

LAB MANUAL FOR GENERAL, ORGANIC AND BIOLOGICAL CHEMISTRY-1-SEMESTER



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GOB Lab Manual

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1: PHYSICAL AND CHEMICAL PROPERTIES

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1.1: Student Learning Outcomes

PDF file of Experiment 1

Upon completion of this experiment, the student will be able to:

- evaluate the difference between a physical property or a chemical property
- *measure mass of an object or a quantity of a granular solid using a milligram balance*
- *measure temperatures using a thermometer*
- *determine volume of a liquid using a graduated cylinder and the volume of an irregular solid using the water displacement method*
- *perform length measurements using a ruler*
- *express all measurements to the correct number of significant figures*
- discuss similarities in chemical properties for groups of elements in the periodic table
- recognize the influence of impurities on the melting (freezing) point of water

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1.2: Materials and Discussion

Materials

Sodium chloride, balance, ruler, 10 mL and 100 mL graduated cylinders, 400 mL and 1000 mL beakers, thermometer, stirrer, wash bottle, ruler, tape, scissors, evaporating dish, thermometer, rubber stopper, baking soda, vinegar, 0.1 M sodium chloride, 0.1 M silver nitrate, 10% sodium hydroxide, 6 M hydrochloric acid, sucrose, solution A & B, small and large test tubes, 0.5 M solutions of BaCl_2 , CaCl_2 , SrCl_2 , LiCl , KCl , NaCl , 0.5 M $(\text{NH}_4)_2\text{CO}_3$ solution and 0.1 M solutions of NaCl , NaBr , and NaI .

Discussion

Physical vs. Chemical Property

A physical property refers to a characteristic that can be observed without changing the composition of the substance. Some of the examples of physical properties are mass, volume, color, density, melting point, boiling point, solubility in water and physical states of matter such as solid, liquid or gas.

A chemical property refers to a property that can be observed during a chemical reaction. As a result of a chemical change new substances are produced. A chemical change is accompanied by the appearance of new sets of properties corresponding to formation of new substances. Some examples of chemical changes are burning of wood, rusting of iron, and boiling of an egg. Observations during a chemical change include release of a gas, a permanent change of color, formation of an insoluble substance (a precipitate) when two solutions are added together, and release of energy. The release in energy is often manifested in terms of heat, light or sound.

Precision and Accuracy of Measurements

Scientific measurements must be as precise as possible. This means that every measurement will include one uncertain or estimated digit. When making measurements we normally estimate at the smallest scale division on the instrument being used. Only the uncertain digit should vary if the measurement is repeated using the same instrument, even if it is repeated by someone else.

The accuracy of a measurement or calculated quantity refers to its agreement with some known value. For example, to determine the density of a substance, we need to make two measurements, mass and volume. The ratio of the mass to the volume will give us the experimental density of the substance. This experimental density can then be compared with the density of the substance listed in a reference such as the *Handbook of Chemistry and Physics*. High accuracy means there is good agreement between the experimental value and the known value listed in the reference.

Mass Measurements

In conducting an experiment, it is often unnecessary to measure an exact quantity of material. For instance, the directions might state, "Weigh about 5 g of sodium chloride." This instruction indicates that the measured quantity of salt should be 5 g plus or minus a small quantity. In this example 4.8 to 5.2 g will satisfy these requirements. To weigh exactly 5.000 g wastes time since the directions call for approximately 5 g.

Sometimes it is necessary to measure an amount of material precisely within a stated quantity range. Suppose the directions read, "Weigh about 5 g of sodium chloride to the nearest 0.001 g with a milligram balance." This instruction does not imply that the amount is 5.000 g but that it should be between 4.8 and 5.2 g and measured and recorded to three decimal places. Therefore, four different students might weigh their samples and obtain 5.141 g, 5.034 g, 4.812 g, and 4.937 g, respectively, and each would have satisfactorily followed the directions.

An analytical balance may also be used in the laboratory. The mass should be measured and recorded to four decimal places if a student uses an analytical balance. Therefore, four different students might weigh their samples and obtain 5.1412 g, 5.0341 g, 4.8123 g, and 4.9371 g, respectively, and each would have satisfactorily followed the directions.

Temperature Measurements

The simple act of measuring a temperature with a thermometer can easily involve errors. Not only does the calibration of the scale

on the thermometer limit the precision of the measurement, but the improper placement of the thermometer bulb in the material being measured introduces a common source of human error. When measuring the temperature of a liquid, one can minimize this type of error by observing the following procedures:

1. Hold the thermometer away from the walls of the container.
2. Allow sufficient time for the thermometer to reach equilibrium with the liquid.
3. Read the thermometer at eye level.
4. If two substances are mixed before the temperature is measured of the mixture, be sure the mixture is adequately mixed.

Freezing Points and Boiling Points of Pure and Impure Substances

The melting point and boiling point are easily determined physical properties that are very useful in identifying a pure substance. When a substance (solvent) is uniformly mixed with a small amount of another substance (solute), the freezing point (melting point) of the resulting solution (an "impure substance") will be lower than that of the pure solvent. For example, the accepted freezing point for pure water is 0.0°C . Solutions of salt in water may freeze at temperatures as low as -21°C depending on the amount of salt added to the water. That is why salt is spread on snowy walkways to melt the snow in winter and antifreeze (ethylene glycol) is added to the water in a car radiator to lower the freezing point of the water. This phenomenon of a lowered freezing point of a mixture of substances is known as freezing point depression.

The addition of ethylene glycol to the water in your auto's radiator coolant not only lowers the freezing point of the water solution so it does not freeze in the winter, but it also elevates the boiling point. This helps prevent the radiator coolant from boiling over in the summer. This phenomenon of elevated boiling point of a mixture of substances is known as boiling point elevation.

Length Measurements

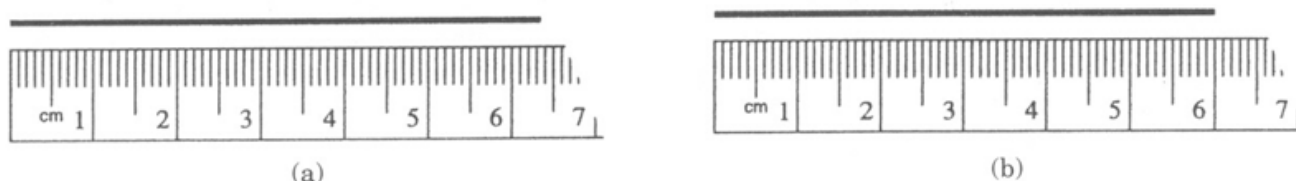


Figure 1.1

Suppose we are measuring two pieces of wire, using the metric scale on a ruler that is calibrated in tenths of centimeters as shown in Figures 1.1 a and b. One end of the first wire is placed at exactly 0.0 cm and the other end falls somewhere between 6.3 cm and 6.4 cm . Since the distance between 6.3 and 6.4 is very small, it is difficult to determine the next digit exactly. One person might estimate the length of the wire as 6.34 cm and another as 6.33 cm . The estimated digit is never ignored because it tells us that the ruler can be read to the 0.01 place. This measurement therefore has three significant digits (two certain and one uncertain figure).

The second wire has a length which measures exactly 6 cm on the ruler as shown in Figure 1.1 b. Reporting this length as 6 cm would be a mistake for it would imply that the 6 is an uncertain digit and others might record 5 or 7 as the measurement. Recording the measurement as 6.0 would also be incorrect because it implies that the 0 is uncertain and that someone else might estimate the length as 6.1 or 5.9 . What we really mean is that, as closely as we can read it, the length is exactly 6 cm . So, we must write the number in such a way that it tells how precisely we can read it. In this example we can estimate to 0.01 cm so the length should be reported as 6.00 cm .

Volume and Meniscus

The curved surface of a liquid (water) is called the meniscus. A meniscus forms because the water molecules are more strongly attracted to the container than to each other. To properly measure the volume of a liquid in a graduated cylinder you must be at eye-level and read the bottom point of the meniscus.

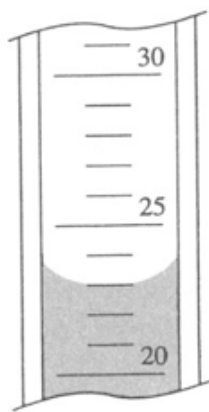


Figure 1.2 **Read the bottom of the meniscus.** The volume is 23.0 mL

Volumes measured in illustrated graduated cylinders are calibrated in 1 ml increments and should be estimated and recorded to the nearest 0.1 ml.

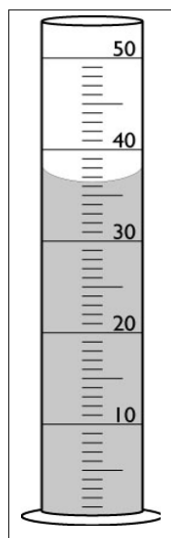


Figure 1.3: The volume is 36.5 mL

1.3: Pre-Laboratory Assignment

Name: -----

Pre-laboratory Assignment # 1 (Due prior to start of lab session)

Please read the discussions and procedure of this experiment to answer the following questions.

1. What is physical property. List two physical properties you **will measure** in this experiment.
2. What is a meniscus? Explain the proper technique for reading a meniscus.
3. In the laboratory you may be provided with a milligram balance or an analytical balance. How do you identify a balance as milligram or analytical?
4. What is the uncertain digit in a measurement? Identify the uncertain digit in the volume measurement of 23.4 mL?
5. What is chemical property. List two examples of chemical properties.
6. List four experimental observations that suggest a chemical change has taken place.

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1.4: PROCEDURE AND OBSERVATIONS

1.4.1: Mass (Weight)

1.4.2: Temperature

1.4.3: Volume of a Rubber Stopper

1.4.4: Length

1.4.5: Chemical Properties vs. Physical Properties

1.4.6: Families of Elements

1.4.1: Mass (Weight)

Procedure and Observations

A. Mass (Weight)

The mass measurements may be performed with a 0.001gram precision balance (milligram balance) or a 0.0001 grams precision balance (analytical balance).

The balance should always be “zeroed” before anything is placed on the balance pan. On an electronic digital balance, this is done with the “tare” or “T” button.

Never place chemicals directly on the balance pan; place them on a porcelain container (evaporating dish) or watch glass outside of the balance prior to mass measurement. Clean up any materials you spill around the balance with a brush.

1. Note the balance provided, as milligram balance or analytical balance. If the reading is 0.000 g on “tare” then you are using a milligram balance. If the reading in 0.0000 g on “tare”, then you have an analytical balance.
2. Do the following, recording all the masses in the next three steps to the nearest 0.001 g with a milligram balance or 0.0001 g with an analytical balance. Since you are provided with a digital balance, record your measurements, so that all significant digits are measured.
3. Request a solid rubber stopper from your instructor. Weigh the dry stopper. **Save the stopper for part C of the experiment.**
4. Weigh a small evaporating dish (porcelain dish).
5. Add approximately 5 g of sodium chloride to the evaporating dish from step 4. Measure to the nearest 0.001 g on a milligram balance or 0.0001 g on an analytical balance.

Note: Add the salt in small quantities with a spoon outside of the balance to reach approximately 5 g. Do not add the salt to the container while it sits in the balance. Clean up any materials you spill around the balance.

6. Calculate the mass of sodium chloride by difference.

Observations:

Balance used (Circle one): Milligram Balance / Analytical Balance

Object or chemicals	Mass measurements in units of grams (g)
Rubber stopper	
Evaporating dish	
Evaporating dish + sodium chloride	
sodium chloride	

Save the rubber stopper for part C and the salt for part B. Do not dispose.

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1.4.2: Temperature

B. Temperature

Record all temperatures to the **nearest 0.1°C** to include one uncertain digit and all certain digits.

1. Fill a 400 mL beaker half full of tap water. Place your thermometer in the beaker. Give it a minute to reach thermal equilibrium. Keeping the thermometer in the water and holding the tip of the thermometer away from the glass, read and record the temperature in the Celsius scale. Save this water for the next step.

2. Fill a 1000 mL beaker with crushed ice from the ice-machine. Add 200 mL of crushed ice with a beaker to the water in the 400 mL beaker from B1. Stir the mixture for about 2 minutes with a glass rod. Add more ice if almost all the ice melts. Holding the thermometer and the beaker at eye level, read and record the temperature of ice water in the Celsius scale. Save the ice-water for the next step.

3. Add the 5 g of sodium chloride from part A to the ice-water mixture. Stir for 2 minutes with a glass rod, to dissolve the salt adequately in the water. Add more ice if needed (i.e., if the ice melts). Read and record the temperature of the mixture in the Celsius scale.

4. Your instructor will fill a 500 mL beaker half full of tap water and warm the water on a hot plate. Read and record the temperature of the warm water. Be sure to hold the thermometer away from the bottom and sides of the beaker.

5. Salt water/ice mixture may be disposed down the sink.

Observations

Sample	Measured temperature to 0.1°C
Tap water	
Ice water	
Salt + ice water	
Warm water	

Compare the temperature of ice water and ice water with salt added. Which sample has a lower temperature? What is this phenomenon called?

Convert the measured temperature of warm water to the Kelvin scale. Pay attention to significant figures.

1.4.3: Volume of a Rubber Stopper

C. Volume of a Rubber Stopper (Water Displacement Method)

1. Fill the 100 mL graduated cylinder with de-ionized water to approximately 50 ml using the wash bottle provided in the locker. **Read the volume at the bottom of the lower meniscus.** Record the volume to the 0.1 place.
2. Carefully place the rubber stopper from part A into the graduated cylinder so that it is submerged. Read, record the new volume to 0.1 place.
3. Calculate the volume of the rubber stopper by difference.

Observations

Read the volume at the bottom of the lower meniscus. Record the volume to the 0.1 place.

Initial volume of water in cylinder -----mL

Final volume of water in cylinder -----mL

Volume of rubber stopper -----mL

Return the rubber stopper to your instructor.

1.4.4: Length

D. Length

1. Add 2.0 mL of deionized water to a 10.0 mL graduated cylinder using a wash bottle. The bottom of the lower meniscus should be at 2.0 ml mark. Pour the liquid into a **large test tube**. Mark the height of water in the test tube with a small piece of tape. With a ruler, measure the height of water in the test tube in cm.
2. Add 2.0 mL of water in a 10.0 mL graduated cylinder. Pour it into a **small test tube**. Mark the height of water in the test tube with a small piece of tape and with a ruler, measure the height in cm.
3. In the future, you will often find it convenient to simply estimate volumes of 1.0 to 2.0 mL simply by observing the height of the liquid in the test tube. So, save the little pieces of tape on the test tube and have your instructor sign off your report.

Observations:

Read the ruler to the nearest 0.01 cm. The length measurements should have three significant digits: two certain and one uncertain digit.

Height of 2.0 mL of water in a **large** test tube -----cm

Height of 2.0 mL of water in a **small** test tube -----cm

Instructor signature -----

1.4.5: Chemical Properties vs. Physical Properties

E. Chemical Properties vs. Physical Properties

Use five large clean test tubes, one for each test. Observe carefully any changes which occur in tests 1 through 5 below. On the basis of these observations classify each result as a physical or a chemical change.

- To 1 ml of vinegar in a test tube add a small amount of (pea-size) baking soda with the help of a spatula.
- To 1 mL of 10 % sodium hydroxide in a test tube add equal amount of 6 M hydrochloric acid. Immediately feel the test tube for heat evolved.
- To 1 mL of deionized water in a test tube add a few crystals of sucrose (table sugar) with the help of a spatula. It is important you take only a few crystals and not an entire scoop of sugar. Mix sugar and deionized water with a glass rod until sugar dissolves.
- To 1 mL of solution A in a test tube add 1 mL of solution B*. Swirl the contents of the test tube **for at least a minute** or until a change is observed.
- To 1 mL of 0.1 M sodium chloride in a test tube add drops of 0.1 M silver nitrate until a change is noticed.

Observations

Test	Observations	Physical or Chemical Change
After mixing the two chemicals		
1		
2		
3		
4		
5		

*Solution A and solution B are starch and iodine solutions respectively. Iodine reacts with starch to form a deep blue-black complex. Amylose is a component of starch and has a helical structure. The inside of the helix is just the right size and polarity to accept an iodine (I_2) molecule. When iodine is lodged within this helix, a deep blue starch-iodine complex result.

Dispose the contents of test tubes 1 thru 5 in the chemical waste container. Wash the test tubes with tap water in your work area and then rinse with distilled water from your wash bottle. Store the clean test tubes in your locker.

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1.4.6: Families of Elements

F. Families of Elements

1. Place three small test tubes in a test tube rack and add 1 mL of chloride, bromide and iodide into separate test tubes. Add drops of silver nitrate solution to each test tube till precipitate formation is observed. Mix the contents of the test tube by swirling the test tube or with the help of a glass rod. Record the color and precipitate formed in the data table.

Solution tested	Group # of the nonmetal	Observations after adding the reagent silver nitrate
Chloride solution		
Bromide solution		
Iodide solution		

2. Place six small test tubes in a test tube rack Add 1 mL of barium, calcium, strontium, lithium, potassium and sodium solutions into separate test tubes. Add drops of ammonium carbonate solution to each test tube. If precipitate forms, record color and precipitate formed in the data table. If there is no precipitate formed, write no reaction.

Solution tested	Group # of the metal	Observations after adding the reagent ammonium carbonate
Barium solution		
Calcium solution		
Strontium solution		
Lithium solution		
Potassium solution		
Sodium solution		

Dispose contents of the nine test tubes in the chemical waste container Wash the test tubes with tap water in your work area and rinse with distilled water from your wash bottle. Store the clean test tubes in your locker.

1.5: Post Laboratory Assignment

You can download this by selecting this link [Eimexpernt 1](#)

Post Laboratory Assignment

Write your name and date on page 1 and turn in the entire experiment for grading to your instructor.

1. In part A of the experiment what is your measured mass of the rubber stopper? How many significant figures are there in your measurement. Circle the digit that is the uncertain digit.
2. Calculate the density of the rubber stopper used in the experiment to the correct # of significant figures? Show your calculations.
3. Can you determine the volume of a cube of sugar using the water-displacement method? Explain.
4. Identify the tests that are chemical changes in part E of the experiment? What observations suggest chemical changes?
5. Evaluate your observations in parts F1 and F2. [Hint: Do you observe a similarity in chemical property in a group/family of elements?]

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2: NOMENCLATURE OF IONIC COMPOUNDS AND BINARY MOLECULAR COMPOUNDS

2.1: Goals and Student Learning Outcomes

2.2: Discussion

2.3: Pre-laboratory Assignment

2.4: Exercise 1- Binary Ionic Compounds

2.5: Exercise 2- Binary and Ternary Ionic Compounds

2.6: Exercise 3- Ionic and Molecular Compounds

2.7: Post-Laboratory Assignment

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2.1: Goals and Student Learning Outcomes

Upon completion of this exercise, the student will be able to:

- name and write the formulas of simple ionic compounds and binary molecular compounds.

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2.2: Discussion

Types of Compounds

Inorganic chemical compounds generally fall into one of two broad categories: ionic compounds and covalent compounds. Ionic compounds are always solid at room temperature. They have high melting points, over 350°C and often as high as 1000°C . Ionic compounds always conduct an electric current when dissolved in water or when heated to the liquid state.

Covalent compounds are also called molecular compounds. Although some are solids at room temperature, many are liquids or gases. Their melting points are low, less than 350°C . Few molecular compounds are conductors of electricity.

Metal/Nonmetal Technique

It is possible to predict whether a compound will be ionic or covalent by first classifying the elements as metals or nonmetals. A bond between a metal and a nonmetal will generally be an ionic bond and the compound that results will be named an ionic compound. A bond between two nonmetals will be a covalent bond and the compounds that result will be named molecular compounds.

Electron Dot Formulae

1A	2A	3A	4A	5A	6A	7A	
H •							
Li •	• Be •	• B •	• C •	• N •	• O •	• F •	•
Na •	• Mg •	• Al •	• Si •	• P •	• S •	• Cl •	•
K •	• Ca •						

The electron dot formulae of some of the main group elements are shown. The number of dots in the electron dot formula of the element is equal to the number of valence electrons. The group number of the main group elements is the number of the valence electrons.

Octet Rule

The octet rule states that atoms tend to gain, lose or share electrons in order to end up with the same electron configuration of one of the noble gases. All of the noble gases have eight valence electrons—except for helium, which has two valence electrons. Hydrogen is the exception to the rule of eight; this nonmetal is usually placed above the alkali metals on the periodic table because it too has one valence electron that it can lose to form a cation. Unlike the alkali metals, hydrogen atoms tend to gain or share electrons to end up with two valence electrons (**Duet Rule**).

Metal atoms can satisfy the **Octet Rule** by losing one or more electrons to become positively charged ions or cations. It is easy to predict the charge of the ions of the alkali metals, the alkaline earth metals, aluminum and a few other metals that always form the same ion. The group number of the representative elements is the number of the valence electrons. These metals lose all their valence electrons. The group number of representative metals is the positive charge of the ion. Exceptions are Pb, Sn and Bi.

Group	Charge of Ion	Example
1A	1+	Na ⁺ sodium ion
2A	2+	Ca ²⁺ calcium ion
3A	3+	Al ³⁺ aluminum ion

Nonmetal atoms can satisfy the Octet Rule in two ways:

- (1) By sharing electrons with another nonmetal to form a covalent bond, or
- (2) By gaining one or more electrons to form negatively charged ions or anions.

It is easy to predict the charge of nonmetal anions. The group number is the number of valence electrons, subtract the group number from 8 to find the negative charge of the nonmetal monoatomic ion. Names of nonmetal monoatomic ions always end with an -ide suffix.

Group	Charge of Ion	Example
5A	3-	N ³⁻ nitride ion
6A	2-	O ²⁻ oxide ion
7A	1-	Cl ⁻ chloride ion

Formulae of Ionic Compounds

Binary ionic compounds contain only two elements and their formulas are easily predicted using the periodic table. Consider the compound calcium chloride, which is used to melt ice on sidewalks. Calcium is a metal in group 2A, therefore it has two valence electrons and so calcium ion has a 2+ charge, Ca²⁺. The nonmetal chlorine is a group 7A element and has seven valence electrons. Chlorine atom needs to gain one electron to satisfy the Octet Rule and so chloride ion has a charge of 1-, Cl⁻. To balance out the charges you need one Ca²⁺ ion and two Cl⁻ ions. So, the formula for calcium chloride is CaCl₂. Subscripts of 1 are not written; when an atom does not have a subscript, it means there is one atom of calcium.

Once you have used the periodic table to predict the charge of the metal cation and the nonmetal anion, list the positive ion first and the negative ion second. For more complex examples simply criss-cross the charges to determine the subscripts of the compound formula. Let us look at the example of aluminum sulfide. The charge of aluminum (3) becomes the subscript of sulfur (2). So, the formula of aluminum sulfide is Al₂S₃.



This “criss-cross” method works in all cases **except where the subscripts are equal**. Formulas of ionic compounds are always written as whole-number ratios. For example, the correct formula of calcium oxide is CaO . The names of binary ionic compounds list metal name first, followed by the nonmetal name with an -ide suffix.

2.3: Pre-laboratory Assignment

Name: _____

Section: _____

Pre-laboratory Assignment # 2

1. Use the periodic table to complete the following table.

Atom/ Ion	# of Protons	# of Electrons	Magnitude and type of Charge	Name of atom/ion?
K				
K ⁺				
Mg				
Mg				
Cl				
Cl ⁻				
N				
N ³⁻				
Pb				
Pb ²⁺				
Pb ⁴⁺				

2. For each of the following main group elements complete the following.

If the element does not form an ion write not applicable (n/a).

Atom	Metal/ Nonmetal	Total # of Electrons	# of Valence Electrons	Electron Dot Structure	Formula and Charge of Ion Formed?
Na					
Ca					
Al					
N					
S					
F					
Ar					

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2.4: Exercise 1- Binary Ionic Compounds

Formulae of Binary Ionic Compounds.

For the following ionic compounds below, write the formula and charge of each ion, then use the charges of the ions to predict the formula of the ionic compound.

Name of the Compound	Formula of Cation (with charge)	Formula of Anion (with charge)	Formula of Compound
Sodium bromide			
Calcium iodide			
Barium oxide			
Aluminum chloride			
Lithium nitride			
Magnesium sulfide			
Potassium nitride			
Strontium fluoride			
Aluminum oxide			
Potassium sulfide			
Sodium iodide			
Lead(II) iodide			
Iron(III) chloride			
Copper(I) sulfide			
Tin(IV) oxide			
Copper(II) oxide			
Tin(II) sulfide			
Cobalt(III) iodide			

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2.5: Exercise 2- Binary and Ternary Ionic Compounds

Formulas of binary & ternary ionic compounds:

Classify each compound as ionic or covalent by using the metal/nonmetal technique. Write the formula and charge of each ion (some ions are monoatomic and some are polyatomic), then use the charges of the ions to predict the formula of the ionic compound.

Name of the compound	Ionic compound or molecular compound?	Formula and charge of cation	Formula and charge of anion	Formula of the compound
Sodium sulfide				
Sodium sulfate				
Calcium hydroxide				
Potassium nitride				
Potassium nitrite				
Potassium nitrate				
Barium chloride				
Sodium phosphide				
Sodium phosphate				
Sodium hydrogenphosphate				
Calcium carbonate				
Calcium bicarbonate				
Sodium hydrogencarbonate				
Potassium oxide				
Ammonium bromide				
Ammonium phosphate				

Note: The ammonium ion is made of nonmetals (exception). Compounds of ammonium ions (NH_4^+) are ionic compounds.

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2.6: Exercise 3- Ionic and Molecular Compounds

Nomenclature of ionic and molecular compounds

Classify each compound as ionic or covalent by using the metal/nonmetal technique. If it is a molecular compound, the compound has covalent bonds. Remember, covalent compounds do not contain ions! Covalent compounds use prefixes as part of their name, ionic compounds do not use prefixes. Write the name of each compound.

Formula	Ionic compound or covalent compound	Name
NaI		
Mg(NO ₃) ₂		
CO ₂		
K ₂ S		
K ₂ SO ₄		
SO ₂		
Na ₂ CO ₃		
NaHCO ₃		
P ₂ O ₅		
Mg ₃ (PO ₄) ₂		
BaO		
CO		
LiBr		
CS ₂		
CaS		
SO ₃		
NO ₂		
Ba(OH) ₂		

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2.7: Post-Laboratory Assignment

Post laboratory Assignment on Nomenclature

1. Complete the table below by combining the ions into a **correct formula of the compound** and **name the compound**.

	Cl^-	SO_4^{2-}	O^{2-}
K^+			
Ca^{2+}			
Al^{3+}			

2. Complete the table below. Combine the ions into a **correct formula**, and **name the compound**.

	Nitrate ion	Carbonate ion	Phosphate ion
Manganese(IV) ion			
Cupric ion			
Plumbous ion			

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3: INTRODUCTION TO ORGANIC COMPOUNDS

3.1: Student Learning Outcomes

3.2: Discussion

3.3: Pre-Laboratory Assignment

3.4: Procedure

3.5: Report-Compounds

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3.1: Student Learning Outcomes

Upon completion of this exercise, the student will be able to:

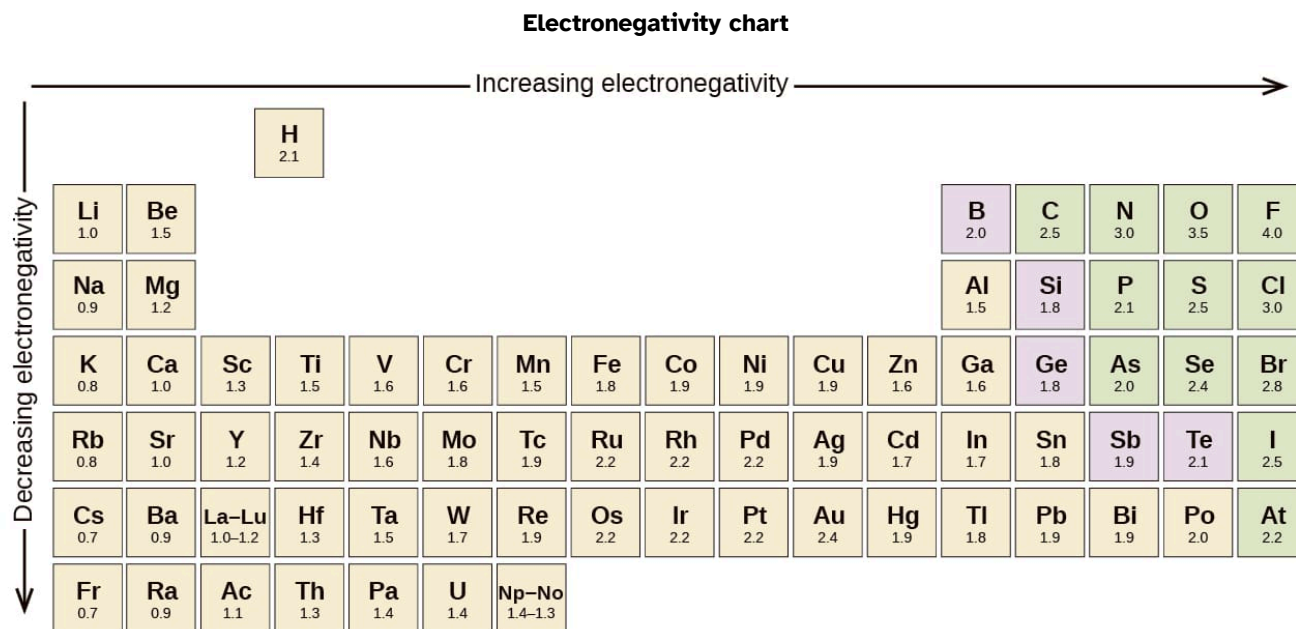
- determine the bond formed between two atoms as polar or nonpolar
- list the molecular formula from the line bond structure
- identify the overall polarity of the molecule.
- identify the strongest IMF in an organic compound
- redraw the organic compound in condensed structural formula
- classify the functional group/family of the organic compound.

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3.2: Discussion

Electronegativity

Electronegativity is the power of an atom to attract bonding pairs of electrons towards itself, when it is a part of a compound. Refer to the electronegativity chart to see the trends among the elements.



Bond Polarity

Electrons shared by two atoms are influenced by the positively charged nuclei of both atoms. For like atoms, these forces are equal. For example, in diatomic molecules such as H₂ or Cl₂, the bonded atoms have exactly the same electronegativity and hence the bond formed is a nonpolar covalent bond. Unlike atoms exert unequal attractions for their shared electrons and form polar covalent bonds

When the differences in electronegativity are less than 0.5, the bond is considered nonpolar.

Differences in electronegativity that ranges from 0.5 to 1.9 result in polar covalent bond.

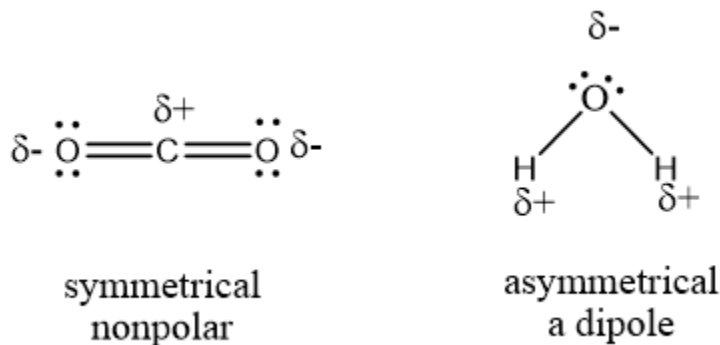
When it is greater than 1.9 the bond is ionic.

Electronegativity is used to determine the direction of the bond polarity, which can be indicated in the line bond formula by labeling the more electronegative atom with a δ^- and the less electronegative atom with a δ^+ . δ implies partial charge.



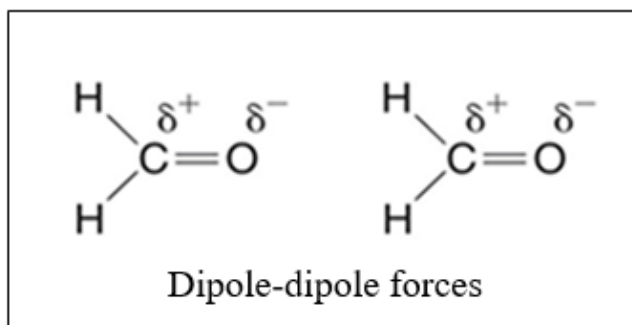
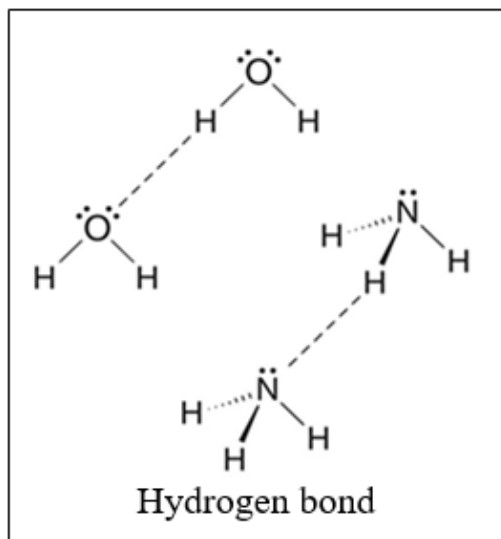
Overall Polarity

When there are two polar bonds within a molecule as in CO_2 the polar effect of these bonds around a central atom can be cancelled if they are arranged symmetrically. On the other hand, if the arrangement of the polar bonds is asymmetrical, as in the bent water molecule, H_2O , the resulting molecule has a definite positive end and oppositely charged negative end, and the molecule is called a dipole. In water the H atoms have a partial positive charge δ^+ , and the O atom has a partial negative charge, δ^- . The water molecule is overall polar.



Intermolecular Forces

Molecules interact with one another through intermolecular forces (IMF's) also known as noncovalent interactions. Polar molecules or polar parts of a molecule interact with one another through hydrogen bond or dipole-dipole forces. A **hydrogen bond** in an organic compound or a biological molecule is due to the interaction of a nitrogen or oxygen atom with a hydrogen atom that is covalently bonded to a different nitrogen or oxygen atom. **Dipole-dipole forces** takes place between polar molecules or polar parts of a molecule by orienting δ^+ end to δ^- end. Hydrogen bond is a special type of dipole-dipole interaction.



Nonpolar molecules or nonpolar parts of a molecule interact with one another through **London forces**. This force results from continuous movement of electrons in a molecule.

3.3: Pre-Laboratory Assignment

Pre-laboratory Assignment # 3 (This assignment is due at the start of lab session)

1. Find the electronegativity (EN) of each atom in the following compounds. Find the difference in electronegativity (DEN) between the atoms in each pair. Predict whether the bond between the two atoms is ionic, polar covalent or nonpolar covalent.

Compound	EN ₁ (atom 1)	EN ₂ (atom 2)	$\Delta EN = (EN_2 - EN_1)$	Bond Type
K-F				
H-Br				
Cl-F				
Cl-Cl				

2. What is the difference in electronegativity in a nonpolar covalent bond? -----

3. What is the difference in electronegativity in an ionic bond? -----

4. In general, what type of bond is formed when a metal bonds with a nonmetal? -----

5. In general what type of bond is formed when a nonmetal bonds with a nonmetal? -----

6. Refer to the periodic table in order to predict the number of valence electrons for each of the following elements:

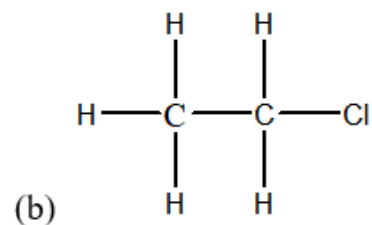
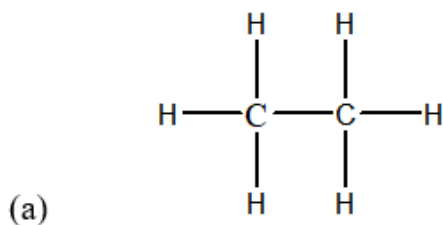
(a) oxygen -----

(b) nitrogen -----

(c) carbon -----

(d) chlorine -----

7. Draw condensed structural formula for each of the molecules whose line bond structures are shown.



8. Define the following terms. Please answer this question in full sentences.

(a) covalent bond

(b) polar covalent bond

(c) octet rule

(d) single bond

(e) double bond

- (f) triple bond
 - (g) molecular formula
 - (h) structural formula
 - (i) lone pairs
9. What is the trend in electronegativity as atomic number increases?
- (a) In a period (row)?
 - (b) In a group (family)?

3.4: Procedure

A. Bond Polarity

Bond Polarity can be determined by looking up the electronegativity values for both atoms in the electronegativity table. In the column for bond polarity, label the bonds with δ^- on the more electronegative atom and δ^+ on the less electronegative atom. Label only polar bonds with δ^+ / δ^- .

B. Molecular Formula

This formula lists the number of each type of atoms that is present in a molecule. For example, the molecular formula of $\text{CH}_3\text{CH}_2\text{CH}_3$ is C_3H_8 .

C. Polarity

Look at the molecule and evaluate the polar bonds. If the molecule contains polar bonds then the result is a polar molecule. If the molecule does not have polar bonds, the molecule is nonpolar.

Note: If the effect of the polar bonds within a molecule cancel each other around the central atom such as in CO_2 , the result is a nonpolar molecule. If they do not cancel one another and result in a polar molecule (dipole).

D. Condensed Formula

Redraw the molecule in its condensed formula in column 4. This formula shows the attachment of atoms without showing all the bonds. A carbon atom with three covalent bonds to hydrogen is drawn as CH_3 and a carbon atom with two covalent bonds to hydrogen is drawn as CH_2 .

E. Functional Group (Family)

A functional group is a group of atoms that gives the molecule a particular set of properties. The family that an organic compound belongs to is determined by the functional group. Identify the functional group or family present in the molecule.

Your choices are alkene, alkyne, aromatic, ester, amide, carboxylic acid, alcohol, phenol, thiol, sulfide, disulfide, aldehyde or ketone. Note: Alkanes are hydrocarbons with single bonds and does not contain a functional group.

F. Intermolecular Forces

Identify the **strongest** IMF possible to other molecules like itself.

Your choices are: H-bonding, Dipole-Dipole interactions, and London forces.

G. Interaction with water

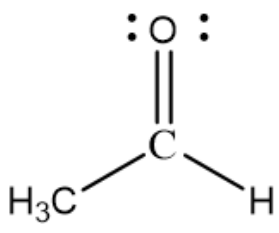
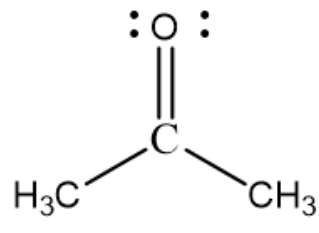
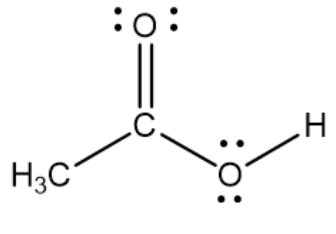
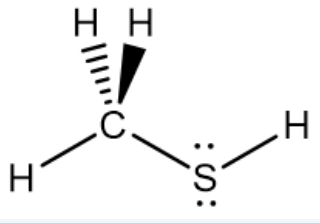
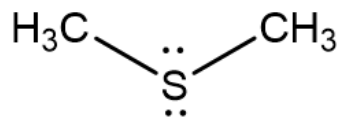
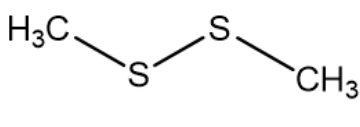
Water (HOH) is a polar molecule. Any molecule that is polar or has polar parts due to the presence of covalent bonds to the elements oxygen or nitrogen will be able to interact with water through H-bond.

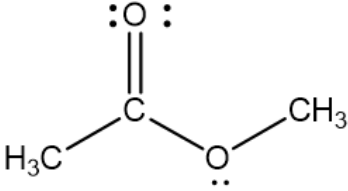
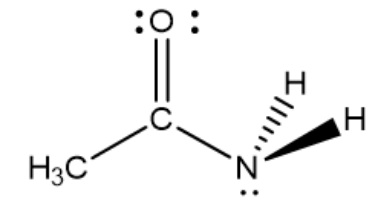
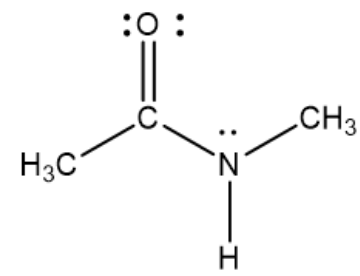
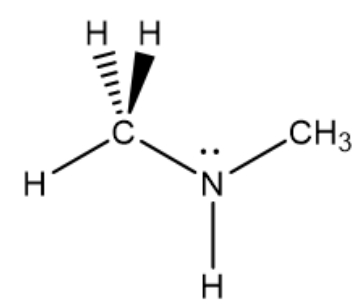
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3.5: Report-Compounds

Read carefully the procedure and complete the following table below:

Bond polarity Label δ^+/δ^- for polar bonds, Molecular Formula	Overall Polar or Nonpolar	Condensed formula & Functional group/Family	Strongest IMF present	H-bond to HOH

Bond polarity Label $\delta+/\delta-$ for polar bonds, Molecular Formula	Overall Polar or Nonpolar	Condensed formula & Functional group/Family	Strongest IMF present	H-bond to HOH
$\text{H}-\text{C}\equiv\text{C}-\text{H}$				
				
				
				
				
				
				

Bond polarity Label δ^+/δ^- for polar bonds, Molecular Formula	Overall Polar or Nonpolar	Condensed formula & Functional group/Family	Strongest IMF present	H-bond to HOH
				
				
				
				

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4: TRANSPORT THROUGH MEMBRANES (PROJECT)

4.1: Description

4.2: Directions

4.2.1: Gain Knowledge

4.2.2: Materials

4.2.3: Procedure-Removing the Shell of the Egg

4.2.4: Procedure-Treatment of Naked Eggs with Solutions

4.3: Evidence of Experimentation

4.3.1: Part A

4.3.2: Part B

4.4: References

4.1: Description

An egg has a semipermeable membrane that responds to diffusion of ions, small polar molecules and osmosis of water. In this experiment we are going to use the egg to represent a cell. The membrane is exposed, by chemically removing the shell with vinegar. The shell provides support and protection for the egg. The membrane under the shell is semipermeable and controls what enters and exits the egg (cell). The membrane allows certain nutrients and toxins to enter or leave the cell via diffusion. There is also diffusion of water molecules through the membrane called osmosis. Once the shell is removed, we will test the membrane by placing the shell-less eggs (naked eggs) in hypotonic and hypertonic solution.

This experiment will meet objectives from unit 1 and 2. In order to understand the theory of this experiment you will listen to several web lectures from Khan Academy, read the procedure provided, and perform the experiment at home. The report for the project is due in two weeks.

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4.2: DIRECTIONS

4.2.1: Gain Knowledge

4.2.2: Materials

4.2.3: Procedure-Removing the Shell of the Egg

4.2.4: Procedure-Treatment of Naked Eggs with Solutions

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4.2.1: Gain Knowledge

Read section [6.6](#) & [6.7](#) of your textbook to gain knowledge of the transport of molecules and ions with or without the semipermeable membrane.

Watch the following lectures from Khan Academy to refine your understanding.

[Diffusion and Osmosis](#)

[Hyper, Hypo, and Isotonic](#)

[Diffusion](#)

[Osmosis](#)

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4.2.2: Materials

Step 2 Materials Required

700 mL of vinegar, 3 eggs, 350 mL tap water, 350 mL corn syrup, 2 spoons, 3 Clear disposable plastic cups.



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4.2.3: Procedure-Removing the Shell of the Egg

Step 3 Procedure: Removing the Shell of the Egg

1. Label 3 cups A, B, and C.
2. Pour 350 mL of vinegar into the cups labeled A and B.
3. Submerge 2 eggs in cups A and B. What is your immediate observation, as the eggs are submerged in vinegar? The third egg is placed in cup C without vinegar, this is your control egg. Please take a picture at this stage and embed this individual work picture in the discussion area.
4. Leave the eggs overnight in vinegar to allow the egg shells to dissolve.

Note: A spoon may be used to keep the egg submerged in the vinegar. The vinegar should dissolve all or most of the shell. If the shells are not completely dissolved, you can gently rub the existing egg shell in a circular motion to remove the remaining egg shell. Be careful not to pop the egg. you may leave the egg in the vinegar for a longer period of time. or replace the existing vinegar with fresh vinegar after 6 hours. We will refer to the egg without the shell as the “naked egg”.

5. Remove the vinegar from cups A and B. Rinse the naked eggs with tap water and complete step 4.

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4.2.4: Procedure-Treatment of Naked Eggs with Solutions

Step 4 Procedure: Treatment of Naked Eggs with Hypotonic and Hypertonic Solutions.

1. Pour 350 ml of tap water into cup A and then place one naked egg in the water. Wait overnight for observations.
2. Pour 350 ml of corn syrup into cup B and then place the second naked egg in the syrup. Wait overnight for observations.
3. For observations, note the changes in size of the naked eggs in water or syrup. Compare with the control egg C.
4. Remove the eggs from the solution to a plate/towel and take picture(s) of the eggs immediately and embed this individual work picture in the discussion area.

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4.3: EVIDENCE OF EXPERIMENTATION

This assignment and experiment cover learning objectives from several chapters. The students will reference the objectives from each unit in their report. The experiment may involve student-to-student interactions via discussion board if the student decides to collaborate with classmates.

4.3.1: Part A

4.3.2: Part B

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4.3.1: Part A

Step 4 Evidence of Experimentation-Part A

The following three labeled pictures will need to be uploaded to this discussion area of canvas. Do not post the answers to the questions here.

1. Labeled pictures of eggs in vinegar after signs of chemical reaction in cups A and B. Control egg in cup C should be next to cups A and B.
2. Naked egg after treatment with tap water. Take the egg out of the solution, place on a plate/towel and take picture immediately.
3. Naked eggs after treatment with corn syrup. Take the egg out of the solution, place on a plate/towel and take picture immediately.

*Embed the jpg file of the labeled pictures of the eggs at various stages by choosing Insert Image, Upload Image. Resize the Image and Submit. Do not attach pictures. Attached pictures need to be downloaded to view. Your instructor will not download your images to view. You must complete this step to make evident your experimentation and upload pictures of the results of your individual work. **Do not post the answers to the questions on canvas.***

4.3.2: Part B

Part B-Project Report

Chem 309

Name: _____

Experiment 4 (Project # 1)

Partner: _____Canvas Group # _____

Transport Through Membranes

Submit all pages of Experiment 4 (Project 1) after completing pages 3, 4, and 5. Please write your answers neatly and in full sentences since this is a project report. Do not attach pictures to the report.

Please answer these questions in full sentences based on your procedure and observations of the project.

1. What is the chemical composition of the eggshell? Give name and formula of the major compound in the shell.
2. What is the composition of vinegar? Give name and line bond formula of the organic compound in vinegar. Identify the functional group present.
3. What is your immediate observation, when the 2 eggs are place in vinegar? Please take a picture of the eggs at this stage and embed the picture in the Project 1 area.
4. What unit learning objective does question 3 cover from unit 1? Please write the learning objective from the resources area.
5. How long did it take to remove the shell completely? Did you have to intervene so that the shell is removed uniformly?
6. Write a balanced chemical equation for the reaction of the egg shell and the household chemical. Classify the type of reaction involved.
7. What unit learning objective does question 6 covers from unit 2? Please write learning objective from the resources area.
8. Movement of solutes and solvents happen via a membrane. Where is the membrane in the egg? What does it look like and feel to touch?
9. What hypotonic solution did you use for your experiment? What happened to the size of the egg after treatment in hypotonic solution used? Please take a labelled picture of the egg treated with the hypotonic solution and embed the picture in the Project 1 area.
10. What hypertonic solution did you use in the experiment? What happened to the size of the egg after treatment in hypertonic solutions used? Please take labelled pictures of the egg treated with the hypertonic solution and embed the picture in the Project 1 area.
11. How long did it take to soak the naked egg in the two solutions (hypotonic and hypertonic) to make the observations?
12. What unit-learning objective do question 9 & 10 cover from unit 2? Please write the unit learning objective from the resources area.
13. Give an example of one other household chemical (besides vinegar) that may be used to remove the eggshell?
14. Did you encounter any unexpected result during this experiment? Suggest any changes in procedure to overcome the unexpected result.
15. In your own words explain why does the naked egg change sizes in hypotonic and hypertonic solutions? What scientific term is used to describe the egg after the treatment with these solutions?

16. Give one example of an application of osmosis in our everyday lives. Explain how osmosis works in your application.

4.4: References

This project was adapted for Chem. 309 by Deboleena Roy using resources from the following sources

- a. CSU Bakersfield
- b. Biology Junction.

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5: PROPERTIES OF SOLUTIONS (ALE EXP)

5.1: Student Learning Outcomes

5.2: Introduction

5.3: Pre-Laboratory Assignment

5.4: Procedure and Observations

5.4.1: Conductivity

5.4.2: Dissolved Oxygen

5.5: Post Laboratory Assignment

5.1: Student Learning Outcomes

Properties of Solutions: Electrolytes and Solubility of Gases

This experiment is a part of the Authentic Learning Experience (ALE)

Goals and Student Learning Outcomes

In this experiment, the student will:

- study the electrolyte properties of solutions
- investigate relationship of solubility of a gas with temperature
- enter observations on conductivity of solutions and collected creek/river water samples.
- measure dissolved oxygen gas concentration in water samples from creek/river.
- develop a sense of belonging in the STEM community

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5.2: Introduction

Electrolytes and Non-electrolytes

When a solute dissolves in a solvent, the solute particles may be ions with positive or negative charges or molecules without a charge. Anions are negatively charged ions and positive ions are cations. Aqueous solutions of electrolytes conduct an electric current because the solutions contain ions. When dissolved in water, sodium chloride dissociates into sodium ions and chloride ions. The charged ions conduct an electric current. Hence sodium chloride is an electrolyte.



Solutions of molecules without a charge do not conduct an electric current. In aqueous solution these molecules are termed non-electrolytes. Glucose, the sugar found in blood, is a molecule and exists in water as uncharged molecules and does not conduct an electric current.

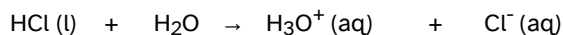


Electrolyte ions are found in all body tissues and fluids, and are routinely measured in blood serum and urine to evaluate acid-base and electrolyte balance. The most commonly measured serum electrolytes are sodium ion (Na^+), potassium ion (K^+), bicarbonate ion (HCO_3^-), and chloride ion (Cl^-). Na^+ and K^+ ions are important in transmission of nerve impulses, HCO_3^- is involved in the transport of CO_2 from tissues to the lungs, and Cl^- help regulate the body's water content. Other electrolytes found in body fluids include calcium ion (Ca^{2+}), magnesium ion (Mg^{2+}), hydrogen phosphate ion (HPO_4^{2-}), and sulfate ion (SO_4^{2-}).

Strong and Weak Electrolytes

Some electrolyte solutions are good conductors of electricity. They are called strong electrolytes. Weak electrolytes are poor conductors of electric current. Both strong and weak electrolyte solutions contain ions. Strong electrolytes are almost completely ionized in water and weak electrolytes are only slightly ionized. A conductivity apparatus shown below is used to determine the electrolyte properties of solutions.

The acid in your stomach hydrochloric acid (HCl) is a strong electrolyte. In a mixture of water and hydrochloric acid, over 99 % of the acid ionizes to hydronium ions and chloride ions, less than 1 percent of the acid remains as uncharged molecules (HCl). Hydrochloric acid is a strong electrolyte because it is almost totally ionized in water.



< 1 %

> 99%

The acid in vinegar, acetic acid (CH_3COOH) is a weak electrolyte. In a mixture of acetic acid and water only about 8 % of the acid ionizes to acetate ion (CH_3COO^-) and H_3O^+ ions. 92 % is present as molecular CH_3COOH . Because it is only slightly ionized, acetic acid solutions weakly conduct electricity.



92 %

8 %

Conductivity Apparatus

In the laboratory we will use the Flinn Conductivity Meter.

This conductivity meter is safe and portable because a 9-volt battery is used.

Precautions must be observed in using the conductivity apparatus.

1. Electrodes must be washed with de-ionized water before testing a new solution. This will prevent cross contamination of samples.
2. When testing the solution, be sure the electrodes are immersed well beneath the surface of the solution.



Solubility of a Gas and Temperature:

Solubility is defined as the amount of solute that dissolves in a solvent as a given temperature. Gases are more soluble in solvent water at lower temperature than they are at higher temperatures. Carbonated beverages contain carbon dioxide gas in water or an aqueous solution of sugar syrup. As the cold beverage warms to room temperature, the CO_2 gas comes out of solution and the drink loses its fizz. On a hot day, fish may show signs of stress, because of low concentrations of oxygen in the water. Less oxygen from the air dissolves in the warm water than in cooler water.

5.3: Pre-Laboratory Assignment

Name:

Section:

Pre-Laboratory Assignment on Experiment # 5

Please answer the following questions in full sentences:

1. How does the water-solubility of a gas change when the temperature is increased?
2. What do you understand by the term electrolytes?
3. What is the difference between a strong electrolyte and a weak electrolyte?
4. What precautions should you observe to prevent cross contamination of samples during conductivity experimentations?
5. You will use the following compounds as aqueous solution during the experiment. Classify each of the following as an ionic compound or a molecule. (Hint: Use the metal/nonmetal technique)

- | | | | |
|--|-------|--------------------------------------|-------|
| a. NaCl | ----- | e. CH ₃ CO ₂ H | ----- |
| b. C ₁₂ H ₂₂ O ₁₁ | ----- | f. HCl | ----- |
| c. Mg(OH) ₂ | ----- | g. C ₃ H ₈ O | ----- |
| d. H ₂ O | ----- | h. NaHCO ₃ | ----- |

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5.4: PROCEDURE AND OBSERVATIONS

5.4.1: Conductivity

5.4.2: Dissolved Oxygen

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5.4.1: Conductivity

A. Conductivity

Obtain two well plates from your instructor. Wash them with tap water and rinse with de-ionized water. A few drops of the NaCl solution to be tested is placed in one of the wells of the spot plate. Please have a de-ionized wash bottle and a beaker to catch waste. Rinse the Cu electrodes of the conductivity tester with DI water, turn on the tester, and then place the electrodes into the solution as shown in the picture. Identify the solution as strong, weak, or nonelectrolyte based on the table below. You will test a few solutions before you test the water collected from the creek/river and compare with tap water and de-ionized water.



Red LED	Green LED	Conductivity
Very Bright	Medium	Strong
Bright	Dim	Weak
Medium	Off	Weak
Dim	Off	Very Weak
Off	Off	Non

Table of observations for conductivity:

Date/Time/Temperature of collection: -----

#	Sample	LED Observations	Strong/Weak/Non Electrolytes
1.	Table Salt, NaCl		
2.	Baking Soda, NaHCO ₃		
3.	Hydrochloric acid or Stomach acid, HCl		
4.	Milk of Magnesia, Mg(OH) ₂		
5.	Vinegar, (5% CH ₃ CO ₂ H)		
6.	Table sugar or Sucrose, C ₁₂ H ₂₂ O ₁₁		
7.	Rubbing alcohol, C ₃ H ₈ O		
8.	Tap water		

9.	Creek/River water		
10.	De-ionized water		

5.4.2: Dissolved Oxygen

B. Dissolved Oxygen (DO) Concentrations in Creek/River water.

You are provided a ProSolo meter with a probe that measures the concentration of O_2 in our water. The probe is stored with a gray rubber sleeve. At the bottom of the sleeve is a sponge that is saturated with water. The probe is calibrated by your instructor to the current barometric pressure, temperature, and 100% water saturation of the water-soaked sponge. The instrument is ready for analyzing samples.

Please turn on the meter (green button). Remove the gray sleeve and place the probe in 200 mL of tap water in a 400 mL beaker. Ensure the probe is immersed in the tap water. Record the temperature, pressure and DO in the table below. Record the units

Next, place the probe in the river/creek water sample. This water is stored in ice after collection to ensure that the gaseous solutes such as O_2 is not lost.

The third sample is warm tap water. Ensure that the probe is immersed in the sample. Record temperature, pressure and DO of the tap water sample in the table below.

Sample	Temperature ($^{\circ}C$) and Pressure (mm of Hg)	Dissolved Oxygen (mg/L)
Tap water at room T		
River/Creek water		
Warm Tap water		

5.5: Post Laboratory Assignment

“Thermal pollution” is a term used to describe hot water added to a lake or river by an electrical generating plant or other sources. Why is increased water temperature an environmental concern?

Name two ions likely to be present in our creek/river water. Give the formula of the ions.

Name two electrolyte ions that are important to our health? What role do they play in the human body or blood serum?

Compare the conductivity of tap water, creek/river water, and de-ionized water. Which sample is most conducting? Why?

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6: ACIDS AND BASES (PROJECT)

6.1: Description

6.2: Directions

6.2.1: Gain Knowledge

6.2.2: Materials

6.2.3: Procedure-Prepare Red Cabbage Juice Indicator

6.2.4: Procedure-Measuring the pH of Common Household Chemicals.

6.2.5: Procedure-Reaction of an Acid and a Base

6.2.6: Procedure-Effect of your Breath on a Neutral Solution of Water

6.3: Evidence of Experimentation

6.3.1: Part A

6.3.2: Part B

6.4: References

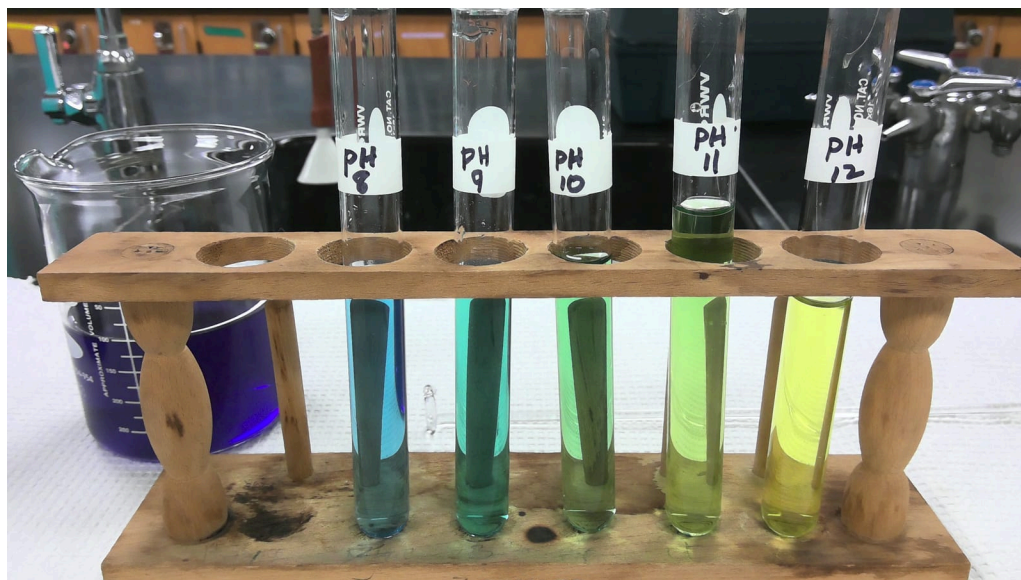
6.1: Description

Many household solutions or chemicals have acidic or basic (alkaline) properties. The pH scale helps us to determine this property. In this home-based experiment, you will use the extract of red cabbage to test liquids and chemicals around you. The red cabbage contains a pigment called flavin, which responds to the concentration of H_3O^+ ions and produces a color change. Based on this color change you will successfully identify the approximate pH of common household chemicals using the color standard posted on canvas.

This experiment will meet objectives from unit 1, 2, and 3. In order to understand the experiment you are encouraged to learn about acids and bases first. The experiment can be performed easily at home. The report for the project is due in two weeks.



Indicator color for pH 2 through 7 (first picture with pink to blue color)



Indicator color for pH 8 through 12 (2nd picture with blue to yellow color)

6.2: DIRECTIONS

6.2.1: Gain Knowledge

6.2.2: Materials

6.2.3: Procedure-Prepare Red Cabbage Juice Indicator

6.2.4: Precedure-Measuring the pH of Common Household Chemicals.

6.2.5: Procedure-Reaction of an Acid and a Base

6.2.6: Procedure-Effect of your Breath on a Neutral Solution of Water

6.2.1: Gain Knowledge

Step 1 Gain Knowledge

Chapter 7 of your textbook is also a good resource for acids and bases.

Gain knowledge on the topic of acids and bases by watching as many videos as necessary from Khan Academy or studying Acids and Bases.

[Acids and Bases Videos](#)

[Acids and Bases Textbook](#)

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6.2.2: Materials

Step 2 Materials

Red cabbage, vinegar, carbonated water, shampoo, baking soda, kitchen measuring cup, 3 other household chemicals of your choice, and 7 clear disposable plastic cups.

Clear plastic cup for the experiment.

Clear plastic cup

6.2.3: Procedure-Prepare Red Cabbage Juice Indicator

Step 3 Procedure

Prepare red cabbage juice indicator

Cut several leaves of red cabbage (0.25 lbs. to 0.50 lbs.) into small pieces and place them in a pot with enough water to cover the leaves. Warm the mixture of water and red cabbage leaves until the water turns a dark purplish color. Do not boil the water. Transfer the purple liquid into a clear glass container, pouring through a strainer to remove the cabbage leaves. Take a picture of this container with the liquid.

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6.2.4: Precedure-Measuring the pH of Common Household Chemicals.

Step 4 Procedure

Measuring the pH of common household chemicals

Some examples of common household chemicals that can be tested for pH are vinegar, carbonated soda or water, dilute solutions of shampoo and baking soda. For powders, dissolve 3 teaspoon of powder in 100 mL of water. For shampoo, dissolve 1 tablespoon of shampoo in 100 mL water. Use a kitchen measuring cup for 100 mL of water. Choose three other chemicals or solutions not in this list for this experiment. Make a solution with the powders as necessary. Add an equal volume of cabbage indicator to the solutions of household chemicals. Note the color and compare the colors to the standards pH 2 to 12. Predict the approximate pH of the household chemicals/solutions and take pictures. Do not dispose of the chemicals yet. Use clear plastic cups to perform this part of the experiment.

Note: Coffee, tea, and milk have strong coloration and may mask the color of the cabbage indicator. This in turn may lead to inconclusive results in the pH determination. For this experiment, it is best to choose solutions that are colorless to pale in color.

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6.2.5: Procedure-Reaction of an Acid and a Base

Step 5 Procedure

Reaction of an acid and a base

Use the cup containing vinegar. Note the color and pH. Slowly add the contents of the cup containing baking soda to the cup of vinegar. Note any color changes or other observations. Predict the approximate pH after the mixing of the two chemicals. Take pictures of the color change.

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6.2.6: Procedure-Effect of your Breath on a Neutral Solution of Water

Step 6 Procedure

Effect of your breath on a neutral solution of water

Use approximately 50 mL of dark purple cabbage indicator solution and dilute with an equal amount of tap water. Note the color and pH. Using a clean straw, blow bubbles into the water until there is a color change. Note the pH change by comparison with the color standards.

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6.3: EVIDENCE OF EXPERIMENTATION

6.3.1: Part A

6.3.2: Part B

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6.3.1: Part A

The following four labeled pictures will need to be uploaded to the discussion area of canvas. **Do not post the answers to the questions in canvas.**

1. Red cabbage indicator in step 3.
2. Household chemicals with cabbage indicator in step 4.
3. Reaction of acid and base in step 5. (before and after)
4. Effect of breath on a neutral solution in step 6. (before and after)

Embed the jpg file of the labeled pictures of your experiment by choosing Insert Image, Upload Image. Resize the Image and submit. Do not attach pictures. Attached pictures need to be downloaded to view. Your instructor will not download your images to view. You must complete this step to make evident your experimentation and upload pictures of the results of your individual work. **Do not upload the answers to the question in the discussion area of canvas.**

6.3.2: Part B

Submit all pages of Experiment 6 (Project 2) after completion of pages 3, 4, and 5. Please write your answers neatly and in full sentences since this is a project report. Do not attach pictures to the report.

Please answer these questions in full sentences based on the procedure, and observations of the project.

1. What is the color of the extract from red cabbage in step 3? Estimate the pH of this solutions by comparing with the color standard in canvas. Please take a picture and embed the picture in the Project 2 area.
2. Complete the following table of observations on the household chemicals. Please write the names of three other solutions in the table below. Observe the color change and predict pH.

Sample	Color Change with Indicator	Predicted pH of the Sample
Vinegar		
Carbonated Soda		
Shampoo solution		
Baking Soda solution		
Other 1		
Other 2		
Other 3		

3. Please take picture(s) and embed in the Project 2 area. The cups should be clearly labelled with household solutions used.
4. Draw a graph with pH 1-14 on the X-axis and the name of the household chemical used distributed on the y-axis; plot the values for each household chemical clearly on the graph. Please review the references prior to this graph and attach a copy of the graph to your report.

5, Record the initial and final color of solutions in step 5.

Initial color and estimated pH of vinegar -----

Final color and pH (after adding baking soda solution) -----

6. What other observations did you note during this part of the experiment?

7. Write a balanced chemical equation for step 5 of the experiment. Classify the type of reaction involved.

8. For step 5 what unit learning objectives can you reference from unit 1, 2, and 3?

9. Record the initial and final color of solutions in step 6.

Initial color and estimated pH of tap water -----

Final color and pH of tap water (after blowing through a straw) -----

10. Explain the **change in color and pH** that occurred when you blew through a straw into water

in step 6 based on the equation $\text{CO}_2 (\text{g}) + \text{H}_2\text{O} (\text{l}) \rightleftharpoons \text{H}_2\text{CO}_3 (\text{aq})$.

Use Le Chatelier's principle and add equilibrium equations necessary to explain.

11. Please reference a unit learning objective from unit 3 that was used to explain step 6 observations.
12. Please take before and after pictures of step 5 and 6 and embed in the Project 2 area. The cups should be clearly labeled with the step #.
13. In this experiment, we used red cabbage indicator to test acid and base properties. Is it possible to use other vegetables/ fruits/flowers to achieve the same results? List a couple of examples.
14. Give one example of applications in health care or daily life of one of the concepts you learned in this topic or this experiment.
15. Did you encounter any unexpected result during the experiment? Suggest any changes in procedure to overcome the unexpected result.

6.4: References

References

This project was adapted for Chem. 309 by Deboleena Roy using resources from Stem Curriculum for K-12 and Stanford Outreach.

[Red Cabbage Lab.](#)

[Red cabbage Chemistry](#)

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7: ACIDS, BASES, BUFFERS, AND PH MEASUREMENTS (ALE EXP)

7.1: Goals and Student Learning Outcomes

7.2: Discussions

7.3: Pre-laboratory Assignment # 7

7.4: Procedure

7.4.1: A. Measuring the pH of solutions using a pH meter

7.4.2: B. Effect of Buffers on pH

7.5: Post-Laboratory Assignment

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7.1: Goals and Student Learning Outcomes

This experiment is part of the Authentic Learning Experience (ALE)

In this experiment, the student will:

- measure the pH of several solutions using a pH meter
- enter individual observations of the pH of solutions and creek/river water sample.
- develop a sense of belonging in the STEM community.
- observe changes in pH as acid or base is added to the buffered and non-buffered solutions

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7.2: Discussions

Acids and bases

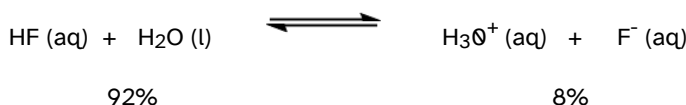
Acids have a sour taste. Vinegar is sour because of the acetic acid present in it and lemon juice is sour due to the citric acid present in it. The word acid is derived from the Latin word *acidus*, meaning “sour.” Bases taste bitter and feel slippery to the touch. Acids react with bases via a double replacement reaction known as a neutralization.

Acids and bases can be defined in several different ways. The most useful is the Brønsted-Lowry definition which states that an acid is a substance capable of donating a hydrogen ion (H^+) and a substance capable of accepting a hydrogen ion (H^+) is a base.

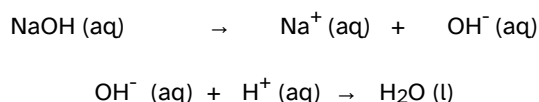
When added to water, all acids can donate a H^+ ion to water; the products of the reaction will be a hydronium ion (H_3O^+) and a negative ion. Because ions are formed, all acids are electrolytes. Strong acids are completely ionized in water and are therefore strong electrolytes. Strong acids easily donate H^+ to acceptors. Hydrochloric acid, is a strong acid. Only a small number of common acids are strong acids. Sulfuric acid is a strong acid and so is nitric acid.



Weak acids are slightly ionized in water and are called weak electrolytes. Weak acids donate H^+ , but with limited success. Hydrofluoric acid is a weak acid. Most common acids are weak acids like vinegar.



Bases, too, are either strong or weak. Strong bases are nearly totally ionized in water and hence are strong electrolytes; weak bases are weak electrolytes because they only partially ionize in water. Strong bases are better H^+ acceptors than weak bases. Sodium hydroxide (NaOH) and potassium hydroxide (KOH) are common strong bases; both ionize in water to release hydroxide ion (OH^-). The hydroxide ion is an excellent H^+ acceptor.



We encounter acids and bases every day. There are acids in oranges, lemons, and vinegar. Some household bases include ammonia, detergents and oven-cleaning products.

The pH scale

The pH scale is a quantitative measure of hydronium ion concentration [H_3O^+]. On this scale the lower the pH the more acidic the solution. When $pH = 7$ the solution is neutral. The solution is basic when the pH is greater than 7. The concentrations of hydronium ion are often very small, involving negative powers of ten. So, pH is obtained from the negative logarithm of the hydronium ion concentration. It is a unitless number and varies from ~ 0 to 14 pH.

$$pH = -\log [H_3O^+]$$

A solution with a $[H_3O^+] = 1.0 \times 10^{-4} M$ has a pH of 4.00, and is acidic.

Experiment # 12, Electrolytes, Acids, and Bases; Laboratory manual; General, Organic and Biological Chemistry an integrated approach; Macaulay, Bauer and Bloomfield.

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7.3: Pre-laboratory Assignment # 7

Name:-----

Pre-laboratory Assignment # 7

Please review the introduction and procedure prior to completing this assignment

1. List one physical property of an acid and one for a base.
2. What information does a pH of a solution give?
3. What is a buffer solution?
4. What is the composition of a buffer solution?
5. Which of the following can you use to determine if the solution is acidic, basic, or neutral?
 - a. purple cabbage dye
 - b. litmus paper
 - c. a pH meter
 - d. all of them
6. An unknown solution is measured to have a pH of 9.6. After adding 5 mL of acid to 1 liter of the solution the pH changed to 9.5. Which term best describes the solution?

Acid	Buffer	Electrolyte	Amphoteric
------	--------	-------------	------------
7. What is the pH range of blood? -----
8. Based on the procedure provided in the experiment what do you use to rinse pH probe?
9. True or False It is important that the bulb of the pH meter is immersed in solution to be tested.
10. Based on your procedure provided, how do you turn OFF the pH meter?

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7.4: PROCEDURE

The procedure is in 2 parts:

Part A: Measuring the pH of solutions using a pH meter.

Part B: Procedure and Observations for Effect of Buffers on pH.

7.4.1: A. Measuring the pH of solutions using a pH meter

7.4.2: B. Effect of Buffers on pH

7.4.1: A. Measuring the pH of solutions using a pH meter

A. Measuring the pH of solutions using a pH meter:

1. Take a clean 400 mL beaker and half fill the beaker with 200 mL of tap water. You will use this beaker with tap water to rinse the probe of the pH meter.
2. Your instructor will provide 3 culture tubes for this experiment. Obtain 2.5 ml of vinegar, hydrochloric acid, and windex cleaner in each of the 3 culture tubes provided.
3. The device we will use for pH measurements is the Flinn pH meter. The electrode bulb of the meter is in a red storage solution in a culture tube in the off position. The storage solution is made of saturated potassium chloride solution, pH 4 buffer, and a red dye.
4. Request a pH meter with the culture tube from your instructor. Have a test tube rack ready to hold the pH meter/culture tube.
5. Short press ON/OFF key to turn on the pH meter.
6. Wash the electrode in tap water in the 400 mL beaker from step 1 and place it in the culture tube with 2.5 mL vinegar solution to measure pH. It is important that the bulb of the pH meter is immersed in the solution being tested. Wait until the number is stable and the symbol  is shown on the screen to record the reading.
7. Between sample measurements, wash the electrode in the 400 mL beaker of tap water. Do not use de-ionized water to rinse the meter. Measure the pH of the samples of hydrochloric acid and windex and record. Label acidic, basic, or neutral.
8. Contents of the three culture tubes are disposed into the waste container. Wash the 3 culture tubes with tap water from the sink at your bench area.
9. Obtain 2.5 mL of each of the next three samples for pH measurements. Continue the measurements for the rest of the samples in the clean culture tubes by rinsing with tap water.
10. Solutions with pH close to 7 may be labeled as slightly acidic or slightly basic.
11. Obtain 2.5 mL of river water in **a clean culture tube from your instructor**. Note the date, time, and temperature of the water collection. Measure the pH of the sample.
12. After use, wash the electrode with tap water, turn off the pH meter (long press for 3 seconds) and return to the culture tube of in 6 ml red storage solution.

Observation Table

Substance	pH measurement	Acidic, Basic or Neutral?
Vinegar		
Hydrochloric acid (stomach acid)		
Windex Cleaner		
Lye or Drano		
Tea		
Coffee		
pH 7 buffer		
River Water		

Date/Time/Temperature of water collection: _____

- What are some of the sources of hydronium ions in our creeks and rivers?
- Calculate the pH of the following solutions. Is the solution acidic, basic, or neutral?
 - $[\text{H}_3\text{O}^+] = 2.5 \times 10^{-10} \text{ M}$
 - $[\text{H}_3\text{O}^+] = 4.8 \times 10^{-3} \text{ M}$

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7.4.2: B. Effect of Buffers on pH

B. Procedure and Observations for Effect of Buffers on pH

Effect of Adding an Acid

1. Please review the procedure on use of the pH meter on part A and apply to this part of the experiment.
2. Place 2.5 mL of each of the following solutions into 4 culture tubes.
 - a. A buffer with pH 9
 - b. A buffer with pH 3
 - c. 0.1 M NaCl
 - d. Tap H₂O
3. Determine the pH of each sample using the pH meter. Record.
4. Add 2 drops of 0.1 M HCl (acid) to test tube a. **Stir with a glass rod to mix.** Measure the pH with a meter and record.
5. Add 2 more drops of 0.1 M HCl. **Stir with a glass rod to mix.** Measure and record the pH. Repeat this procedure with each of the other solutions.
6. Determine the change in pH, after adding drops of acid. The pH change should be calculated using the difference in initial pH of the sample and pH after drops of acid are added. Identify the solutions that are buffers.

Solution	Initial pH	pH after adding 2 drops of HCl	pH after adding 4 drops of HCl	pH change on adding drops of HCl	Is this sample a buffer?
pH 9 buffer					
pH 3 buffer					
0.1 M NaCl					
Tap H ₂ O					

Disposal

Contents of the test tube in this section are disposed in the chemical waste container. Wash the culture tubes with tap water from the sink at your work bench before you proceed to the next step.

Effect of Adding a Base:

- Please review the procedure on use of the pH meter on page 7 and apply to this part of the experiment.
- Place 2.5 mL of each of the following solutions into 4 culture tubes.
 - A buffer with pH 10
 - A buffer with pH 4
 - 0.1 M NaCl
 - Tap H₂O
- Determine the pH of each sample using a pH meter. Record.
- Add 2 drops of 0.1 M NaOH (base) to test tube a. **Stir with a glass rod to mix.** Measure the pH with a meter and record.
- Add 2 more drops of 0.1 M NaOH. **Stir with a glass rod to mix.** Measure and record the pH. Repeat this procedure with each of the other solutions.
- Determine the change in pH after adding drops of base. The pH change should be calculated using the difference in initial pH of the sample and pH after drops of base are added. Identify the solutions that are buffers.

Solution	Initial pH	pH after adding 2 drops of NaOH	pH after adding 4 drops of NaOH	pH change on adding drops of NaOH	Is this sample a buffer?
pH 10 buffer					
pH 4 buffer					
0.1 M NaCl					
Tap H ₂ O					

Disposal

Contents of the culture tube in this section are disposed in the chemical waste container. Wash the culture tubes with tap water from the sink at your work bench and rinse with de-ionized water from the wash bottle provided in your locker. Return the 4 culture tubes to your instructor.

7.5: Post-Laboratory Assignment

Post laboratory Assignment

1. Which samples in part B of your experiment showed the greatest change in pH when drops of acid or base were added? Explain why there are changes in pH.
2. Which samples in part B of your experiment showed little or no change in pH when drops of acid or base were added? Why?
3. Is a buffer supposed to keep the pH of a solution at 7? Explain.
4. The label of a shampoo claims that it is pH balanced. What do you think “pH balanced” means?
5. Explain how the $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ buffer system helps maintain blood serum at constant pH with the help of chemical equations. At what pH is it most effective?

8: LE CHATELIER'S PRINCIPLE AND BLOOD BUFFERS

8.1: Goals and Student Learning Outcomes

8.2: Discussion

8.3: Pre-laboratory Assignment # 8

8.4: Report

8.4.1: Respiration Helps Maintain Blood pH During Acidosis

8.4.2: Respiration Helps Maintain Blood pH During Alkalosis

8.4.3: Kidneys Help Maintain Blood Serum pH

8.4.4: Dihydrogen phosphate and Hydrogen phosphate Buffer System

8.5: Post laboratory Assignment # 8

8.1: Goals and Student Learning Outcomes

In this exercise, the student will:

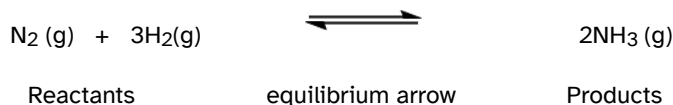
- use Le Châtelier's principle to explain how a reaction in equilibrium responds to being disturbed.
- apply Le Châtelier's principle to maintain the pH of blood
- describe the role of buffers, respiration, and the kidneys in maintaining a stable blood serum pH.

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8.2: Discussion

Chemical equilibrium

Some chemical reactions will, after forming product reverse and reform reactants. These are reversible reactions. For example, the generation of ammonia is a reversible reaction.



Once ammonia is formed, the reaction will reverse, re-forming nitrogen and hydrogen. Eventually, the rate of the formation of ammonia (the forward reaction) and the rate of formation of nitrogen and hydrogen gases (the reverse reaction) becomes equal. The balance of the rates of reaction is called chemical equilibrium. A special type of arrow, called an equilibrium arrow (shown in the equation), is used in this chemical equation to indicate that both the forward and reverse reactions can take place simultaneously. This does not mean that the concentrations of ammonia and of nitrogen and hydrogen are equal. It means there is no net change in the concentrations. In other words, the concentrations of products and reactants stay the same.

The Equilibrium Constant K

If we measure the concentrations of ammonia, nitrogen and hydrogen present, the ratio of products to reactants would be constant. This value is called the equilibrium constant, K , and it is a characteristic of equilibrium reactions at a given temperature. K is defined as

$$K = \frac{[\text{Products}]}{[\text{Reactants}]}$$

The brackets, [], mean “molar concentration of.” So, the equation states that the equilibrium constant, K , is equal to the molar concentration of the products divided by that of the reactants. If there is more than one reactant or product, the concentrations are multiplied together. For the generation of ammonia, the expression for K is shown below. The superscripts in the expression come from the coefficients in the balanced chemical equation.

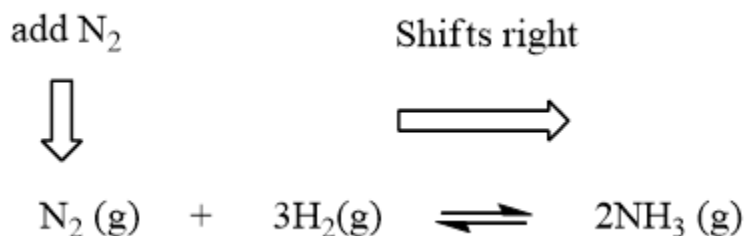
$$K = \frac{[\text{NH}_3]^2}{[\text{N}_2] [\text{H}_2]^3}$$

Only substances with concentrations that change appear in an equilibrium expression. Solids (s) and pure liquids (l) do not appear in the equilibrium expression.

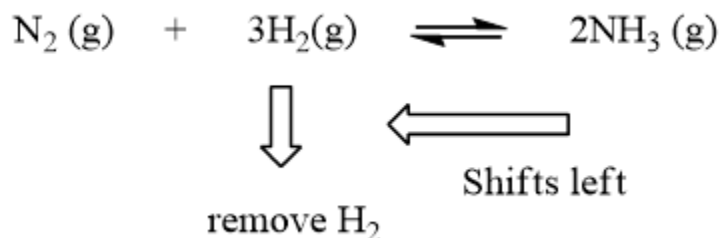
Effect of Concentration on Equilibrium – Le Chatelier’s Principle

Suppose that we had our reaction for the generation of ammonia sitting in a reaction vessel and that it was at equilibrium. What would happen if we decided to inject more nitrogen into the vessel?

According to the Le Chatelier's principle, if we disturb an equilibrium-chemists refer to this as applying stress (upset) to the equilibrium-the rate of the forward or reverse reaction will change to offset the stress and regain equilibrium. Applying this principle, if we add more N_2 to our system (it appears on the reactant side of our equation), the rate of the forward reaction will increase, shifting the equilibrium to produce more products. This rate increase occurs because adding more N_2 molecules increases the chance that N_2 will collide with H_2 molecules, forming ammonia more rapidly.



As a contrast, suppose that we must remove one of the reactants, H_2 . To regain equilibrium, the reverse reaction must be faster than the forward reaction, allowing the H_2 to be replenished. The equilibrium shifts to the left, forming more reactants and reestablishing the equilibrium between the two rates.



In general, we can think of equilibrium as a balancing act between forward and reverse reaction. If one side of the reaction gains a substance, the reaction shifts to the other side to regain its equilibrium. If one side of the reaction loses a substance, the reaction will shift towards that side to regain its equilibrium.

8.3: Pre-laboratory Assignment # 8

Name: _____

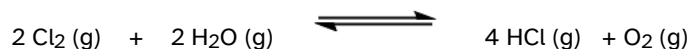
Section: _____

Pre-laboratory Assignment # 8

1. Given the following equilibrium $A + B \rightleftharpoons C + D$, what will be the effect on the equilibrium if each of the following takes place?

Applying Stress: Upset	Will the rate of forward or reverse reaction increase?	Will the concentration of these substances increase or decrease?
a. Increase in the concentration of C		A and B
b. Decrease in the concentration of A		C and D
c. Increase in the concentration of B		C and D
d. Decrease in the concentration of D		A and B

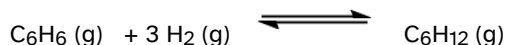
2. Consider the following chemical system at equilibrium.



For each of the following adjustments of conditions, indicate the effect (shifts left, shifts right, or no effect) on the position of the equilibrium.

- a. Increasing the Cl_2 concentration _____
- b. Increasing the O_2 concentration _____
- c. Decreasing the water vapor concentration _____
- d. Decreasing the HCl concentration _____

3. Consider the following chemical system at equilibrium.



For each of the following adjustments of conditions, indicate the effect (shifts left, shifts right, or no effect) on the position of the equilibrium.

- a. Increasing C_6H_{12} concentration _____
- b. Decreasing the C_6H_6 concentration _____
- c. Adding a catalyst to the equilibrium mixture _____

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8.4: REPORT

This exercise on blood buffers consists of 5 parts. Please answer the questions on the following pages to gain an understanding on how the organs play a role in maintaining the pH of blood. The exercise utilizes Le Chatelier's Principle. So it is important to understand applying stress (upset) to an equilibrium followed by the response to an equilibrium. The first four topics utilizes the Carbonic Acid and Bicarbonate ion buffer system. Both lungs and the kidneys have roles in compensation. The last topic utilizes the Dihydrogen phosphate and Hydrogen phosphate buffer system. Only the kidneys have a role with this buffer system.

1. Respiration helps maintain blood serum pH during acidosis.
2. Respiration helps maintain blood serum pH during alkalosis.
3. Kidneys helps maintain blood serum pH during acidosis.
4. Kidneys helps maintain blood serum pH during alkalosis.
5. Dihydrogen phosphate and the hydrogen phosphate buffer system.

8.4.1: Respiration Helps Maintain Blood pH During Acidosis

8.4.2: Respiration Helps Maintain Blood pH During Alkalosis

8.4.3: Kidneys Help Maintain Blood Serum pH

8.4.4: Dihydrogen phosphate and Hydrogen phosphate Buffer System

8.4.1: Respiration Helps Maintain Blood pH During Acidosis

1. Explain how respiration (lungs) helps maintain blood serum pH during acidosis.

Write the two equilibrium equations responsible for maintaining pH

Equation 1: Show CO_2 dissolving in water.

Equation 2: Show the dissociation of carbonic acid in water.

- What happens to the pH of the blood under acidosis conditions?
- What happens to the concentration of H_3O^+ (aq) under acidosis conditions?
- In which direction does the rate increase in equation 2?
- What happens to the concentration of H_2CO_3 (aq) as a result of the shift?
- In which direction does the rate increase in equation 1?
- What happens to the concentration of CO_2 (g) as a result of the shift?
- How do the lungs compensate for the changes in $[\text{CO}_2]$?

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8.4.2: Respiration Helps Maintain Blood pH During Alkalosis

1. Explain how respiration (lungs) helps maintain blood serum pH during alkalosis.

Write the two equilibrium equations responsible for maintaining pH

Equation 1: Show CO_2 dissolving in water.

Equation 2: Show the dissociation of carbonic acid in water.

- What happens to the pH of the blood under alkalosis conditions?
- What happens to the $[\text{H}_3\text{O}^+]$ (aq) under alkalosis conditions?
- In which direction does the equilibrium shift in equation 2?
- What happens to the $[\text{H}_2\text{CO}_3]$ (aq) as a result of the shift?
- In which direction does the equilibrium shift in equation 1?
- What happens to the $[\text{CO}_2]$ (g) as a result of the shift?
- How do the lungs compensate for the changes in $[\text{CO}_2]$?

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8.4.3: Kidneys Help Maintain Blood Serum pH

3. Explain how kidneys help maintain blood serum pH during acidosis.

Write the equilibrium equation responsible for maintaining pH

Equation:

- What happens to the pH of the blood under acidosis conditions?
- What happens to the concentration of H_3O^+ (aq) under acidosis conditions?
- In which direction does the equilibrium shift in the equation?
- What happens to the concentration of HCO_3^- (aq) as a result of the shift?
- What ion can the kidneys remove through the urine to correct acidosis?
- What ion can the kidneys add to the blood stream to correct acidosis?

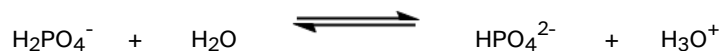
4. Explain how kidneys help maintain blood serum pH during alkalosis.

- What happens to the pH of the blood under alkalosis conditions?
- What happens to the concentration of H_3O^+ (aq) under alkalosis conditions?
- In which direction does the equilibrium shift in the equation?
- What happens to the concentration of HCO_3^- (aq) as a result of the shift?
- What ion can the kidneys remove through the urine to correct alkalosis?
- What ion can the kidneys add to the blood stream to correct alkalosis?

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8.4.4: Dihydrogen phosphate and Hydrogen phosphate Buffer System

5. The mixture of anions H_2PO_4^- / HPO_4^{2-} is an effective buffer with the following acid-base equilibrium.



- In acidosis, explain how the kidneys would compensate? What ion(s) appear in the urine? What ion(s) appear in the blood?
- In alkalosis, explain how the kidneys would compensate? What ion(s) appear in the urine? What ion(s) appear in the blood?

Even with all these controls, the pH of blood can still sometimes fall outside the normal range. The acid/base disorders are classified as being either respiratory or metabolic in nature. Respiratory disorders occur when exhaling does not remove CO_2 from the blood at the same rate that it is produced in cells. Metabolic disorders arise from the inability to remove acids produced by metabolism, inability to control concentrations of HCO_3^- , ingestion of compounds that are acids or bases or compounds that can be converted to acids or bases. The following lists conditions that can lead to either acidosis or alkalosis.

Acidosis:

- Lung disease, asthma
- Excessive alcohol consumption
- Ketosis, ketoacidosis (starvation, diabetic conditions)
- Holding your breath too long
- Hypoventilation (not enough CO_2 is exhaled)

Alkalosis:

- Excessive use of antacids
- Anxiety, fever
- Hyperventilation (CO_2 is blown off faster than it is produced in the cells)

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8.5: Post laboratory Assignment # 8

Post laboratory Assignment # 8

Now you have gained proficiency in identifying the shifts, increase or decrease in the concentrations of $H_3O^+(aq)$, $H_2CO_3(aq)$, $HCO_3^-(aq)$, $CO_2(g)$. Without the use of prompts answer in your own words the following questions.

1. Explain with chemical equation(s) and Le Chatelier's principle how the kidneys compensate for acidosis.
2. Explain with chemical equation(s) and Le Chatelier's principle how the lungs compensate for alkalosis.
3. How does cystic fibrosis upset the blood buffer system to cause acidosis? (Note: Gas exchange is slow and so $[CO_2]$ builds up. This also happens in asthma.)

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9: DENATURATION OF PROTEINS

9.1: Goals and Student Learning Outcomes

9.2: Introduction

9.3: Pre-laboratory Assignment # 9

9.4: Procedure and Observations

9.4.1: Effect of Alcohol on Albumin

9.4.2: Effect of Heavy Metal Ions on Albumin

9.4.3: Biuret Test

9.4.4: Effect of Heat on Albumin

9.4.5: Effect of Heat and pH on Casein

9.5: Post-Laboratory Assignment # 9

9.1: Goals and Student Learning Outcomes

In this experiment, the student will experiment on factors that lead to protein denaturation.

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9.2: Introduction

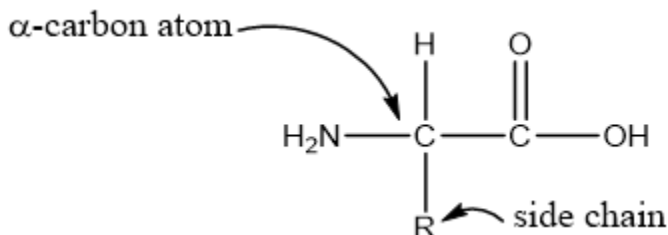
Introduction

Next to water, proteins are the most abundant substances in nearly all cells. They account for about 15% of the cells overall mass. Proteins are one of the three major classes of foods. Carbohydrates and fats are the others. Proteins differ from carbohydrates and fats not only in their functions in the living organism, but also in elemental composition. In addition to carbon, hydrogen and oxygen, which are present in carbohydrates and fats, proteins contain nitrogen. Most proteins also contain sulfur; sulfur is present in proteins due to the amino acids, cysteine and methionine. Additional elements such as phosphorus, iron, copper, zinc and iodine also occur in certain complex proteins. These elements are not part of the primary protein structure but are generally constituents of non-protein substances called prosthetic groups which combined with proteins.

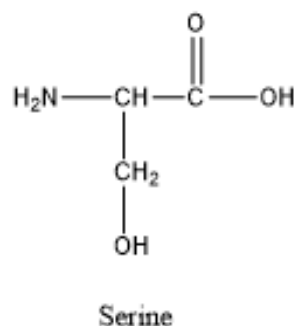
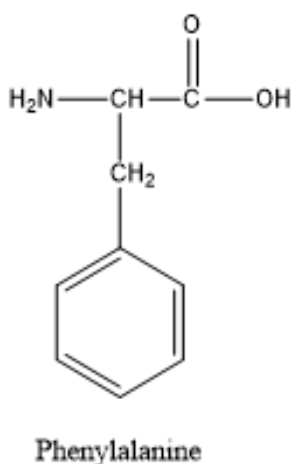
Amino acids

An amino acid is an organic compound that contains both an amine (-NH_2) functional group and a carboxylic acid (-COOH) group. Amino acids are building blocks for proteins. The amino acids found in proteins are always α -amino acids. The two functional groups are separated by one carbon atom and hence the symbol alpha (α). There are twenty common α -amino acids. Human beings cannot synthesize ten of them and they are called essential amino acid that must be supplied in the diet.

The general structural formula for an α -amino acid is:

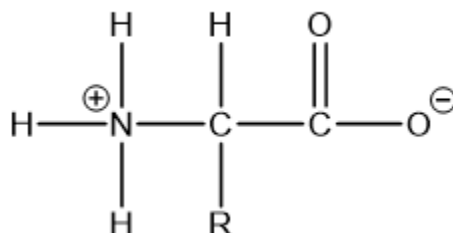


The group represented by R is the side chain. The side chain may be polar or nonpolar. For example, shown below are two amino acids. Phenylalanine has a nonpolar aromatic ring and serine contains a polar alcohol functional group in the side chain.



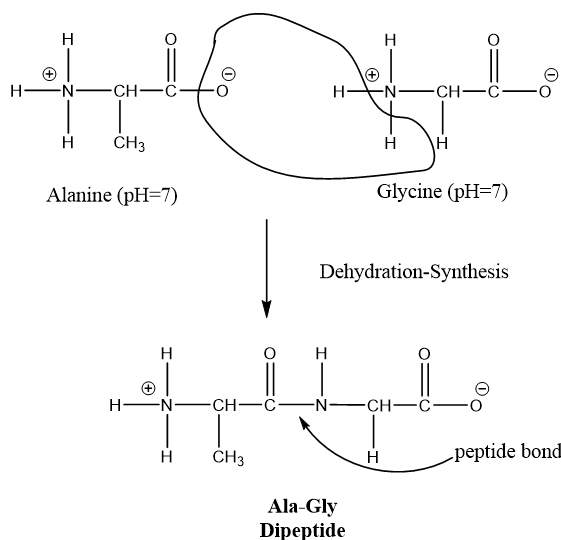
Acid-Base Properties of Amino Acids

The carboxylic acid functional group ($-\text{COOH}$) tend to donate protons (H^+) producing the carboxylate anion ($-\text{COO}^-$). The amine functional group ($-\text{NH}_2$) tend to accept protons (H^+) producing the quaternary ammonium cation ($-\text{NH}_3^+$). An internal acid-base reaction results in the formation of a zwitterion meaning “double ion”. A zwitterion has a positive charge on one atom and a negative charge on another atom. The zwitterion has no net charge.



Proteins and Peptide bond

Proteins are high molecular weight compounds made of amino acids. In a protein the amino acids are linked together by a series of peptide bonds. A peptide bond is an amide functional group obtained by elimination of a water molecule followed by condensation between two adjacent amino acids. Shown below is a dipeptide made of the two amino acids, Alanine and Glycine. The dipeptide called Ala-Gly is made via dehydration followed by synthesis reaction.



Protein structure

The **primary structure** of a protein refers to the order in which the amino acids are linked together by means of peptide bonds.

The three-dimensional configuration of segments of the protein chain is called its **secondary structure**. Common secondary structures are the α -helix and the β -sheet. The secondary structure is formed when amino acids hydrogen bond to other amino acids farther along the polypeptide chain.

The **tertiary structure** is the three-dimensional shape of the entire polypeptide chain. Globular proteins, for example, are very tightly folded into a compact spherical form. This folding result from interactions or bonds between the side chains of amino

acids. The interactions involve hydrogen bond, and hydrophobic interactions. Covalent bonds such as disulfide bonds and ionic bonds called salt bridges allow the protein to maintain its 3^o structure.

When proteins contain more than one chain, they have a quaternary structure. The manner in which the several peptide chains fit together describes a protein's **quaternary structure**. An example of a protein that has a 4^o structure is hemoglobin.

Denaturation of proteins

If the secondary, tertiary or quaternary structures of a protein are disrupted, the protein loses its biological activity and is said to be **denatured**. Denaturation usually results in precipitation of the protein from solution. The factors that may cause the denaturation of proteins are changes in pH, heat, alcohols, heavy metal ions, soaps, and detergents. In this experiment you will study the denaturation of albumin, the protein found in egg white. Casein is the name for a family of related proteins found in mammalian milk. Casein is not denatured with heat. This protein can be denatured by addition of acids during the process of making cheese.

Biuret test

The biuret test is used for predicting the presence of peptide bonds. The biuret test relies on the reaction between copper (II) ions and peptide bonds in a high pH solution. A violet color indicates the presence of proteins. Proteins give a strong biuret reaction because they contain a large number of peptide bonds. Dipeptides do not give a positive biuret test.

9.3: Pre-laboratory Assignment # 9

Name: _____

Pre-laboratory Assignment # 9

1. Review the chapter on proteins and this laboratory experiment prior to this assignment.
2. What is an amino acid? What two functional groups are found in all amino acids?
3. What are proteins?
4. Give the name of the functional group that connect the amino acids in a protein. _____
5. Give the name of the protein found in eggs? _____
6. Give the name of the protein found in milk? _____
7. What is denaturation?
8. What are some of the ways to denature a protein? List four.
9. List the intermolecular forces and chemical bonds responsible for the tertiary structure of the protein. List four.

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9.4: PROCEDURE AND OBSERVATIONS

Fill a 1000 mL beaker with 400 mL tap water and heat the water to around 90°C on a hot plate. Please clean 8 small test tubes and get them ready for the five tests. Use a red, round test tube rack to hold the test tubes in the hot water bath. The hot water bath is for parts d and e of the experiment and may be shared by students.

9.4.1: Effect of Alcohol on Albumin

9.4.2: Effect of Heavy Metal Ions on Albumin

9.4.3: Biuret Test

9.4.4: Effect of Heat on Albumin

9.4.5: Effect of Heat and pH on Casein

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9.4.1: Effect of Alcohol on Albumin

a. Effect of Alcohol on Albumin

1. Pour about 1 mL of albumin solution into two test tubes. Add 1 mL of rubbing alcohol to the first test tube. Save the second test tube as a comparison tube for this test and for subsequent tests.
2. Mix the components in the first test tube and let it sit for **5 min**. Compare the contents of the test tube with the second one which is the untreated albumin solution.
3. Record your observations of adding alcohol to the albumin solution in full sentences.

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9.4.2: Effect of Heavy Metal Ions on Albumin

b. Effect of Heavy Metal Ions

1. Pour 1 mL albumin solution in two test tubes. Number the test tubes, and add 10 drops of the following reagents to each test tube,

No.1: 1% silver nitrate (AgNO_3) solution

No.2: 1% sodium chloride (NaCl) solution

2. Mix the components in the two test tubes with a clean glass rod. Compare them to the control/comparison test tube from part a.

3. Record your observations in full sentences after adding AgNO_3 solution to albumin.

4. Record your observations in full sentences after adding NaCl solution to albumin.

9.4.3: Biuret Test

c. Biuret test

1. Pour 1 mL of the albumin solution into a test tube and add an equal volume of 10% NaOH solution.
2. Add about 5 drops of 0.5% CuSO₄ solution. Mix the contents of the test tube using a glass rod.
3. Record your observations in full sentences.

9.4.4: Effect of Heat on Albumin

d. Effect of heat on albumin

1. Pour about 1 mL of albumin solution a test tube. Warm the test tube in the hot water bath for about 10 min. Compare the contents of the test tube with the control/comparison tube.
2. Record your observations on heating the albumin solution in full sentences.

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9.4.5: Effect of Heat and pH on Casein

e. Effect of heat and pH on milk protein

1. Pour about 1 mL of **milk** in two small test tubes. Warm both test tubes for 5 min. Add 5 to 8 drops of vinegar to the first test tube. Use the second test tube of milk as control for observation.
2. Mix the components of test tube # 1 with a stirrer and compare with the control tube. Record your observations in full sentences.
3. Observations of the test tube with warm milk and vinegar:
4. Observations of the test tube with warm milk:

Turn off the hot plate, cool and dispose the hot water, return the beaker to your locker. Dispose of the contents of test tubes a through e in the chemical waste container. Wash the test tubes with tap water to remove the residue. Use the sink in the work bench. Rinse with small volumes of de-ionized water from the wash bottle in locker. The wash may be disposed down the sink.

9.5: Post-Laboratory Assignment # 9

Post laboratory Assignment # 9

Answer questions 2, 3, 4, and 5 based on the observations in the experiment.

1. Draw the polypeptide fragment in its line bond formula or condensed structural formula as it would appear at the **physiological pH**. Label C-terminus & N-terminus. Circle the peptide bonds.

Val-Asp-Ser-Ala

2. Why does denaturation happen with heat? Which level of structure(s) are disrupted when the protein **albumin** is subjected to heat?

3. Rubbing alcohol (isopropyl alcohol) a good disinfectant. What IMF is likely to be disrupted in a protein structure that results in denaturation?

4. Based on your observations which reagent is responsible for denaturation of protein silver nitrate or sodium chloride? Explain why.

5. Why is dilute silver nitrate solution applied to the eyes of newborn infants as an antibiotic?

6. Vinegar is a dilute acetic acid solution. Explain how an acid disrupts a salt bridge in the tertiary structure of a protein. Use the polypeptide drawn in questions 1 to answer this question.

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10: DNA AND RNA

- 10.1: Student Learning Outcomes
- 10.2: Introduction
- 10.3: Pre-laboratory Assignment
- 10.4: Report
- 10.5: Post-Laboratory Assignment

10.1: Student Learning Outcomes

Upon completion of this exercise, the student will be able to:

- deduce the sequence of the daughter DNA in replication
- construct the sequence of m-RNA in transcription
- assemble the protein that is synthesized in transcription
- analyze the polarity of the segment of the protein.

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10.2: Introduction

Deoxyribonucleic acid (DNA)

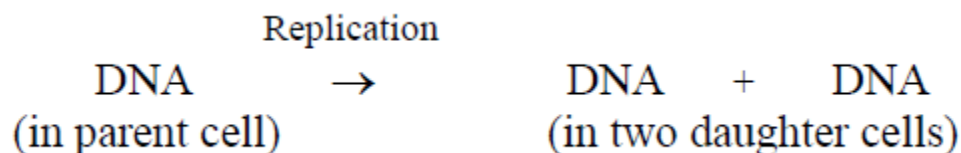
DNA is the hereditary molecule in all cells. The double helix structure of DNA has a backbone consisting of a five-carbon sugar (deoxyribose) and phosphate groups connected by phosphodiester linkages. Attached to the backbone are four nitrogen bases. The sequence of the four bases is the genetic code, the hereditary information of the cell. The bases of one strand of the double helix are hydrogen bonded to the bases of the other strand in a complementary manner. The base thymine (T) is always hydrogen bonded to adenine (A); cytosine (C) is always hydrogen bonded to guanine (G).

Nucleoside vs. Nucleotide

A **nucleoside** is the combination of a sugar with a nitrogen base. A sugar, a nitrogen base, and a phosphate group combine to form a **nucleotide**. DNA and RNA are polymers of nucleotides in which the sugars are deoxyribose and ribose, respectively.

Replication

When a cell divides, each daughter cell receives DNA identical to the parent cell through the process of replication. During replication of DNA, the double helical structure of DNA unwinds and each strand serves as the template for the synthesis of a new complementary strand of DNA.



Transcription of DNA to RNA

Each cell must synthesize protein and each protein must have a specific order of amino acids (the primary structure of the protein). The sequence of nitrogen bases in DNA contains the genetic information required to assemble the amino acids in the correct order. Another nucleic acid called the ribonucleic acid (RNA), copies the genetic information from DNA and uses that information to direct the synthesis of protein in the cell's ribosomes. Like DNA, RNA has a backbone made of a five-carbon sugar with ester linkages to phosphate groups; however, the sugar in RNA is ribose rather than deoxyribose. Four nitrogen bases are attached to the backbone; three of the bases (adenine, cytosine and guanine) are the same as in DNA. The RNA base uracil replaces the DNA base thymine. Several types of RNA exist, usually with a single-strand of nucleic acid.

The preliminary step in protein synthesis is a transcription of the genetic information from a segment of DNA to messenger RNA (mRNA). This step is similar to the replication of DNA but only a segment of the DNA chain is copied, not the entire chain. During transcription the nitrogen bases of DNA hydrogen bond to the nitrogen bases of mRNA. Cytosine (C) is hydrogen bonded only to guanine (G) and adenine (A) is always hydrogen bonded to uracil (U) or thymine (T).

Translation of mRNA (Protein synthesis)

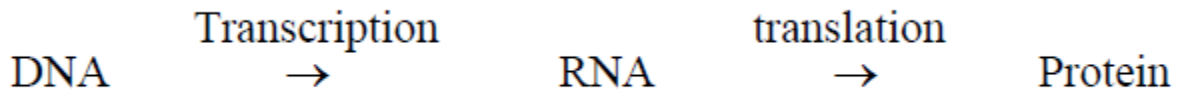
The genetic information is transcribed from DNA to mRNA in the nucleus of the cell. After being synthesized, the mRNA migrates to the cell's ribosomes.

Another form of RNA, called the transfer RNA (tRNA) is much smaller than other nucleic acids. Each tRNA molecule is specifically designed for a particular amino acid. The amino acid becomes attached to the end of its tRNA molecule. One after another, tRNA molecules hydrogen bond to mRNA; the amino acids carried by the tRNA are assembled in sequence to make a

protein. The order of the amino acid assembly is determined by the sequence of nitrogen bases in the mRNA. A sequence of three nitrogen bases on mRNA is a codon. The nitrogen bases of the mRNA codon, hydrogen bond to complementary nitrogen bases in the anticodon of the tRNA. Most of the 20 common amino acids have more than one codon. For example, UUU and UUC are both codons for the amino acid phenylalanine.

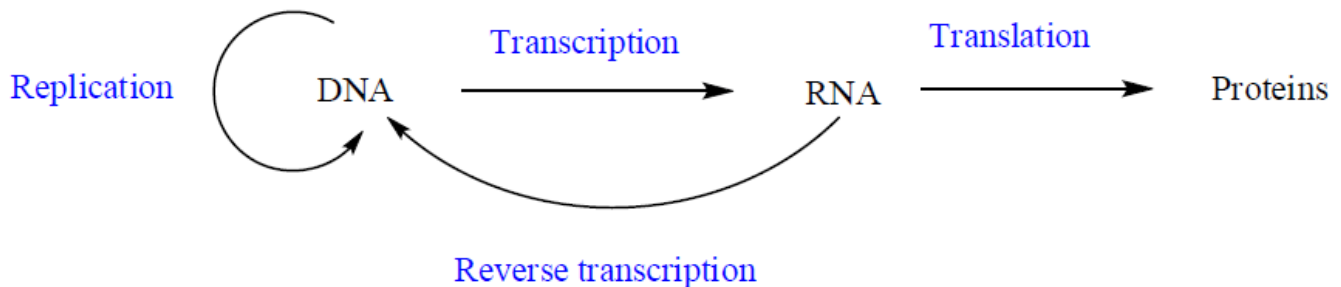
Summary

Transcription copies genetic information from DNA to RNA. The actual synthesis of the protein occurs during translation when the sequence of nitrogen bases is translated into the sequence of amino acids of the protein. The key to the translation is matching of specific amino acids to the codons of the mRNA using complementary anticodons on tRNA.



Reverse transcription “Why medications are not available to completely cure certain diseases”.

Some viruses called retroviruses do not contain DNA. One such example is the HIV (the virus that transmits AIDS) which is a RNA containing virus. This virus reproduces by taking over the host cell's machinery to produce new DNA that carries the information for the manufacture of viral nucleic acids and proteins. Therefore, the viruses thrive. Influenza, mumps and polio are other examples of RNA viruses. The enzyme involved in the reverse transcription is called reverse transcriptase. It is not as efficient as the DNA polymerase and so the virus rapidly mutates into new forms.



How does one control the AIDS virus?

The anti-AIDS drug zidovudine (AZT) has been around for a while and is a competitive inhibitor of an enzyme involved in the production of the HIV virus in the reverse transcription process. After being administered, AZT is converted to AZT triphosphate which closely resembles in structure one of the enzyme substrates dTTP responsible for the production of viral DNA from the viral RNA. The growing DNA chain now having incorporated an AZT triphosphate is unable to complete the synthesis of the DNA.

Mutation, genetic disease and the malaria connection

Any permanent change in the primary structure of the DNA is called mutation. Errors in replication, exposure to mutagens which include X-rays, UV radiation, nuclear radiation and chemicals are common causes of mutations. While some mutations have no effect on an individual or their off-spring, if these mutations take place in the egg or sperm cells it can be inherited and are called genetic diseases. There are over 4000 identified genetic diseases in humans. Some populations have higher rates of genetic anomalies. Sickle cell anemia is one such genetic disorder and results from a single amino acid change in a hemoglobin molecule. This disease is most prevalent among people with ancestors that were originally from lands that are close to the

equator where there are lots of mosquitoes. It is thought that a mild form of the disease protects people from malaria infection.

The woman behind the scenes

Rosalind Franklin made crucial contribution to the discovery of DNA although her name did not appear in the Nobel Prize awarded for this achievement. Her X-ray diffraction pictures of DNA were used by James Watson, Francis Crick and Maurice Wilkins as they correctly predicted the structure of DNA. Rosalind was working on the possible structure of DNA and even suggested the double helix structure but she did not publish her hunches, she was a very meticulous scientist and was doing very thorough analyses of her x-ray crystallography data. Even though they were working in the same lab, at King's College in the University of London, personality differences precluded Franklin and Wilkins from being collaborators. Thus, Wilkins showed her data to Watson and Crick (without her consent or knowledge) and this allowed them to complete the puzzle utilizing 3-D molecular models.

The Nobel Prize went to Watson, Crick and Wilkins for the discovery of how DNA encodes information that is passed from one generation to the next. This prize was awarded in 1962, four years after Franklin died from ovarian cancer. The Nobel Prize is not awarded posthumously, and it is awarded to no more than three scientists. Thus, there is some controversy over whether she would have shared the prize if she had lived long enough[§].

§Reference:

McGrayne, Sharon Bertsch, Nobel Prize Women in Science, Carol Publishing Group, New York, 1993 pp 304-332.

This exercise was modified from the following experiment without the use of the model kit recommended:

Experiment # 24: DNA and RNA; Laboratory Experiments; Chemistry and the living organism; 6th edition by Bloomfield and Stephens.

10.3: Pre-laboratory Assignment

Name: -----

Section: -----

Pre-laboratory for Exercise 10

Based on your knowledge of nucleic acids, answer the following questions.

1. List the names and symbols of the nitrogen bases found in DNA.

a. -----

b. -----

c. -----

d. -----

2. List the names and symbols of the nitrogen bases found in DNA and in RNA.

a. -----

b. -----

c. -----

d. -----

3. List the symbols of the nitrogen base pairs that hydrogen bond.

a. -----

b. -----

c. -----

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10.4: Report

Name: -----

Section: -----

Report for Exercise 10

1. Shown below is a base sequence in a segment of a DNA strand.

a. List the base sequence in a new DNA strand produced by DNA polymerase. What process is this called? Label the 5' and 3' ends.

b. Now list the base sequence on the mRNA molecule using the new DNA strand as template. What process is this called? Label the 5' and 3' ends.

c. List the codons in the 5' to 3' sequence of mRNA strand.

d. List the anticodons in the tRNA molecules that would hydrogen bond to the mRNA of 1c.

e. Using the standard methods of illustrating polypeptide structure (ex. Ala-Glu-Ser-Val) write the polypeptide formed from the m-RNA strand. What process is this called?

f. Draw the polypeptide fragment as it would appear at the physiological pH. Label the C-terminus and the N-terminus.

- g. In an aqueous environment, will the peptide fragment synthesized more likely be buried inside the tertiary structure or located on its surface? Why?
- h. What is most likely to change the biological activity of the protein e: Replacement of the leucine residue with isoleucine or the replacement of leucine with aspartic acid? Please discuss why.
2. What are the three base sequences on the mRNA molecule called? -----
3. What are the three base sequences on the tRNA called? -----
4. What happens to the mRNA after transcription prior to the synthesis of the polypeptide?
5. What happens to the polypeptide or protein after it is synthesized?
6. What is mutation?
7. Give examples of how a mutation can occur as a result of exposure to mutagens. List a couple of mutagens.
8. Must two mRNA molecules be identical to produce the same polypeptide? Explain.

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10.5: Post-Laboratory Assignment

Name: -----

Section: -----

Post Laboratory Assignment 10

Answer the following questions based on the segment of the DNA strand shown below:

5' → AGGAAATCAGACGAGGGA ← 3'

1. List the base sequence in the complimentary segment of the DNA strand. Label the 5' and 3' ends.
2. List the base sequence in the mRNA strand that is formed from the new DNA template strand in #1. Label the 5' and 3' ends. What is the name of this process?
3. Using the standard method of polypeptide structure, write the polypeptide that would be formed on the mRNA strand from question 2. Label the C-terminus and the N-terminus. What process is this called?
4. Draw the polypeptide fragment as it would appear at the physiological pH. (Ignore post-translational modification)
5. In an aqueous environment, will the peptide fragment more likely be buried inside the tertiary structure or located on its surface? Why?
6. What is most likely to change the biological activity of the protein above: replacement of a serine residue with threonine or the replacement of serine with phenylalanine? Please discuss why.

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11: PREPARATION AND TESTING OF SOAP (ALE EXP)

11.1: Student Learning Outcomes

11.2: Introduction

11.3: Safety

11.4: Pre-Laboratory Assignment

11.5: Procedure and Observations

11.5.1: Preparation of Soap

11.5.2: Testing of Soap

11.6: Post-Laboratory Assignment

This page titled [11: Preparation and Testing of Soap \(ALE exp\)](#) is shared under a [CC BY-NC 4.0](#) license and was authored, remixed, and/or curated by Deboleena Roy.

11.1: Student Learning Outcomes

In this experiment, the student will:

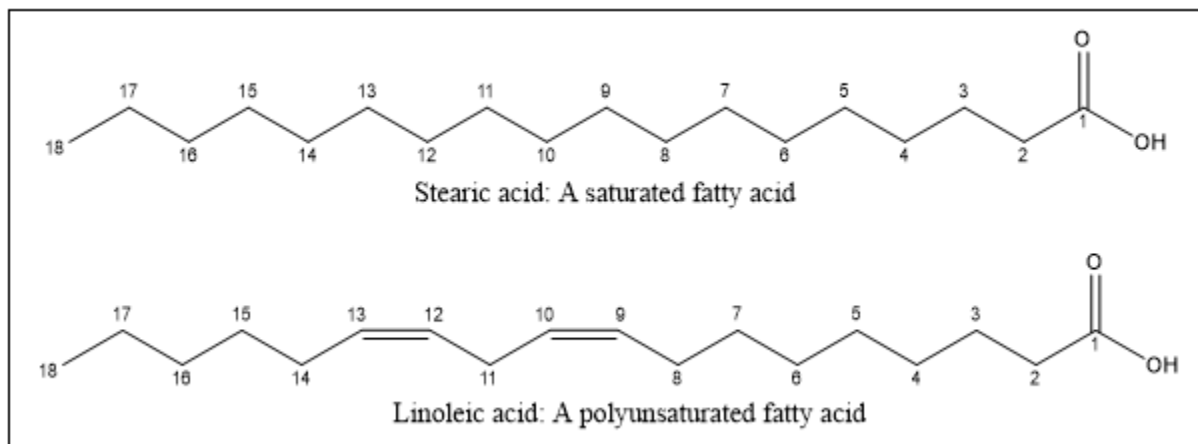
- make soap and compare the action of soap with hard and soft water
- enter individual observations on hardness of creek/river water samples.
- develop a sense of belonging in the STEM community.

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11.2: Introduction

Fatty acids

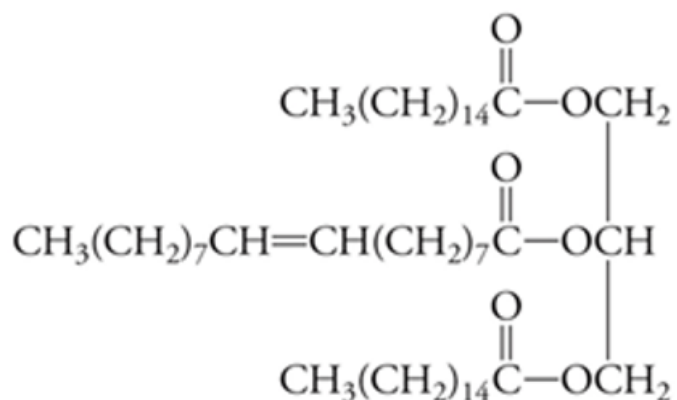
Fatty acids are long chain carboxylic acids that contain 12 to 20 carbon atoms. Fatty acids can be classified as saturated or unsaturated. Saturated fatty acids have only single bonds between carbon atoms. Unsaturated fatty acids contain one or more carbon-to-carbon double bonds. The fatty acids in animal fat tend to be higher in saturated fatty acids. The fatty acids in vegetable oils are generally higher in unsaturated fat but vary considerably from vegetable to vegetable.



Triglycerides

Triglycerides are tri-esters of glycerol and three fatty acids. About 95 percent of the lipids in foods and in human bodies are triglycerides. Triglycerides from plants tend to be liquids and are oils. Solid triglycerides from animals are fats.

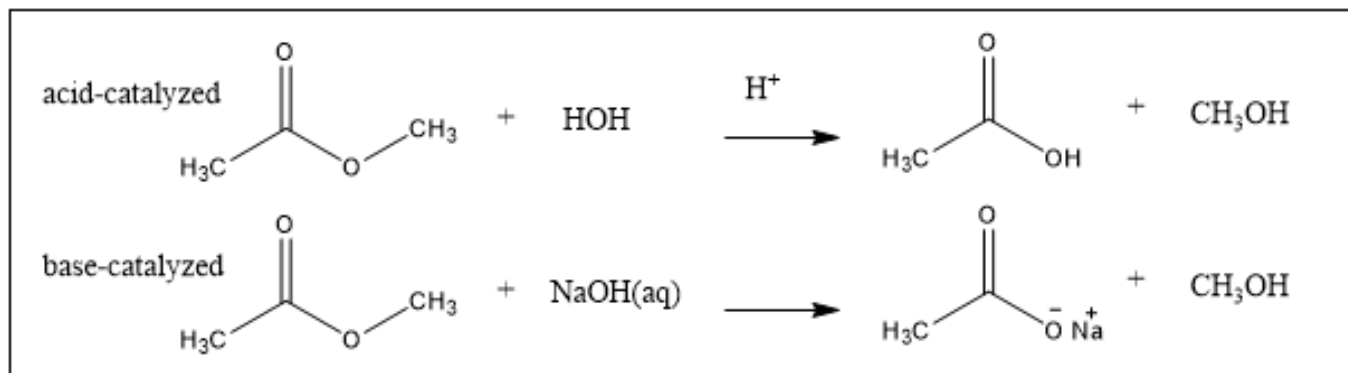
Shown below is a structure of a triglyceride from glycerol, two palmitic acids and an oleic acid. Such a triglyceride which is low in unsaturated fatty acids is likely to be a fat (solid) at room temperature.



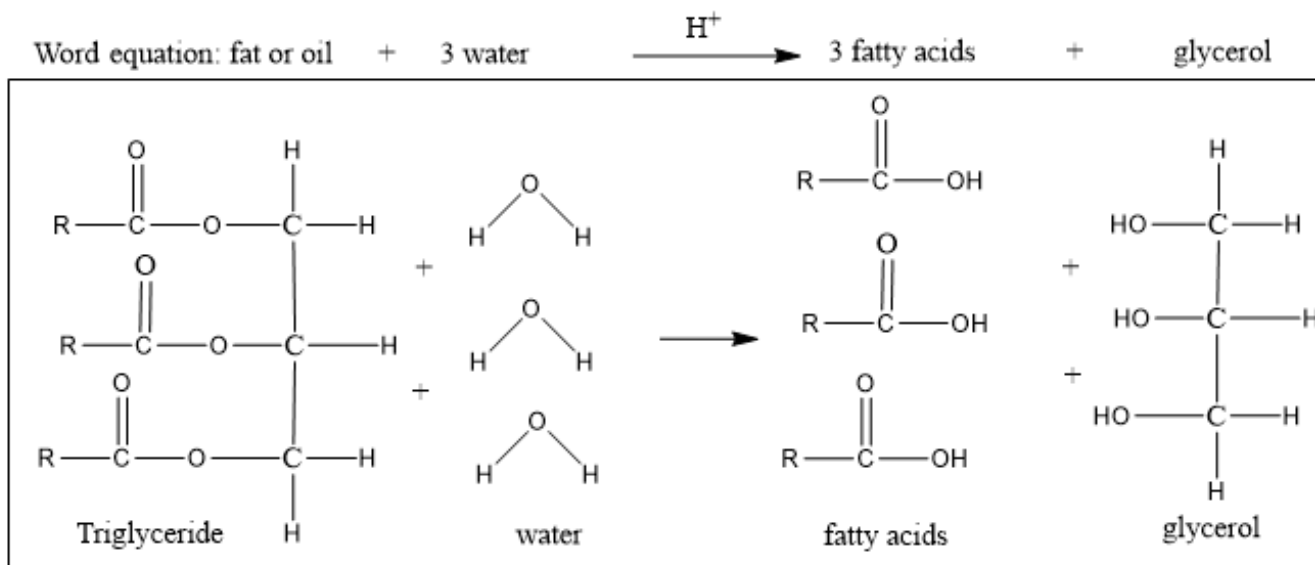
Some examples of fats include tallow in beef, chicken fat, butter, and lard. Tropical oils such as coconut oil and palm oil are high in saturated fats. Olive oil and canola oil are high in monounsaturated fatty acids while safflower oil, sunflower oil, and corn oil are high in polyunsaturated fats.

Hydrolysis of Esters

Hydrolysis is splitting apart with water. Hydrolysis can be catalyzed by an acid (H^+) or a base (NaOH). In an acid catalyzed ester hydrolysis, the ester is split to form a carboxylic acid and an alcohol. In a base catalyzed reaction, the sodium salt of carboxylic acid and an alcohol is produced.

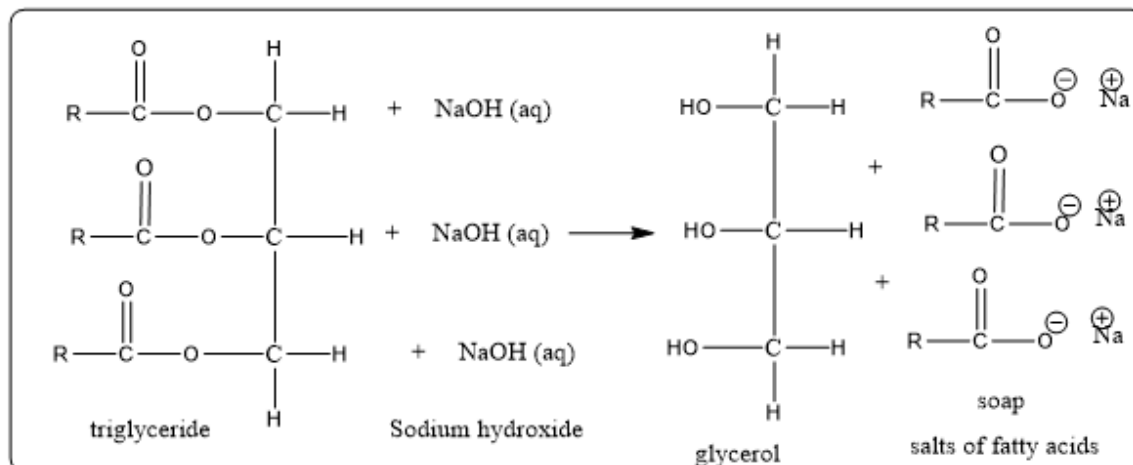


Hydrolysis of triglycerides (fat or oil) can be done in an acidic or basic solution. If this is done in the presence of an acid (H^+) then the products are three fatty acids and glycerol. The fatty acids are formed as carboxylic acids as shown below. The R in the triglyceride and fatty acids contain 11 to 19 carbons.



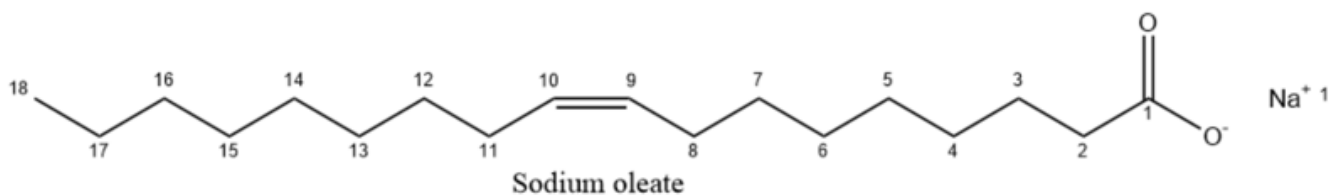
When hydrolysis occurs in the presence of a base such as sodium hydroxide, glycerol and soaps are formed. This is called saponification. Soap is the salt of a fatty acid. Shown below is the sodium salt of the fatty acid.

Word Equation: fat or oil + 3 sodium hydroxide (aq) $\longrightarrow$ glycerol + soap

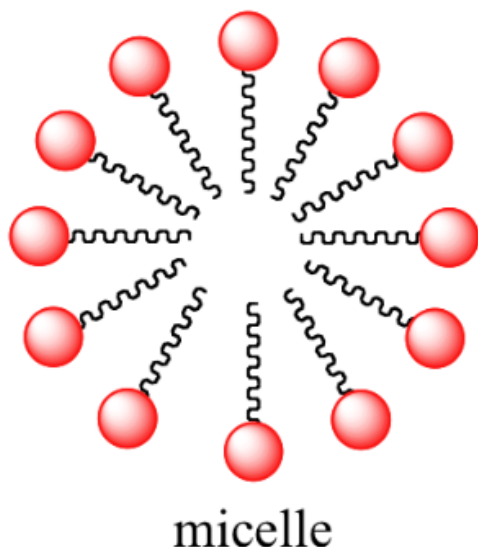


Soaps and Micelles

Fatty acids are hydrophobic. They are insoluble in water and have no cleansing properties. Soap is the salt of a fatty acid. Shown below is the skeletal formula of sodium oleate. Sodium oleate is soap and it is an amphipathic compound.



In aqueous solutions, soaps will spontaneously form **micelles**, a spherical structure that allows the hydrophobic tails to avoid contact with water and simultaneously form cluster in the center via London dispersion forces. The hydrophilic heads make up the outer surface and dissolves in water. Dirt and grime usually adhere to skin and clothing by combining with body oils, cooking fats, and lubricating greases. These substances are nonpolar and so they are not miscible in water. Washing with water alone will not remove them. These nonpolar oils and grease along with dirt will be trapped in the hydrophobic interior of the micelle of soap, while the micelle is easily rinsed away with water.

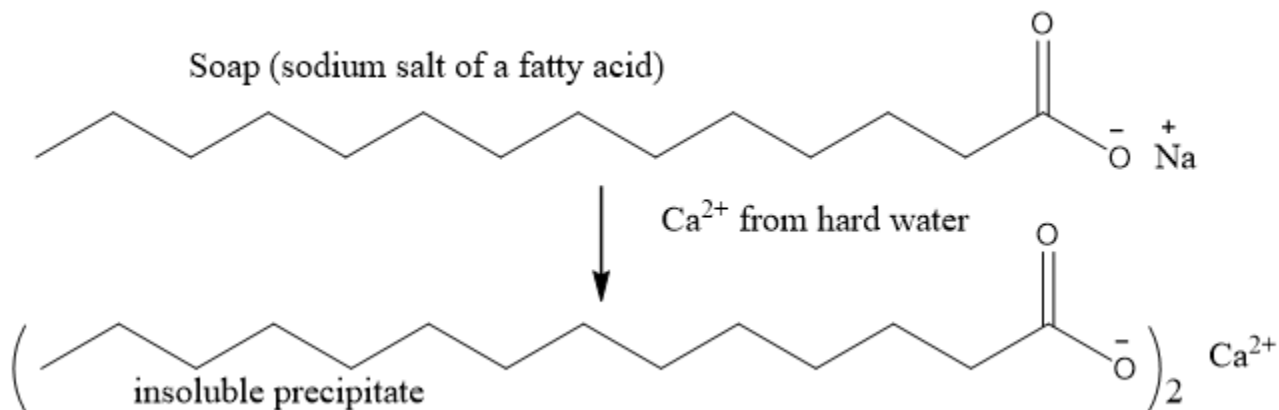


Hard water

Water is considered hard water, if it contains soluble Ca^{2+} , Mg^{2+} and Fe^{3+} ions. These ions are present in water as soluble bicarbonate compounds. If these ions are removed or absent, the water is considered soft water. Hard water can be purified by using water softeners or ion-exchange resins that replace Ca^{2+} and Mg^{2+} with Na^+ ions.

Soaps

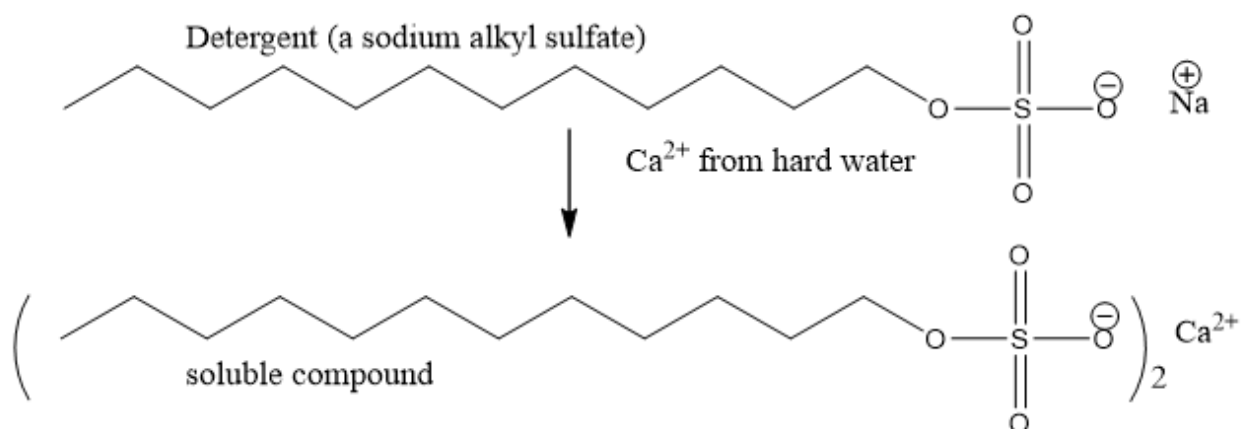
A soap is the sodium or potassium salt of a fatty acid. The tail part of the soap molecule is nonpolar, but the polar “head” at the carboxyl end permits a small portion of the molecule to be water soluble. Hence soaps are amphipathic. The cleansing power of soap is dependent on this water solubility. However, hard water containing Ca^{2+} and Mg^{2+} reacts with soap and forms insoluble scales (scum) with soap. This reduces the cleansing properties of soap since less soap molecules are available to form the micelle.



Detergents

A detergent is the sodium salt of a sulfonic acid. It has similar cleansing properties to soap since detergents can form micelles just like soap. However, detergents differ from soaps in its interactions with hard water. The calcium and magnesium ions from hard water interact with a detergent and form compounds that are soluble in water. So, calcium and magnesium ions from hard

water have no effect on the cleansing properties of a detergent.



11.3: Safety

CAUTION

Oils and fats are flammable. Heat the mixture slowly at low temperature.

Sodium hydroxide (lye) is very corrosive (caustic) to animal tissue, especially the hot, concentrated solution used this experiment. Use eye goggles to protect your eyes and gloves to protect your hands while making soap.

11.4: Pre-Laboratory Assignment

Name: _____

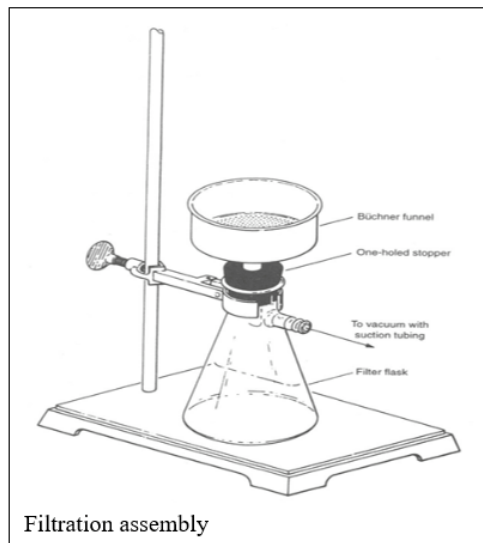
Pre-laboratory Assignment

Please read the introductions of this experiment and your textbook to complete this assignment.

1. What safety precautions are to be followed in this experiment?
2. What is an unsaturated fatty acid? Give an example.
3. Write the structural formula of the lipid (triglyceride) formed from glycerol and three molecules of stearic acid. Identify the nonpolar parts of the triglyceride and label with a box.
4. What is saponification?
5. What is a micelle? Explain how a soap micelle cleans grease.
6. What is the difference between hard water and soft water?

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11.5: PROCEDURE AND OBSERVATIONS



6. Dispose the liquid from the filter flask into the waste container in the hood. The cheese cloth may be disposed in the trash. Wash the funnel and the filter flask with tap water. Please reassemble the apparatus prior to departure from the lab.

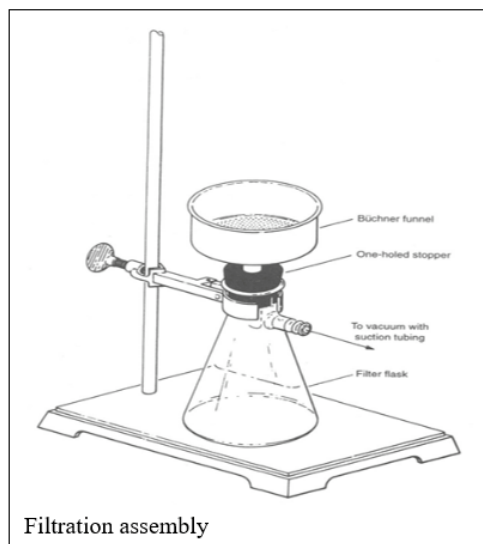
11.5.1: Preparation of Soap

11.5.2: Testing of Soap

11.5.1: Preparation of Soap

Saponification of a lipid-Preparation of Soap (Work in groups of 4 to 5 students.)

1. Place 10 g of canola oil in a 250-ml beaker, and add 40 ml of 50% sodium hydroxide solution. (CAUTION: The solution of sodium hydroxide is caustic. Please use gloves and goggles.) Coconut oil, vegetable shortening, lard, or corn oil may be used in place of canola oil.
2. Place the beaker of the oil and sodium hydroxide on a hot plate. Place a retaining ring around the beaker, to prevent the beaker from slipping. Use a stand to hold the retaining ring around the beaker. Place an evaporating dish filled with ice on top of the beaker, to prevent the escape of hot vapors. Heat the beaker for about 30 min on a hot plate at low heat (setting of 3 or 4). Stir the mixture with a glass rod twice at 10 min intervals. Replace the water in the evaporating dish with ice as needed.
3. When clumps of soap precipitate appear in the beaker, you are ready to do the next step. Add 50 ml of a saturated sodium chloride (NaCl) solution into the hot solution from step 2. Stir to mix. Sodium chloride is readily soluble in water and it reduces the solubility of soap in water. This process allows the soap to precipitate out.
4. Do not touch the strongly basic solution with bare skin. Wear gloves. Use a Buchner funnel set up to separate the soap from the solution it is suspended in. Use a layer of cheese cloth on the Buchner funnel to do filtration. The filter flask may be connected to a vacuum pump to increase the speed of filtration. Wash the soap with ice water. Lower temperatures reduce the solubility of a solid (soap) in water. This step removes excess lye and salt from the soap. Avoid adding the ice to the soap. Keep the pump on until the soap appears dry.
5. Save the lumps of soap in a clean 100 ml beaker for use in the next part. Label the beaker with the names of members of your group, the triglyceride used and the date of the experiment. Wash your hands and submit your beaker to your instructor. Check disposal directions on the next page.



6. Dispose the liquid from the filter flask into the waste container in the hood. The cheese cloth may be disposed in the trash. Wash the funnel and the filter flask with tap water. Please reassemble the apparatus prior to departure from the lab

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11.5.2: Testing of Soap

Cleansing properties of soaps (work in pairs)

1. Take a few clean lumps of soap in a 100 ml beaker and add 60 ml de-ionized water. Stir until you have a soap solution. It is not necessary to dissolve all of the soap. Let the solution stand for 2 min.

(If the soap solution is murky then redo this step with de-ionized water)

2. Add 5 mL of the following in six large test tubes.

1. de-ionized water
2. tap water
3. 0.1% CaCl_2 solution
4. 0.1% MgCl_2 solution
5. 0.1% NaCl solution
6. creek water/river water

3. Add 3 mL of soap solution to each. Stopper each of the test tubes and shake for about 30 seconds. Compare the foam/precipitate formed in each case and record your observations.

Solution added	Amount of foam produced with soap	Amount of precipitate produced with soap
Deionized water		
Tap water		
0.1 % CaCl_2		
0.1 % MgCl_2		
0.1 % NaCl		
Creek/River water		

Date/Time/Temperature of creek water collection: -----

Disposal: Please dispose contents of all test tubes in the waste container. Wash test tubes with tap water and rinse with deionized water before you store them in a locker.

11.6: Post-Laboratory Assignment

Complete your post laboratory assignment, write your name and date on page 1, and turn in the entire experiment for grading to your instructor.

Post Laboratory Assignment

1. Answer the following questions based on Part A of the experiment.
 - a. What lipid did you use to make soap? What are your observations of the product of the saponification reaction?
 - b. Why was saturated NaCl solution used in the preparation of soap?
 - c. Hydrolysis of esters can happen in acid or basic solutions. Why is sodium hydroxide used instead of hydrochloric acid in the making of soap?
2. Answer the following questions based on Part B of the experiment.
 - a. List the solutions tested that produced the most foam with soap.
 - b. List the solutions tested that produced the most precipitate with soap?
 - c. Which of the cations tested are usually present in hard water? Circle. Ca^{2+} Mg^{2+} Na^{+}
 - d. With the help of a chemical equation, please show the formation of insoluble scales with a hard water cation and soap.
 - e. Insoluble scales not formed with a detergent. Why?

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12: THE GAS LAWS

12.1: Goals and Student Learning Outcomes

12.2: Introduction

12.3: Pre-Laboratory Assignment

12.4: Procedure and Observations

12.4.1: Boyle's Law- Volume-Pressure Relationship

12.4.2: Gay-Lussac's law- The Pressure-Temperature Relationship

12.4.3: Charles' law- The Volume-Temperature law

12.4.4: Physiological Applications-The Lung Model

12.4.5: Solubility of Gas and Temperature

12.4.6: Henry's Law- The Pressure-Solubility of Gas Relationship

12.1: Goals and Student Learning Outcomes

- investigate Boyle's Law, Charles' law, Henry's Law, Gay-Lussac's law, relationship of solubility of a gas with temperature
- recognize physiological applications of Boyle's Law and Henry's Law.

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12.2: Introduction

There are three states of matter: gas, liquid, and solid. It is hard to overlook solids and liquids; however, gases are often invisible and so their presence can go unnoticed. Gases nevertheless are enormously important. We are immersed in a sea of invisible gases that make up the air of the earth's atmosphere. Unpolluted air is a mixture of gases that consists of 78% elemental nitrogen (N_2), and 21% elemental oxygen (O_2), 1% is mostly argon, and trace amount of other gases such as CO_2 and H_2O (in vapor form). Oxygen is a colorless and odorless gas which we need to survive.

The behavior of gases is simple and easy to predict compared to liquids and solids, because the particles (atoms and molecules) of a gas are far apart and do not interact with each other very much. Scientists quantified the behavior of gases some 300 years ago. All gases behave similarly. In this experiment we will investigate a set of laws that describe the behavior of gases.

Boyle's law

Boyle's Law states that the volume of a gas is inversely proportional to the pressure when the temperature remains constant. In simple terms, the volume of a gas increases as the pressure decreases; conversely, the volume of a gas decreases with an increase in pressure. In this experiment a vacuum pump is used to decrease the pressure of a gas (air) in a flask.

The physiological application of Boyle's Law is illustrated with a lung demonstration model. The model consists of two rubber balloons (lungs) attached to a Y-tube (trachea) inserted in a rubber stopper in the top of a bell jar (chest cavity). The rubber stopper end of the 'Y'-tube is open to the air (nose), this allows air to enter and leave the balloons. A thin sheet of rubber (diaphragm) is stretched across the bottom of the bell jar. The stopper and rubber sheet seal the bell jar air-tight. No air can enter or leave the bell jar; air already present in the bell jar can change in volume and pressure.



Gay-Lussac's law

Temperature is a measure of the average kinetic energy of the molecules in a sample of matter. At room temperature, molecules are moving rapidly. When temperature decreases, the molecules slow down. In a rigid container (one in which the volume remains constant), increasing the temperature will increase the pressure of the gas. Lowering the temperature will lower the pressure. This relationship is known as the Pressure-Temperature Law. The pressure of a gas varies directly with the absolute temperature provided that the volume remains constant.

In this experiment, a pressure gauge has been attached to a hollow, rigid steel ball. The hollow ball contains a gas. Any gas (such as nitrogen, oxygen, or helium) will work. Our steel ball is filled with air, a mixture of mainly nitrogen, oxygen and argon. The ball is immersed in water at different temperatures and the pressure of the gas in the steel ball is measured.

Charles's Law

In 1787, Jacques Charles, a French scientist, investigated the effect of changing the temperature of a given amount of air while holding the pressure constant. He found that when room temperature air was heated to 100°C , the volume increased; whereas when the air was cooled to 0°C its volume decreased. This relationship between volume and temperature is a direct and linear relationship.

Solubility of a Gas and temperature

Solubility is defined as the amount of solute that dissolves in a solvent at a given temperature. Gases are more soluble in solvent water at lower temperature than they are at higher temperatures. Carbonated beverages contain carbon dioxide gas in water or an aqueous solution of sugar syrup. As the cold beverage warms to room temperature, the CO_2 gas comes out of solution and the drink loses its fizz. On a hot day, fish may show signs of stress, because of low concentrations of oxygen in the water. Less oxygen from the air dissolves in the warm water than in cooler water.

Henry's Law

Pressure has a significant effect on the solubility of a gas. According to Henry's Law, the solubility of a gas in a liquid is directly proportional to the pressure of the gas on the liquid. Increasing the pressure increases gas solubility; decreasing the pressure decreases its solubility. An unopened bottle of a carbonated beverage has carbon dioxide (CO_2) gas dissolved in the aqueous solution. While dissolved, the gas is colorless and the mixture is homogenous. An open bottle of carbonated beverage has bubbles of carbon dioxide gas rising to the surface. These bubbles of gas are not dissolved and the mixture is heterogeneous. Opening the bottle decreases the pressure and decreases the water solubility of carbon dioxide. For this part of the experiment you will work with carbonated water.

This experiment was modified and adapted for the Chem. 309 curriculum by Deboleena Roy from the following source:

Experiment # 6; Properties of Gases; Laboratory manual; General, Organic and Biological Chemistry an integrated approach by Macaulay, Bauer & Bloomfield.

12.3: Pre-Laboratory Assignment

Please read the introduction and answer the following questions in full sentences:

1. How does the volume of a gas change when the pressure is increased?
2. How does the volume of a gas change when the temperature is decreased?
3. How does the water-solubility of a gas change when the pressure is decreased?
4. How does the water-solubility of a gas change when the temperature is increased?
5. How does the pressure of a gas change when the temperature is decreased?
6. What parts of the human body are best represented by the following pieces of the lung demonstration model:
 - a. The two rubber balloons -----
 - b. The Y-tube -----
 - c. The open end of the Y-tube -----
 - d. The bell jar -----
 - e. The thin sheet of rubber -----

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12.4: PROCEDURE AND OBSERVATIONS

12.4.1: Boyle's Law- Volume-Pressure Relationship

12.4.2: Gay-Lussac's law- The Pressure-Temperature Relationship

12.4.3: Charles' law- The Volume-Temperature law

12.4.4: Physiological Applications-The Lung Model

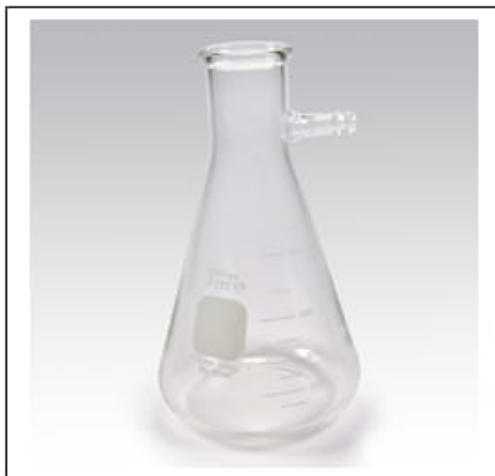
12.4.5: Solubility of Gas and Temperature

12.4.6: Henry's Law- The Pressure-Solubility of Gas Relationship

12.4.1: Boyle's Law- Volume-Pressure Relationship

1. Your instructor will have an apparatus set up by attaching rubber vacuum tubing from the side arm of a filtration flask to the vacuum pump. The vacuum tubing has thick walls to prevent its collapse.

Filtration flask



2. When you turn on the vacuum pump and place the palm of your hand on top of the flask, you should feel a strong suction due to the low pressure inside the suction flask if the vacuum pump is working properly.

3. Remove your hand and turn off the vacuum pump.

4. A small balloon is inflated to a size so that it can be placed inside the 250-ml suction flask. Place the stopper on the flask. Observe the behavior of the balloon in the flask while the vacuum pump is turned on.

5. Turn off the vacuum line. Detach the vacuum tubing from the vacuum line and allow air to enter the flask. Observe the behavior of the balloon.

6. Remove the balloon from the flask and place it in the beaker provided, Save the suction flask for the next set of students.

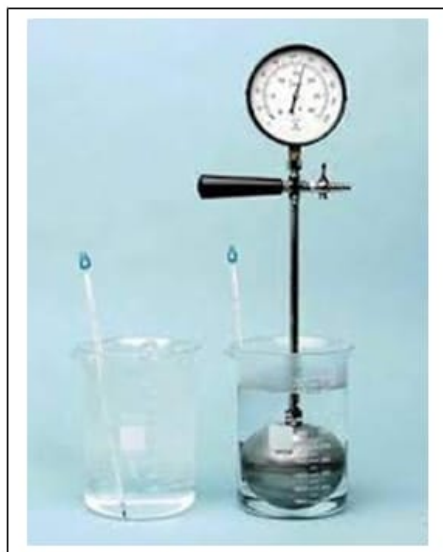
7. Answer questions a through c based on your observations in full sentences.

- What happened to the balloon in the suction flask when you turned on the vacuum pump?
- What happened to the balloon when you turned off the vacuum, detached the tubing, and increased the pressure in the flask?
- Explain your observations in questions a and b using the appropriate gas law.

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12.4.2: Gay-Lussac's law- The Pressure-Temperature Relationship

1. Measure the room temperature to 0.1°C with the help of a thermometer. Record the pressure of the gas sample in the steel ball.
2. Your instructor will have a 2000 mL beaker, $\frac{3}{4}$ -full of ice and tap water. Mix the ice and water with a glass rod. Measure and record the temperature of ice-water mixture with a thermometer to 0.1°C . Immerse the bulb of the steel ball in ice-water for 60 seconds and record the pressure.



3. Note the temperature of the hot water in a 2000 mL beaker in the fume hood. Immerse the bulb of the steel ball in the hot water for 60 seconds and record the pressure. At the same time record the temperature of hot water to 0.1°C .
4. Let the steel ball cool to room temperature for the next pair of students.
5. Fill in the table below with measurements and answer questions a through c based on your observations in full sentences.

Surroundings	Temperature of the surroundings and gas in steel ball ($^{\circ}\text{C}$)	Pressure of the gas in the steel ball (psi)
Room		
Ice and water		
Hot water		

- a. How does the pressure of the gas change as the steel ball is place in ice-water?
- b. How does the pressure of the gas change when the steel ball is place in hot water?
- c. Explain your observations in questions a and b using the appropriate the gas law.

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12.4.3: Charles' law- The Volume-Temperature law

1. Measure the room temperature with the help of a thermometer to 0.1°C .



2. Pull out the plunger of a 50-mL syringe, so that the syringe contains about 20 mL of air. Use a syringe cap to plug the hole in the tip of the syringe. Record the volume of the gas sample to 0.1 mL in the syringe as 20.0 ml in the table.

3. Use the warm/hot water in a 1000 mL beaker set up in the fume hood. Be sure there is sufficient water in the beaker to completely submerge the air contained in the syringe. Allow about 1 minutes for the air in syringe to attain the same temperature as that of the hot water and then record the new volume of the air in the syringe. At the same time measure and record the temperature of the hot water to 0.1°C .

4. Use the ice-water in a 1000 mL beaker set up for the experiment. Add more ice so the temperature is maintained at the freezing point of water. Allow about 1 minutes for the syringe of air to attain the same temperature as the ice/water bath and then record the temperature of the water bath and the volume of the air in the syringe at the same time. Allow the syringe, plunger, and cap to dry for the next pair of students by disassembling them.

5. Fill in the table below with your measurements and answer questions a through c based on your observations in full sentences.

Surroundings	Temperature of the surroundings and air in syringe ($^{\circ}\text{C}$)	Volume of the gas in the syringe (ml)
Room		
Hot water		
Ice and water		

a. How does the volume of the gas change as the syringe is place in hot water compared to room temperature?

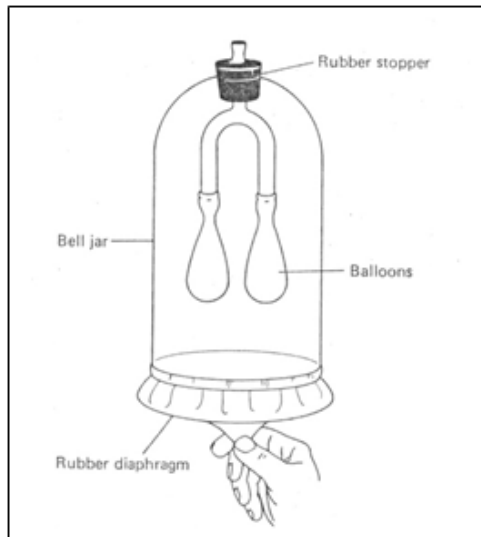
b. How does the volume of the gas change as the syringe is placed in cold water compared to hot water?

c. Explain your observations in questions a and b using the appropriate gas law.

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12.4.4: Physiological Applications-The Lung Model

1. You will work with a “lung demonstration model”.
2. Pull out the thin rubber diaphragm as illustrated in figure below. Care has to be taken not to puncture the rubber with finger nails. Observe the behavior of the balloons (lungs).



3. Release the rubber diaphragm and allow it to return to its original position. Observe the behavior of the balloons (lungs).
4. Answer the following a through c based on your observations in full sentences.
 - a. **Inhaling:** What happened to the balloons (lungs) when you pulled out the rubber diaphragm?
 - b. **Exhaling:** What happened to the lungs when the diaphragm returned to its original position?
 - c. Explain your observations in **respiration** with the help of a gas law.

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12.4.5: Solubility of Gas and Temperature

The following information is from *CRC Handbook of Chemistry and Physics* 75th (1994—95) edition. The information is on the solubility of oxygen in water.

Solubility of O₂ in water at 0°C is 4.89 cc O₂ per 100 cc H₂O

Solubility of O₂ in water at 25°C is 3.16 cc O₂ per 100 cc H₂O

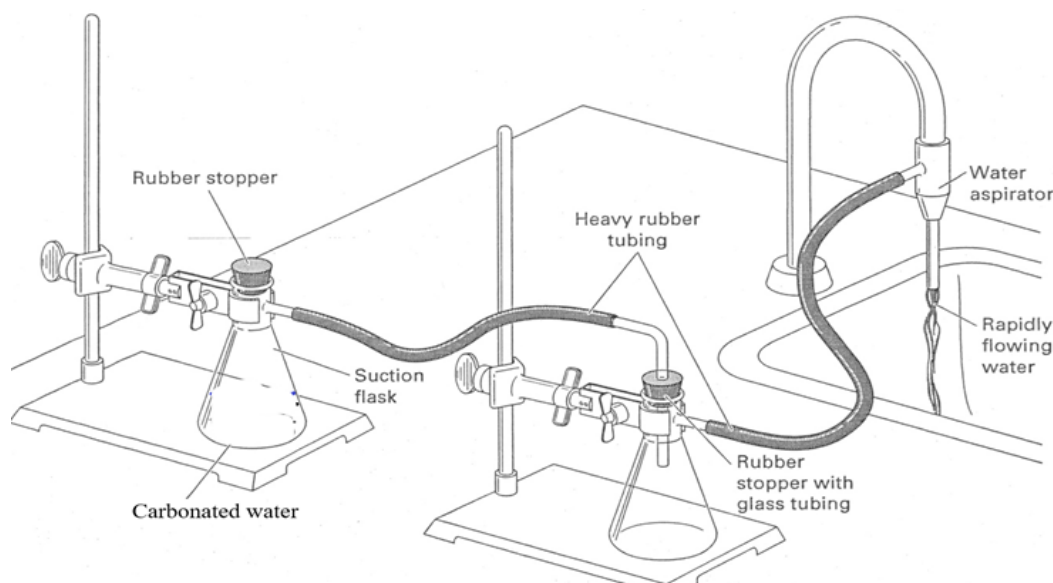
Answer questions a and b based on the data provided.

- a. How does increasing temperature affect the solubility of oxygen in water?
- b. “Thermal pollution” is the term used to describe hot water added to a lake or river by an electrical generating plant or other sources. Why is increased water temperature an environmental concern?

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12.4.6: Henry's Law- The Pressure-Solubility of Gas Relationship

1. Your instructor has an apparatus set up like the one shown below and will demonstrate this experiment for you.
2. Pour 200 ml of carbonated water into the suction flask. Connect the vacuum tubing to the vacuum line or the water aspirator. Run the vacuum line/aspirator at full force. This decreases the pressure inside the suction flask. Observe the behavior of the carbonated water at low pressure inside the flask.



3. Disconnect the vacuum tubing and observe the behavior of the carbonated water. Shut off the vacuum line or water. Dispose of the carbonated water down the sink and reset the apparatus for the next group of students.
4. Answers questions a through c based on your observations in full sentences.
 - a. What did you observe with the carbonated water when the pressure in the flask is decreased?
 - b. Explain your observations with the help of the appropriate gas law.
 - c. Explain how this observation and gas solubility law applies to the medical condition called the "bends".

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