHP mass	_____	_____
KHP moles	_____	_____
$V_{NaOH}$ @ equiv.	_____	_____
Real $c_b$ , [NaOH]	_____	_____
Mean $c_b$ , [NaOH]	_____	

2. Calculate the volume of NaOH for the halfway point (half of the volume added at the equivalence point) and locate in the graph the pH for the solution after that volume was added. Use the value of the pH to calculate the KHP  $pK_a$ .

$V_{NaOH}$ @ halfway	_____	_____
pH @ halfway	_____	_____
$pK_a$	_____	_____
Mean $pK_a$	_____	

3. Using the  $pK_a$  mean value, calculate KHP's  $K_a$ . Write the whole expression for  $K_a$ .

4. Using the KHP's  $K_a$  and its initial concentration, calculate the theoretical pH before any NaOH was added, and compare the calculated values to the pHs measured experimentally.

	Trial 1	Trial 2
KHP moles	_____	_____
$c_a$ , [KHP]	_____	_____
Theoretical pH	_____	_____
Experimental pH	_____	_____

5. Using the  $K_a$  for KHP, calculate  $K_b$  for its conjugated base,  $C_8H_4O_4^{2-}$ , given that  $K_w = K_a \times K_b$ .

## EXPERIMENT

# Strength of Laundry Bleach

## Goal

The goal of this laboratory is to perform an **oxidation-reduction titration** to determine the concentration of sodium hypochlorite (NaOCl) in commercial bleach samples.

## Materials

- |                                                       |                                               |
|-------------------------------------------------------|-----------------------------------------------|
| <input type="checkbox"/> A buret, a stand and a clamp | <input type="checkbox"/> 10% diluted bleach   |
| <input type="checkbox"/> 100mL, 50mL beakers          | <input type="checkbox"/> 10mL transfer pipet  |
| <input type="checkbox"/> 125mL Erlenmeyer             | <input type="checkbox"/> 2M HCl solution      |
| <input type="checkbox"/> 0.025M $Na_2S_2O_3$ solution | <input type="checkbox"/> 0.2% starch solution |
| <input type="checkbox"/> solid KI                     |                                               |

## Background

The properties of chlorine for bleaching fabrics were discovered in the 18th century by the French scientist Claude Berthollet. He was able to make an aqueous solution of sodium hypochlorite which was named after the quarter in Paris where it was produced, Javel. *Eau de Javel* (Javel water) is still used in French to name bleach.

At the beginning of the 18th century, the pharmacist Antoine Germain Labarraque discovered the disinfecting ability of another hypochlorite solution, this time calcium hypochlorite. The so-called "Eau de Labarraque" was employed in the disinfection of animal guts to make musical instrument strings. Hypochlorite was the first antiseptic product, helping to treat gangrene and putrescent wounds in people in the 1820s, and ending with the propagation of "cadaveric particles" in morgues.

Nowadays bleach is a common product at home, used to whiten clothes, remove stains, and as a disinfectant. Because of its strong bactericidal properties, hypochlorite is also used for disinfecting and to prevent the proliferation of algae in swimming pools. Bleach can be consumed. The CDC recommends adding eight drops of bleach per gallon of clear water. For cloudy water, the bleach quantity should be doubled to 16 drops per gallon. Bleach can be found in powder form as well, in the form of Bleaching powder. This is indeed, a solid combination of chlorine and slaked lime, introduced in 1799 by Scottish chemist Charles Tennant.

## Chlorine

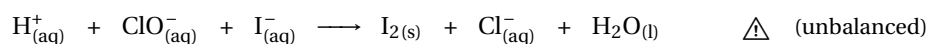
Chlorine forms four different polyatomic anions with oxygen:



These are called **oxoanions** and illustrate 4 possible oxidation states of chlorine atoms; +1, +3, +5, and +7 respectively. When the oxidation state of the chlorine decreases, for instance going from  $\text{ClO}_4^-$  to  $\text{ClO}_3^-$  (+7  $\rightarrow$  +5), it reduces. For reduction to happen another species must be oxidized. Chlorine is then acting as the **oxidizing agent**. Perchlorate it is the strongest oxidizing agent of the 4 oxoanions, but even  $\text{ClO}^-$  can still be reduced to  $\text{Cl}^-$  (Cl oxidation state -1).

## The first oxidation

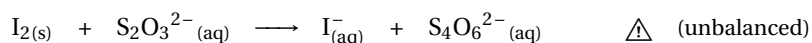
To determine the concentration of  $\text{ClO}^-$ , two consecutive oxidation-reduction reactions will be performed. In the first reaction, all the  $\text{ClO}^-$  present in the sample will be oxidized into  $\text{Cl}^-$ . The **oxidizing agent** will be  $\text{I}^-$ , which will be in turn reduced into  $\text{I}_2$ . The reaction, carried out in an acidic medium, will read as:



The reaction above is not balanced. You can determine the number of electrons being transferred using the oxidation states to help you balance the reaction. An excess amount of  $\text{I}^-$  will be used to ensure that all the  $\text{ClO}^-$  is oxidized. The iodine ( $\text{I}_2$ ) formed in this reaction will be therefore limited by the amount of hypochlorite present in the sample. Keep in mind that the amount of hypochlorite is what we want to determine, and it can be stoichiometrically calculated from the amount of iodine produced. In the following step, the amount of  $\text{I}_2$  will be determined using an oxidation-reduction titration. The solution might turn slightly brownish due to the presence of a Lewis complex between iodine and water.

## The oxidation-reduction titration

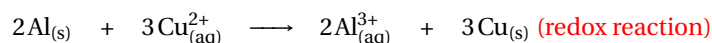
The titration to be performed accounts for the following redox reaction:



Again, this reaction is not balanced. Iodine ( $\text{I}_2$ ) is reduced back to iodide  $\text{I}^-$ . Thiosulfate  $\text{S}_2\text{O}_3^{2-}$  is the reducing agent, being itself oxidized into tetrathionate ions ( $\text{S}_4\text{O}_6^{2-}$ ). The titration endpoint will be determined by adding starch as an indicator at the beginning of the reaction. Remember from previous experiments that starch and  $\text{I}_2$  react turning the solution into a dark blue color, therefore, the titration should be stopped when all the iodine has been reduced and the solution becomes transparent.

## Redox reactions

Redox reaction are different than acid-base or precipitation reaction. They contain the same chemical element in two different states resulting from the loss or win of electrons. For example:



We have that neither  $\text{Al}_{(\text{s})}$  or  $\text{Cu}^{2+}_{(\text{aq})}$  are an acid or a base. Also there is no product precipitate. Hence, this reaction is not an acid-base reaction or a precipitation reaction. This is a redox reaction, as we have Al in two different states: as metallic  $\text{Al}_{(\text{s})}$  and as ionic  $\text{Al}^{3+}_{(\text{aq})}$ , which result from the loss of an electron. In redox reaction there is always elements in the chemicals that

lose electrons and chemicals winning electrons. The redox number, also called oxidation number or oxidation state helps compare elements that have lost (or gained) different charge. When comparing two elements with different redox numbers, the larger this number the more electrons the element has lost. Similarly, the smaller this number the smaller this number the more electrons the element has won.

We indicate redox numbers with roman number on top of the element. For example the redox number of manganese in this compound is +7:  $\text{Mn}^{\text{VII}}\text{O}_4^-$ . The redox number can be a non integer number. There is five rules to identify the redox number of an element.

🔗 **Rule 1** Single atoms or elements have zero redox number. Examples are Na or  $\text{H}_2$ , both with redox zero.

🔗 **Rule 2** Monoatomic ions have redox number equal to their charge. Examples are  $\text{Na}^+$  or  $\text{Cl}^-$  with redox +1 and -1, respectively.

🔗 **Rule 3** The redox number of fluorine is -1

🔗 **Rule 4** The redox number of hydrogen on its covalent (e.g.  $\text{H}_2\text{O}$ ) compounds is +1.

🔗 **Rule 5** The redox number of oxygen in normal oxides (e.g.  $\text{MgO}$ ) is normally -2, with the exception of peroxides (e.g.  $\text{H}_2\text{O}_2$ ) in which is -1.

How do we calculate the redox number for example of manganese in this chemical:  $\text{MnO}_4^-$ , permanganate. In order to do this, we need to set up a formula so that the redox numbers of all elements in the molecule—taking into account the number of atoms in the molecule—equals to the charge. In the case of permanganate, let us call  $x$  to the redox number of manganese. We know the redox of oxygen is -2 and we have four oxygens in the molecule. We also know the charge of the ion is -1. Therefore we have:

$$x + 4 \cdot (-2) = -1$$

If we solve for  $x$  we obtain a redox number of manganese of VII.

### Example

*Calculate the redox number of the elements underlined in the following molecules: (a)  $\text{K}_2\underline{\text{C}}\text{O}_3$  and (b)  $\text{H}_2\underline{\text{C}}\text{O}$ .*

**Answer:** Let us set up the redox equation for the first compound, knowing that the redox of oxygen is -2 and potassium +1. The unknown variable  $x$  represents the redox number of the underlined element. We have:

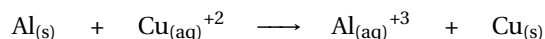
$$2 \cdot (+1) + x + 3 \cdot (-2) = 0$$

Mind we have two potassium and three oxygens hence we need to time the redox of K by two and the redox of O by three. If we solve for  $x$  we obtain a redox number for carbon of IV. The redox equation for the second example is:

$$2 \cdot (+1) + x + (-2) = 0$$

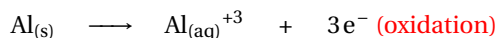
Mind that according to the redox rules, the redox number of oxygen is +1. Solving for  $x$  we have a redox number of zero.

By comparing the redox number of the same element in two different compounds we can figure out in what compound the element has lost or gained electrons. Look for example the case of  $\text{Cr}^{\text{VI}}_2\text{O}_7^{2-}$  and  $\text{Cr}^{\text{III}}_2\text{O}_3$ . The same element in two different molecules has two different redox numbers. In the case of dichromate, the redox of Cr is VI, whereas in the case of chromium(III) oxide the redox of Cr is III. The larger the redox number the more oxidized is an element, and that means the element has lost electrons. The smaller the redox number the more reduced is the element and that means it has gained electrons. If we compare both case, we have that Cr in dichromate is oxidized—it lost electrons—and Cr in chromium(III) oxide is reduced—it gained electrons. The goal is to identify the element that undergoes oxidation and reduction in a chemical reaction. We can reach this goal by using the half-reaction method. Every redox reaction is composed of two process, a reduction and the oxidation. These two processes can be separated into two half-reactions so that the combination of both half-reactions lead to the balanced redox. Let us work on a simple unbalanced redox reaction:

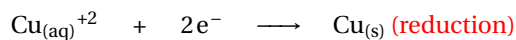


Solid  $\text{Al}_{(\text{s})}$  on the reactant side has zero redox number, whereas ionic  $\text{Al}_{(\text{aq})}^{+3}$  on the product side has redox number equal to III. Al has undergone oxidation as its redox number increases from zero to three. Al has lost three electrons. We can write the oxidation half-reaction:

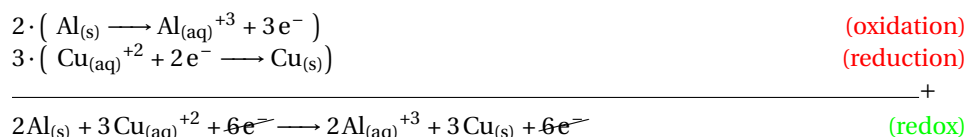




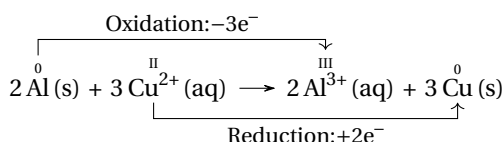
Mind that electrons have negative charge and we add electrons to compensate the charge of  $\text{Al}_{(\text{aq})}^{+3}$ . Now let us compare the redox number of Cu. In the reactant side we have  $\text{Cu}_{(\text{aq})}^{+2}$  with redox of II. In the product side we have  $\text{Cu}_{(\text{s})}$  with zero redox number. Cu has undergone reduction as its redox number has decreases. This means it has gained electrons, in particular two electrons:



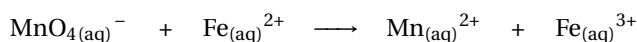
The goal here is to balance a redox chemical reaction by combining two half-reactions. In the example above the oxidation and reaction involve different number of electrons. Hence in order to be able to add both redox we need to time each half-reaction by a number so that the number of electrons cancel out. As the first reaction involved three electrons and the second two, we will do:



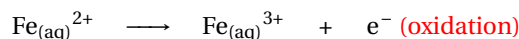
The overall balanced redox equation is:



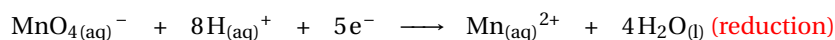
Redox reactions happen in either basic or acidic medium. Here we will go over how to balance redox reactions in acidic medium. In order to do this we will first separate the reaction in two half-reactions. In each semi-reaction we will balance all elements but hydrogen and oxygen. When all elements are balanced, we will proceed to balance O by adding  $\text{H}_2\text{O}$  molecules and we will balance H by adding  $\text{H}^{+}$ . Finally, we all add electrons to compensate the charge of the reaction. Let us work on an example:



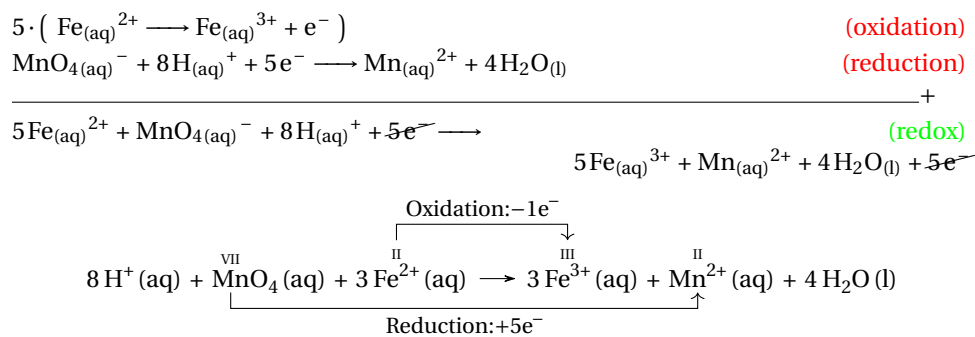
One of the semi-reactions involve Manganese whereas the other involves Iron. The redox number of Mn in permanganate is VII hence Manganese is being reduced, as its redox number decreases from VII to II, whereas Iron is being oxidized as its redox number increases from II to III. The oxidation half-reaction does not contain hydrogen or oxygen hence we will only have to balance the charger with one electron:



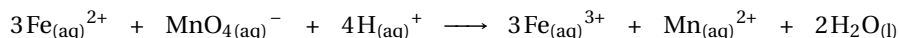
The reduction half-reaction contains oxygen. Hence, we will have to add  $\text{H}_2\text{O}$  molecules to balance oxygen and  $\text{H}^{+}$  to balance hydrogen. In particular, we will need two water molecules—as  $\text{MnO}_4^{-}$  has four oxygens and we will have to add eight protons as we are adding four molecules of water. Finally, we need to add three electrons to equalize the charge:



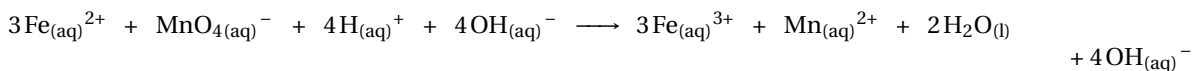
As the oxidation involves one electron and the reduction three, we need to time the oxidation by three:



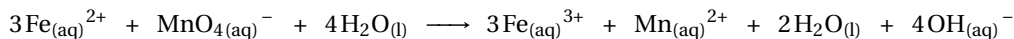
In order to balance a redox in basic medium we need first to balance the reaction in acidic medium. After, we will compensate all  $\text{H}^{+}$  with  $\text{OH}^{-}$  in both sides of the reaction. Mind that when combining  $\text{H}^{+}$  with  $\text{OH}^{-}$  we obtain  $\text{H}_2\text{O}$ . For example, in order to balance the following reaction in basic medium:



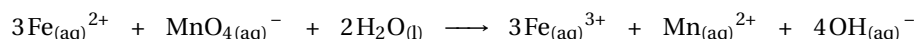
we will add four  $\text{OH}^{-}$  in both sides:



And after cancelling the four protons with the four hydroxyls, we have:



Now we have four water molecules in the left and two in the right. We will cancel two molecules from each side in order not to report water molecules twice:



### Example

*Balance the following redox in acidic medium:*

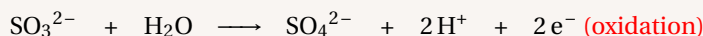


#### Answer:

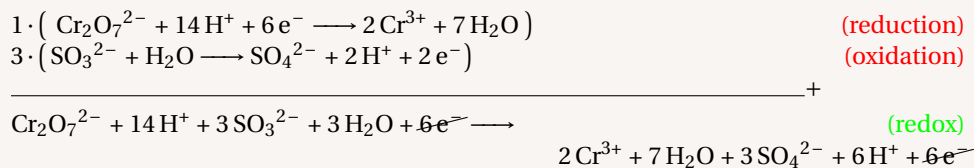
We locate the same element in both sides of the reaction with different redox number. We found Chromium in the form of dichromate ( $\text{Cr}_2\text{O}_7^{2-}$ ) with redox number VI and Chromium in the product side with redox III. Therefore Chromium is being reduced, as its redox number decreases. Differently, we found sulphur in the reactants side in the form of sulfite ( $\text{SO}_3^{2-}$ ) with redox number of IV and in the product side with redox number of VI. Therefore Sulphur is being oxidized. We will first set up the oxidation half-reaction knowing that we have different amounts of Cr in both side and that we will have to add water molecules to balance O and protons to balance H. As we have seven oxygens in  $\text{Cr}_2\text{O}_7^{2-}$  we will have to add seven water molecules. Also, as we add seven water molecules we will have to add fourteen protons. We will need six electrons to compensate the charge:



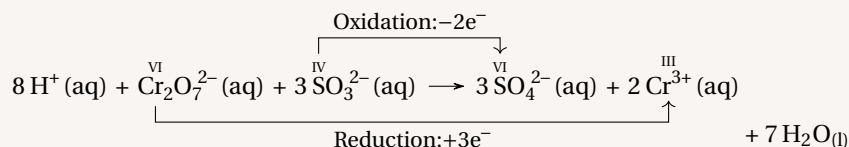
For the reduction half-reaction, we have a difference of one oxygen atoms and hence we will need one water molecule and two protons; we will need two electrons to compensate the charge:



In order to add both half-reactions as the reduction involves two electrons and the oxidation six, we will have to multiply the reduction half-reaction by three:



After we cancel the electrons and eliminate protons and water molecules, the balanced reaction is:



### The experiment

The experimental part consists of a trial titration and an exact titration. The trial titration will be used to estimate the volume added at the endpoint of the titration. It is conducted quickly, adding portions of 1 mL until the endpoint is exceeded. The

final result is not used for calculations. In the exact titration, an initial amount is added rapidly. This amount is calculated by subtracting 5 mL from the volume used in the trial titration. Like this it is possible to stop before the endpoint of the titration and to continue carefully, adding the titrant on a drop-by-drop basis until the precise endpoint. If time allows, the exact titration should be repeated and the mean volume used in calculations.




## Procedure

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### Part A. Trial titration.

- ☐ *Step 1:* – Obtain a solution of bleach (the instructor might give you an unknown bleach sample, if so write down the Unknown # in the results section).
- ☐ *Step 2:* – Obtain a buret, a stand, and a clamp. Rinse the buret with distilled water and make sure the valve works properly.
- ☐ *Step 3:* – In a 100 mL beaker, obtain 70 mL of 0.025 M solution of  $\text{Na}_2\text{S}_2\text{O}_3$ . Use a small portion of the solution to rinse the buret. Discard the rinsing liquid and fill up the buret above the zero mark. Bring the volume in the buret to zero or below, discarding the excess liquid through the tip of the buret.
- ☐ *Step 4:* – In a 125 mL Erlenmeyer, obtain about one inch of a spatula of KI.

#### CAUTION!

-  Bleach is a base.
-  Bases and acids can cause burns and clothes color loss.
-  Always wear eye protection and handle carefully.

- ☐ *Step 5:* – In a 50 mL beaker, obtain about 30 mL of a 10% bleach solution.
- ☐ *Step 6:* – Obtain a 10 mL transfer pipet. Pipet 10.0 mL of the bleach into the KI containing Erlenmeyer. Add 20 mL of distilled water from a graduated column and 20 drops of HCl 2 M. Homogenize the mixture by stirring the flask.
- ☐ *Step 7:* – Record the buret initial reading to the closest hundredth of mL.
- ☐ *Step 8:* – Place the Erlenmeyer under the buret and start the titration by adding 1 mL portions until the solution turns from brown to yellow.
- ☐ *Step 9:* – Add 40 drops of a 0.2% starch solution to the yellow solution in the Erlenmeyer.
- ☐ *Step 10:* – Proceed with the titration adding the  $\text{Na}_2\text{S}_2\text{O}_3$  in 1 mL portions until it turns from dark color to colorless.
- ☐ *Step 11:* – Record the buret final reading and calculate the volume added.
- ☐ *Step 12:* – Discard the solution in the Erlenmeyer in the corresponding waste container.

### Part B. Exact titration.

- ☐ *Step 1:* – Refill the buret with  $\text{Na}_2\text{S}_2\text{O}_3$  0.025 M. At this point you know how much volume you will need from the trial titration. You don not need to fill the buret up to zero but add a few milliliters more than the expected volume.
- ☐ *Step 2:* – In a 125 mL Erlenmeyer, obtain about one inch of a spatula of KI.
- ☐ *Step 3:* – Pipet 10.0 mL of the diluted bleach into the KI containing Erlenmeyer. Add 20 mL of distilled water from a graduated column and 20 drops of HCl 2 M. Homogenize the mixture by stirring the flask.

- ☐ *Step 4:* – Record the buret initial reading to the closest hundredth of mL.
- ☐ *Step 5:* – Place the Erlenmeyer under the buret and add 3 mL less than the volume used in the trial titration.
- ☐ *Step 6:* – Add 40 drops of a 0.2% starch solution to the solution in the Erlenmeyer.
- ☐ *Step 7:* – Continue the titration adding the titrant drop by drop until the solution turns transparent.
- ☐ *Step 8:* – Record the buret final reading in your notes.
- ☐ *Step 9:* – Discard the solution in the Erlenmeyer. If the time allows, do two exact titration replicates.
- ☐ *Step 10:* – When finished empty the buret and rinse it with plenty of distilled water. The buret should be stored empty, open, and upside down.

## Calculations

---

- ① This is the initial reading of the buret (often is zero if the buret is all filled up).
- ② This is the final reading of the buret after the endpoint has been reached.
- ③ This is the volume of titrant used,  $\nu_{\text{S}_2\text{O}_3^{-2}}$ :

$$\nu_{\text{S}_2\text{O}_3^{-2}} = \textcircled{2} - \textcircled{1}$$

- ④ This is the moles of thiosulfate used in the titration:

$$n_{\text{S}_2\text{O}_3^{-2}} = \textcircled{3} \times 0.025 \times 10^{-3}$$

- ⑤ These are the moles of bleach in the (diluted) sample:

$$n_{\text{OCl}^{-1}} = \frac{n_{\text{S}_2\text{O}_3^{-2}}}{2} = \frac{\textcircled{4}}{2}$$

- ⑥ This is the molarity of bleach in the (diluted) sample:

$$c_{\text{OCl}^{-1}}^{\text{diluted}} = \frac{n_{\text{OCl}^{-1}}}{1 \times 10^{-2}} = \frac{\textcircled{5}}{1 \times 10^{-2}}$$

- ⑦ This is the molarity of bleach in the (original, concentrated) sample:

$$c_{\text{OCl}^{-1}}^{\text{concentrated}} = c_{\text{OCl}^{-1}}^{\text{diluted}} \times 10 = \textcircled{6} \times 10$$



**STUDENT INFO**

Name:

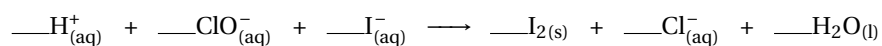
Date:

**Pre-lab Questions**

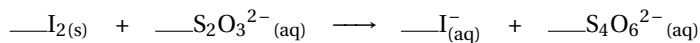
# Strength of Laundry Bleach

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1. Balance the following redox reaction, indicating the oxidizing and reducing agents:



2. Balance the following redox reaction, indicating the oxidizing and reducing agents:



3. What is the proportion of bleach to water necessary to make water safe to drink?
4. What is the active ingredient in liquid bleach? And the active ingredient in bleach powder?
5. Indicate the oxidation state of all elements in the following species:

×  $\text{H}^{+}$

×  $\text{H}_2\text{O}$

×  $\text{Cl}^{-}$

×  $\text{ClO}^{-}$

×  $\text{I}^{-}$

×  $\text{I}_2$

×  $\text{S}_2\text{O}_3^{2-}$

×  $\text{S}_4\text{O}_6^{2-}$



**STUDENT INFO**

Name:

Date:

**Results  
EXPERIMENT****Strength of Laundry Bleach**

		Trial		Exact	
		1	2	3	
1	Initial buret reading (mL)	_____	_____	_____	_____
2	Final buret reading (mL)	_____	_____	_____	_____
3	$\text{Na}_2\text{S}_2\text{O}_3$ added (mL)	_____	_____	_____	_____
4	Moles $\text{Na}_2\text{S}_2\text{O}_3$	_____	_____	_____	_____
5	Moles NaOCl	_____	_____	_____	_____
6	[NaOCl] (diluted) (M)	_____	_____	_____	_____
7	[NaOCl] (original) (M)	_____	_____	_____	_____
Average Molarity,[NaOCl] (M)			_____		
Unknown number ( if given )			_____		





**STUDENT INFO**

Name:

Date:

**Post-lab Questions**

# Strength of Laundry Bleach

---

1. Commercial bleach claims a 5.25% NaClO by weight. Assuming a density of 1.0 g/mL for NaClO obtain the molarity of the solution. Consider your results in the laboratory, is the claim true?
2. Considering the bleach-water ratio of eight bleach drops per water gallon, what is the concentration, in molarity, of NaClO in the dilution?



## EXPERIMENT

# Acids Strength

## Goal

---

The goal of this laboratory is to learn three different techniques to measure pH.

## Materials

---

- |                                                    |                                                                                                                                                                                                   |
|----------------------------------------------------|---------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|
| <input type="checkbox"/> 5 jumbo testtube and rack | <input type="checkbox"/> 0.1 M $\text{H}_3\text{PO}_4$ , 0.1 M $\text{CH}_3\text{COOH}$ , 0.1 M $\text{NaH}_2\text{PO}_4$ , 0.1 M $\text{Zn}(\text{NO}_3)_2$ , and 0.1 M $\text{NH}_4\text{NO}_3$ |
| <input type="checkbox"/> 100mL, 50mL beaker        | <input type="checkbox"/> Distilled water and tap water                                                                                                                                            |
| <input type="checkbox"/> Indicator solutions       | <input type="checkbox"/> Unknown solution                                                                                                                                                         |

## Background

---

Acids are considered dangerous substances and while it is true that special care must be taken when handling some of these products, many acids are used in daily life with great simplicity. For example, aspirin's active principle ingredient is acetylsalicylic acid, vinegar is a 5% solution of acetic acid, and vitamin C is ascorbic acid.

Acids are defined by Brønsted-Lowry as proton ( $\text{H}^+$ ) donors because when dissolved in water, acids dissociate by releasing protons. Strong acids dissociate completely into two ions: the anion is the acid conjugate base and the cation is a proton. For stoichiometric reasons, the concentration of each ion in the solution equals the initial concentration of the acid. Weak acids do not dissociate completely but, instead, only do so to a certain extent. The strength of weak acids depends on how much they dissociate, a ratio characterized by the acid dissociation constant,  $K_a$ .

For a certain weak acid



The equilibrium can be characterized by the expression:

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

where  $K_a$  (capital letter) is the dissociation constant and the brackets represent the concentrations of the respective species.

## pH

The acidity depends on the concentration of  $\text{H}^+$  in the solution, which will always be a value between  $10^{-14}$  and 1. To avoid using such a large and complex range, a logarithmic scale is preferred, ranging from 0 to 14.

This scale is called the pH, defined by Danish biochemist Sørensen in 1909 as the "power of Hydrogen".

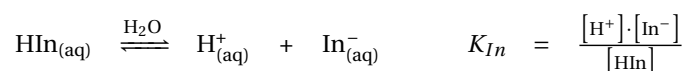
$$\text{pH} = -\log[\text{H}^+]$$

Considering that the strength of acids is their power to generate  $\text{H}^+$ , and to compare the strength of different acids, pH should be measured.

In this experiment, the pH for different solutions will be measured using three different approaches; indicators, pH paper, and a pH meter. Each of these techniques has a different accuracy. While using **pH indicators** it is only possible to determine a range for the pH, **pH paper** will provide an estimated unit, and the **pH-meter** will read down to the decimals.

## Indicators

There are several techniques to locate the endpoint during a titration. On one hand, we can use a PH meter and measure PH during the titration. The endpoint locates between the initial and final plateau. We can also use an automatic titrator, a machine that automatically mixes the acid and the base while displaying the PH. Another technique uses an indicator that changes color when the PH changes from acidic to basic or the opposite. Acid-base indicators are water-soluble organic dyes that have a different color depending on PH. Halochromism is the effect of changing color due to changes on a external force such as PH. Many flowers or plants exhibit this quality. For example, the french hydrangea is a plant that changes color depending on the acidity of the soil. Turmeric is a specie that turn from yellow to brown when the PH increases. indicator can be extracted from many plants and fruits such as red cabbage, geranium, poppy, rose petals or blueberries. Indeed, indicators are weak acids that exist in a conjugate base form as well. Both the acid (the indicator form with the hydrogen attached) and the conjugate base have a different color due to differences in their electronic structure that transform into differences absorbing light. For example, phenolphthalein is a common acid-base indicator that is transparent when the PH is acidic and pink when the PH is basic. The equivalency point of a titration (when acid and base are neutralized) is not necessarily the end point (when the indicator changes color). However, if both points are close enough chemical indicators are an accurate way to visually locate the equivalency point during a titration. As indicators are weak acids, each particular indicator is characterized by a constant of acidity referred to as  $K_{In}$  so that








The change in color is considered to happen when the difference between the acidic and basic form of the indicator is an order or magniture–ten times–different. This is the ten times rule. For this difference, we have that

$$\text{PH} = \text{PK}_{In} \pm 1$$

We can always use a different value for the difference ( $\Delta$ ) between the acid and basic form of an indicator so that the PH range will be given by

$$\text{PH} = \text{PK}_{In} \pm \log(\Delta)$$

This expression helps with the selection of the indicator for a given titration. For example, if we want to select an indicator that changes color when the difference between acid and base form is 5 times ( $\Delta = 5$ ), then the OH range would be:  $\text{PH} = \text{PK}_{In} \pm 0.7$ . A good indicator will turn into a different color within the transition between acid and basic PH during a titration. For example, for the titration of hydrochloric acid with sodium hydroxide, the PH changes from 5 before the equivalency point to 8 after. Any indicator with  $\text{PK}_{In}$  between 5 and 8 would be able to accurately indicate the stoichiometric point. Examples are bromothymol blue, m-nitrophenol, and even phenolphthalein that transition color at 8. For the titration of a weak acid and a strong base, the selection of an indicator is more difficult as the difference between the acidic and basic plateaus during the titration is less pronounced. We must choose an indicator that changes color more closely to the end point. Litmus paper is an indicator heavily used for quick PH testing in labs. It contains a mixture of dyes that were extracted originally from lichens. There are two main types of litmus paper, red and blue litmus. Light blue litmus paper turns red when the medium is acidic, whereas red litmus paper turns blue in basic mediums. By wetting a piece of litmus paper one can also test for the acidity of gases, and for example, ammonia vapor would turn blue a piece of wet red litmus paper.

Acid-base indicators showing ample representations of their colors				
Name	PH Range	Color acid form		Color basic form
Thymol blue (acid range)	1.2-2.8	red		yellow
Thymol blue (basic range)	8.0-9.6	yellow		blue
Methyl orange	3.1-4.4	red		orange
Methyl red	4.4-6.2	red		yellow
Bromomethyl blue	6.0-7.6	yellow		blue

## The experiment

The experiment today is divided into the following parts:

- Determining pH range using indicators.
- Determining pH using pH paper.
- Determining pH using a pH meter.

## Part A. Determining pH range using indicators.

A pH indicator is a chemical substance that, when added to a solution, changes its color depending on the solution's pH. The pH at which color changes is called the **equivalence point**. Some indicators have one equivalence point while others have more than one equivalence points. By using an indicator it will be possible to determine a pH range only. Combining different indicators the range can be narrowed down.

### Example

*We test a sample with three different indicators by dividing the sample into three small samples. Each small sample is tested using an indicator from the following list: i) Indicator A is colorless for pH values lower than 10 and turns pink for a pH greater than 10. ii) Indicator B is blue up to pH 3, then it turns colorless. iii) Indicator C is colorless for a pH lower than 6, when it turns red. The red color changes to yellow between pH 9 and pH 10, and stays yellow for higher pH values. Estimate the pH range for the solution if the three samples are colorless.*

**Answer:** Indicator A tells that the solution's pH is lower than 10. Therefore, the range must be  $0 < \text{pH} < 10$ . Indicator B shows that the solution's pH is greater than 3 (3 to 14). This information, combined with that of indicator A, narrows the pH range to  $3 < \text{pH} < 10$ . Indicator C tells that the solution's pH is lower than 5. The final range is  $3 < \text{pH} < 5$ .

## Parts B. Determining pH using pH-paper.

pH paper, or **universal pH paper** is a strip of paper impregnated with a mixture of indicators. The paper changes color smoothly for each pH value and the pH is determined by matching the color of the paper with that in a chart attached to the pH paper. In the best case, by using pH paper it will be possible to select one value for the pH, although in some situations when the color is not one in the chart, it will be possible to indicate a range of two numbers only.

## Parts C. Determining pH using a pH-meter.

pH meters are scientific instruments that measure the pH of solutions. The tool has a probe that is inserted in the liquid and measures the concentration of protons based on the voltage between two electrodes. Their use is similar to that of a conductivity-meter, and some instruments can perform both tasks. Probes need to be cleaned after each measurement to avoid contamination and should never be touched with the hands or paper towels. Calibration of the tool with solutions of known pH ensures the best accuracy. After use, probes are best kept moist, typically in distilled water. A pH meter can offer a precision down to the hundredth of the unit of pH.

## Procedure

### Part A. Determining pH range using indicators.

- ☐ *Step 1:* – Obtain 5 jumbo test tubes and a rack. Clean the test tubes and let them drain upside down in the rack.
- ☐ *Step 2:* – Obtain a 100 mL beaker with plenty of distilled water and transfer 4.0 mL in a graduated cylinder.
- ☐ *Step 3:* – Label one test tube with the letter R, and pour the 4.0 mL of distilled water into it. This test tube will be used as a volume reference for the following test tubes.
- ☐ *Step 4:* – In a 50 mL beaker get 40 mL of one of the following list of solutions (you will use the others further in the experiment). These solutions are going to be referred to here as the acid or base solutions.

- |                                   |                                     |                                   |                     |
|-----------------------------------|-------------------------------------|-----------------------------------|---------------------|
| 1. 0.1 M $\text{H}_3\text{PO}_4$  | 3. 0.1 M $\text{NaH}_2\text{PO}_4$  | 5. 0.1 M $\text{NH}_4\text{NO}_3$ | 7. Tap water        |
| 2. 0.1 M $\text{CH}_3\text{COOH}$ | 4. 0.1 M $\text{Zn}(\text{NO}_3)_2$ | 6. Distilled water                | 8. Unknown solution |

**⚠ CAUTION!**

- ⚠ Handle concentrated acids with care to avoid chemical burns.
- ⚠ Are you wearing your goggles? Do you know where the eye-washer is?

- ☐ *Step 5:* – Obtain a 250 mL beaker for waste, a medicine dropper or a plastic pipette
- ☐ *Step 6:* – Label the rest of the test tubes from 1 to 4. Use test tube R as a reference for the volume to add in the following test tubes. Rinse the dropper with the distilled water (discarding it in the waste beaker) between different solutions. Add 4.0 mL of the solution chosen to each test tube.
- ☐ *Step 7:* – Add two drops of Thymol Blue to test tube 1. Using the color table provided in the background section, indicate the PH range of the solution.
- ☐ *Step 8:* – Add two drops of Methyl Orange to test tube 2. Using the color table provided in the background section, indicate the PH range of the solution.
- ☐ *Step 9:* – Add two drops of Methyl Red to test tube 3. Using the color table provided in the background section, indicate the PH range of the solution.
- ☐ *Step 10:* – Add two drops of Bromothymol Blue to test tube 4. Using the color table provided in the background section, indicate the PH range of the solution.
- ☐ *Step 11:* – Record the colors on the results page.
- ☐ *Step 12:* – Reserve the acid or base solution for Parts B and C. Set the beaker on a paper towel and label it.
- ☐ *Step 13:* – Discard the solutions tested with the indicator and clean the test tubes.
- ☐ *Step 14:* – Repeat steps 4 to 12 for another acid or base solution from the list given in Step 4. Make sure to rinse the dropper.

**Part B. Determining pH using pH-paper.**

- ☐ *Step 15:* – Obtain a stirring rod and the universal indicator paper. Be careful not to contaminate the paper with dirty gloves or bare hands. Always set the (small) pieces of paper on a clean paper towel.
- ☐ *Step 16:* – Rinse the stirring rod with the wash bottle over the waste beaker. Rinse thoroughly between different solutions.
- ☐ *Step 17:* – Immerse the tip of the stirring rod in one solution and touch a piece of pH paper with the rod.
- ☐ *Step 18:* – Match the color of the drop on the paper with the color chart on the plastic case. Record the pH value in the results section.
- ☐ *Step 19:* – Repeat the same process; rinsing, dipping, and touching, for the rest of the solutions.
- ☐ *Step 20:* – Reserve the solutions in the beakers for part C.

**Part C. Determining pH using a pH-meter.**

- ☐ *Step 21:* – Install the pH meter. The probe of the instrument is immersed in a clean solution that must be kept aside until the end of the experiment.
- ☐ *Step 22:* – Cleaning and measuring. It is very important to follow the cleaning steps before each measurement. You will need a 100 mL beaker labeled waste a wash bottle and a 100 mL beaker with distilled water.
  - 1. Rinse the probe using the wash bottle over the waste beaker.
  - 2. Dip the probe 3-4 times into the 250 mL beaker with clean distilled water.
  - 3. Dip the probe into the solution to be measured. Use a 50 mL beaker with at least 25 mL of solution.
  - 4. Repeat cleaning steps 1 and 2.
- ☐ *Step 23:* – Following the procedure outlined in the step above, measure the eight solutions.
- ☐ *Step 24:* – Record all the values.





**STUDENT INFO**

Name:

Date:

**Pre-lab Questions**

# Acids Strength

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1. Find the definition of the following concepts:
  - (a) Acids (according to Brønsted-Lowry definition).
  - (b) pH.
  - (c) Acid dissociation constant.
  - (d) Equivalence point.
2. Name the three methods to determine pH to be used in this laboratory.
3. Order the three methods from the most accurate to the least accurate, according to your expectations.
4. We test a sample with three different indicators by dividing the sample into three small samples. Each small sample is tested using an indicator from the following list: i) Indicator A is colorless for pH values lower than 10 and turns pink for a pH greater than 10. ii) Indicator B is blue up to pH 3, then it turns colorless. iii) Indicator C is colorless for a pH lower than 6 when it turns red. The red color changes to yellow between pH 9 and pH 10, and stays yellow for higher pH values. Estimate the pH range for the solution if the samples tested with indicators A and B become colorless and the one tested with indicator C turns reddish.



STUDENT INFO

Name: Date:

Date:

# Acids Strength

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**STUDENT INFO**

Name:

Date:

**Post-lab Questions****Acids Strength**

1. Use the following indicator's color chart to determine the pH range for the solutions.

Solution	pH-range	Solution	pH-range
1. $\text{H}_3\text{PO}_4$	_____	5. $\text{NH}_4\text{NO}_3$	_____
2. $\text{CH}_3\text{COOH}$	_____	6. Distilled water	_____
3. $\text{NaH}_2\text{PO}_4$	_____	7. Tap water	_____
4. $\text{Zn}(\text{NO}_3)_2$	_____	8. Unknown	_____

2. Identify the unknown.

3. Which of the three techniques is the most accurate?

4. Calculate the concentration of  $\text{H}_3\text{O}^+$  ions in the most acidic solution, using the pH-meter value.

5. Calculate the concentration of  $\text{H}_3\text{O}^+$  ions in the less acidic solution, using the pH-meter value.

