CHEM3L Introductory Inorganic Chemistry Laboratory

1 units
3 hours Laboratory

Laboratory Manual

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1 1A																	18 8A
1 H Hydrogen 1.008	2 2A											13 3A	14 4A	15 5A	16 6A	17 7A	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.30	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.84	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.41	31 Ga Gallium 69.72	32 Ge Germanium 72.64	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9	46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3
55 Cs Cesium 132.9	56 Ba Barium 137.3		72 Hf Hafnium 178.5	73 Ta Tantalum 180.9	74 W Tungsten 183.8	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.1	79 Au Gold 197.0	80 Hg Mercury 200.6	81 Tl Thallium 204.4	82 Pb Lead 207.2	83 Bi Bismuth 209.0	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium (226)		104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (277)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (281)	111 Uuu Unununium (272)	112 Uub Ununbium (285)	113 Uut Ununtrium (284)	114 Uuq Ununquadium (289)	115 Uup Ununpentium (288)	116 Uuh Ununhexium (289)	117	118
	Lanthanides		57 La Lanthanum 138.9	58 Ce Cerium 140.1	59 Pr Praseodymium 140.9	60 Nd Neodymium 144.2	61 Pm Promethium (145)	62 Sm Samarium 150.4	63 Eu Europium 152.0	64 Gd Gadolinium 157.2	65 Tb Terbium 158.9	66 Dy Dysprosium 162.5	67 Ho Holmium 164.9	68 Er Erbium 167.3	69 Tm Thulium 168.9	70 Yb Ytterbium 173.0	71 Lu Lutetium 175.0
	Actinides		89 Ac Actinium (227)	90 Th Thorium 232.0	91 Pa Protactinium 231.0	92 U Uranium 238.0	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (262)

American Chemical Society's

Record Keeping Fact Sheet

The American Chemical Society's Committee on Patents and Related Matters has prepared this fact sheet as a guideline for maintaining complete research records. Such records are crucial to the advancement of invention and to the protection of intellectual property rights.

1. **Do** keep the record factual.
   Do record novel concepts and ideas relating to the work project.
   Don't editorialize.

2. **Do** use a record book with a permanent binding.
   Don't use a loose-leaf, spiral bound or otherwise temporarily bound book that provides for page deletion and insertions.

3. **Do** enter data and information including formulas and/or drawings directly into the record book promptly as generated.
   Do sign and date each page of the record book at the time the page was completed.
   Don't rely on memory or use informal loose sheets for entries with the intention of later putting these into the bound record book.
   Don't leave any completed page unsigned and undated.
   Don't postpone signing and dating all completed pages.

4. **Do** use permanent ink, preferably black, which will reproduce well when photocopied in black and white.
   Don't use pencil or non-permanent inks.
   Don't use colored ink.

5. **Do** write legibly.
   Don't make entries in handwriting that later on can be subject to interpretation, translation or wrong meaning.

6. **Do** identify errors and mistakes and explain them.
   Don't ignore errors and mistakes.
   Don't obliterate, delete or otherwise render errors unreadable.

7. **Do** completely fill each page.
   Do sign and date each page immediately after the last entry.
   Do draw vertical lines through unused portions of the page where an experiment takes less than a full page.
   Don't leave part of a page blank.

8. **Do** attach support records to the record book where practical; where volume and size prohibit this action, store such records, after properly referencing and cross-indexing, in an orderly form in a readily retrievable manner.

Don't file supporting records in a haphazard, helter-skelter manner without any record of their relationship or connection to the research reported in the record book.

9. **Do** use standard accepted terms; avoid abbreviations, code names, trademarks, trade names, or numbers if possible; if abbreviations or code names, trademarks, trade names or numbers are used, make certain these are defined at least once in every record book.
   Don't use any abbreviations, code names, trademarks, trade names, or numbers without giving its meaning or definition, or identifying the compound or the trademark and/or trade name and source.

10. **Do** keep the record book clean; avoid spills and stains.
    Don't subject the pages of the research notebook to chemical or physical destruction from spills.

11. **Do** see that the record is promptly witnessed by a knowledgeable person who understands what is being reported and, preferably, who assisted in or witnessed the work, but who is not a contributor to the research being conducted.
    Don't postpone having notebooks witnessed.
    Don't have notebooks witnessed by someone who is not technically skilled in the art being reported and who does not understand the contents of the record.
    Don't use a witness someone who has contributed professionally, conceptually, or technically to the work being reported.

12. **Do** maintain the confidentiality of the record until properly released.
    Don't treat the record book as a publication that is freely available to the public.

13. **Do** maintain control of an assigned record book at all times, keeping it in a fireproof safe, file or vault when not in use.
    Don't let the book lie open around the laboratory when not in use.
    Don't remove the record book from the company's or institution's premises.

14. **Do** index and close out the record book as soon as it is filled or a project is completed and check it back in for filing and storage to the person who issued it.
    Do reference the location where the book is being stored to assure ready retrieval.
    Don't keep a closed out and completed record book in the possession of the author.
    Don't file or store a book without referencing its location.

15. **Do** remember the record book is a legal document and should b3e treated as such and made available to your legal and patent counsel if needed.
    Don't keep a record book beyond the company's or institution's established record retention policy for such a record.

Darby & Darby

Proper Maintenance of a Laboratory Notebook

The United States laws implementing the general agreement on tariffs and trade (GATT) came into force on June 8, 1995. As a result, as of January 1, 1996, it has been possible for foreign inventors to establish a date of invention in a U.S. patent application by reference to knowledge or use of the invention in a foreign country which is a member of the World Trade Organization (WTO), or activity with respect to it in a foreign country which is a member of the WTO. Since United States patents are awarded to the first person to invent, not the first applicant, an applicant's ability to obtain a patent for his invention may often depend upon how well he documents inventive foreign activity with respect to his invention. This applies, not only to the situation in which an inventor becomes involved in an interference, but also when he wishes to avoid or "swear behind" a reference applied against an application by proving a prior date of invention.

Under the new law, it therefore becomes essential for foreign inventors to maintain an accurate and well-documented laboratory notebook.

The following guidelines should be followed with respect to all laboratory notebooks.

1. The notebook should have permanently bound pages which are consecutively numbered and should be used by a single engineer or scientist.
2. Ideas, calculations and experimental results should be entered into the notebook as soon as possible, preferably the same date they occur, so that the laboratory notebook becomes a daily record of the inventor's activities.
3. All entries should be made in the notebook in permanent black ink and should be as legible and complete as possible. Do not use abbreviations, code names or product codes without defining them clearly.
4. Draw a line through all errors. *Do not erase!*
5. Entries should always be made in the notebook without skipping pages or leaving empty spaces at the bottom of a page. If you wish to start an entry on a new page, draw a line through any unused portion of the previous page.
6. Never tear out or remove a page from the notebook.
7. Each page should be signed with the inventor's full name and dated. No entry should be changed or added to after signature. If the inventor has any additional information or corrections, a new entry should be made.
8. Each page of the notebook should be witnessed, signed and dated by a colleague who understands the inventor's work. This should preferably occur daily and certainly no less frequently than weekly. The witness should not be a direct contributor to the work being reported.
9. If an additional entry is made between the initial and final pages recording an experiment, the entry should identify the page on which the previous entry for that experiment occurs.

10. When the laboratory notebook is completely filled and is no longer required for reference, it should be indexed and stored in a safe location and, thereafter, handled in accordance with the company's established record retention and destruction policy for such documents.

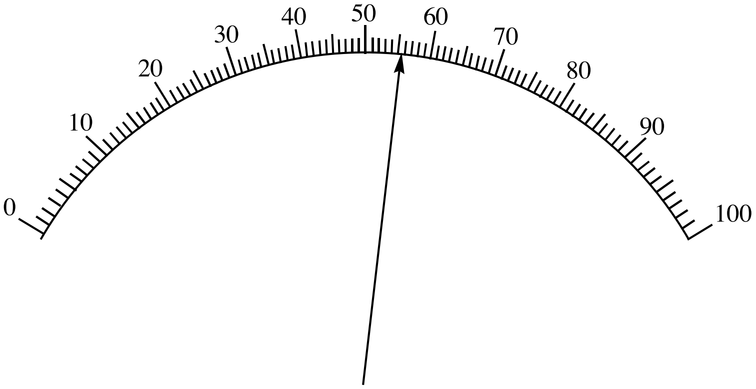
Some factors which reduce the value or credibility of your laboratory notebook:

- Illegible entries are totally worthless;
- Unsigned or undated pages are almost totally worthless;
- Notebook pages which have not been witnessed are almost as bad as unsigned and undated pages;
- A long delay between the signing of the page by the inventor and the witness raises question;
- Consecutive note book pages which are not dated in chronological order raise questions;
- Missing notebook pages raise questions;
- Erasures and deletions raise questions – instead, any later discovered mistakes should be corrected and explained on the next available blank page, referencing the page with the mistake.

In modern laboratories it is often the habit of engineers to maintain records of their work in computer files. We don not believe that computer files can provide sufficient evidence of invention. The reason for having invention records is to be able to prove the earliest date of invention. Since computer records can be updated and changed at will, and their dates are subject to tampering, they cannot serve as evidence that their content was created at a particular time. With a bound notebook, it is clear that the work occurred in a particular sequence and was witnessed by others. Also, scientific experiments can be conducted on the ink and paper in a notebook to prove their age.

Fundamentals of Measurement

In the analytical chemistry laboratory, many measurements are made, and the accuracy of these measurements obviously is a very important consideration. Different measuring devices give us different degrees of accuracy. A measurement of 0.1427 g is more accurate than a measurement of 0.14 g simply because it contains more digits. The former (0.1427 g) was made on an analytical balance, while the latter was made on an ordinary balance. A measurement recorded in a notebook should always reflect the accuracy of the measuring device. It does not make sense to use a very accurate measuring device and then record a number that is less accurate. For example, suppose a mass on an analytical balance was fond to be 0.1400 g. It would be a mistake to record the weight as 0.14 g, even if you know personally that the weight is exactly 0.14 g. Presumably, there are other people in the laboratory using the notebook, and your entry will be construed as to contain only two digits. The following example further illustrates this point.



The figure shows a meter, such as a pH meter, or readout meter on the face of some other laboratory instrument. The correct reading on this meter is 55.7. The temptation may be to write down 56. This latter reading is not correct in the sense that it does not reflect the accuracy of this measuring device. Measuring devices should always be used to their optimum capability, and this means recording all the digits that are possible from the device. The general rule of thumb for a device such as the meter in the figure is to write down all the digits you know with certainty and then estimate one more. Obviously, the meter in the figure shows a reading between 55 and 56, or "55 point something." This "something" is the estimated digit and is estimated to be a 7. The correct reading is 55.7. For digital readouts, such as with an analytical balance, this estimation is done for you by the device.

The digits that are actually part of analytical measurements have come to be known as "significant figures." A knowledge of the subject of significant figures is important from the standpoint that (1) one needs to know the accuracy of a measurement from just seeing it in a notebook (and not necessarily from actually seeing it displayed on the measuring device), (2) calculations are usually performed using the measurement, and (3) the correct number of significant figures must be shown in the result of the calculation.

To cover point number 1 above, the following rules apply.

1. Any nonzero digit is significant. Example, 1.27 (three significant figures.)
2. Any zero located between nonzero digits is significant regardless of the position of the decimal point. Example, 1.027 (four significant figures).
3. Any zero to the left of nonzero digits is not significant, unless it is covered also by Rule #2. Such zeros are shown only to locate the decimal point. Example, 0.0127 (three significant figures).
4. Any zero to the right of nonzero digits and to the right of a decimal point is significant. Example, 1.270
5. Any zero to the right of nonzero digits and to the left of a decimal point may or may not be significant. Such zeros may be shown only to locate the decimal point or they may be part of the measurement; one does not know unless he/she personally made the measurement. Such numbers are actually incorrectly recorded. They should be expressed in scientific notation to show the significance of the zero because then Rule #4 would apply.

   Example,

   1270 (incorrect)

   1.270×10^3 (four significant figures) or 1.27×10^3 (three significant figures)

To cover point number 2, the following rules apply:

1. The correct answer to a multiplication or division calculation must have the same number of significant figures as in the number with the least significant figures used in the calculation. Example, $1.27 \times 4.6 = 5.842 = 5.8$. Hence, 5.842 is the calculator answer, and 5.8 is the correctly rounded off answer.
2. The correct answer to an addition or subtraction has the same number of digits to the right of the decimal point as in the number with the least such digits in the calculation.

   Example,

   $$\begin{array}{r} 4.271 \\ +\ \ 6.96 \\ \hline 11.231 \end{array} \qquad \begin{array}{r:l} 4.27 & 1 \\ +\ \ 6.96 & \\ \hline 11.23 & 1 \end{array} = \boxed{11.23}$$

   11.231 is the calculator answer, and 11.23 is the correctly rounded answer.

3. When several steps are required in a calculation, no rounding would take place until the final step.
4. When both Rules #1 and #2 apply in the same calculation, the number of significant figures in the answer is determined by following both rules in the order in which they are needed, keeping in mind that Rule #3 also applies:

Example,

$$(7.27 - 4.8) \times 56.27 =$$

$$\begin{array}{ccc}
1^{st} & 2^{nd} & 3^{rd} \\
\begin{array}{r} 7.27 \\ -\ 4.8 \\ \hline 2.47 \end{array}
&
\begin{array}{r:l} 7.2 & 7 \\ -\ 4.8 & \\ \hline 2.4 & 7 = 2.5 \end{array}
&
\begin{array}{r} 56.27 \\ \times\ 2.5 \\ \hline 140.675 \end{array} = \boxed{1.4 \times 10^2}
\end{array}$$

138.9869 is the calculator answer, and 1.4×10^2 is the correctly expressed answer.

5. In cases in which a conversion factor that is an exact number is used in a calculation, the number of significant figures in the answer depends on all other numbers used in the calculation and not the conversion factor. To say that a number is exact means that it has an infinite number of significant figures and as such would never limit the number of significant figures in the answer.

   Example,

   $$1.247 \text{ m} \times 100 \text{ cm/m} = 124.7 \text{ cm (four significant figures)}$$

6. In cases in which the logarithm of a number needs to be determined, such as in converting $[H^+]$ to pH or in the conversion of percent transmittance to absorbance, the number of digits in the mantissa of the logarithm (the series of digits to the right of the decimal point) must equal the number of significant figures in the original number.

   Example,

   $$[H^+] = 4.9 \times 10^{-6} \text{ M}$$

   $$\text{pH} = -\log[H^+] = 5.31$$

   This would appear to be an increase in the number of significant figures compared to the original number (three in 5.31 and only two in 4.9×10^{-6}), but the characteristic of the logarithm, the digit(s) to the left of the decimal point, represents the exponent of 10, which serves to only designate the position of the decimal point in the original number and, as such, is not significant.

Kenkel, J. Analytical Chemistry Refresher Manual. 1.4 Fundamentals of Measurement. pp. 6 – 9; Lewis Publishers, London; **1992**.

Introduction to Small Scale Chemistry

Goals

- ☐ Introduce small-scale techniques
- ☐ Record both qualitative and quantitative observations
- ☐ Draw conclusions from results

Introduction

Small scale chemistry techniques involving a scaling down of reaction size and doing chemistry *in plastico* rather than using traditional glassware. Due to the decreased quantities of chemical reagents, the hazards, chemical waste and costs of experiments are also decreased. The obvious physical differences in small-scale chemistry leads to a very useful chemical difference. A distinct advantage to small-scale chemistry is the opportunity to work with reactions that would be difficult or dangerous on a larger scale.

"The drop is the container"

The surface tension of water allows us to replace our traditional test tube with drop size quantities, the drop being its own container. The volume of the drop can become reproducible with practice and consistent approaches.
Smaller amounts of reagents may require visual optimization using a hand lens (magnifying glass). This is usually kept at hand during all small-scale experiments.

"Good practice at any scale"

Consistent results that lead to meaningful conclusions require reproducibility. The most important technique in providing meaningful results in small-scale chemistry is avoiding contamination. The tip of any small scale pipet must be kept clean by avoiding contact with anything else, e.g., another drop, the reaction surface, another pipet. *Any contaminated pipet should be given to your instructor*. Reproducible results require *precision*. Whenever more than one trial or more than one sample is being tested for comparison, all of the trials/samples must be treated as similarly as possible. Many times students cannot control the *accuracy* of a balance but they can make sure that they use the same one throughout an experiment in an appropriate and precise manner.

Principles of Observation

Observations in the laboratory can be either quantitative or qualitative. Quantitative observations are called **measurements** and usually require the recording of a number and the unit of measure. Qualitative observations describe what is seen in the laboratory. Words that are more than superficially descriptive must be used when qualitative observations are made, e.g., the word "blue" does not convey anything specifically meaningful to someone who was not present. Adjectives describing color, size, vigor of reaction, or any other observable are a critical tool in the recording of experimental observations.

When reacting two substances together it is imperative that you monitor the reaction immediately upon mixing and continue to observe until no further change occurs. Some chemical reactions are slow and some are fast. Sometimes more than one reaction can occur as laboratory conditions may influence results (air, light, fumes, etc.).

"But what does it all mean?"

The interpretation of observations is what occurs in the formulation of **conclusions**. To simply state what you see is important. To extrapolate some meaningful conclusion from your data or sets of data requires more than observation.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact with all chemical reagents and dispose of reactions using appropriate waste container.

Materials

10-mL graduated cylinder, top-loading centigram balances, deionized water, transfer pipets, calculator, green food dye solution

Experimental Procedure

Techniques:

***Microburet "Pipet" Techniques I* – The Basics**

The microburet, often referred to as the pipet, is a versatile tool.

1. Fill 2 wells of a 24-well tray with water that has been colored with green food dye. The green dye solution is in a plastic bottle obtained from Reagent Central. Be sure to promptly return the bottle containing the green dye solution to Reagent Central.

2. Fill 2 wells of the 24-well tray with distilled water. The distilled water is available in plastic squeeze bottles located in a cabinet at the side of the room.

3. To make a microburet select a long stem pipet and cut the stem with your scissors at a point about 2 cm from the bulb. NOTE: *make the cut at right angles!*

4. Begin by cleaning the microburet. Draw a little distilled water up into the bulb, shake it so that all internal surfaces have been wetted.

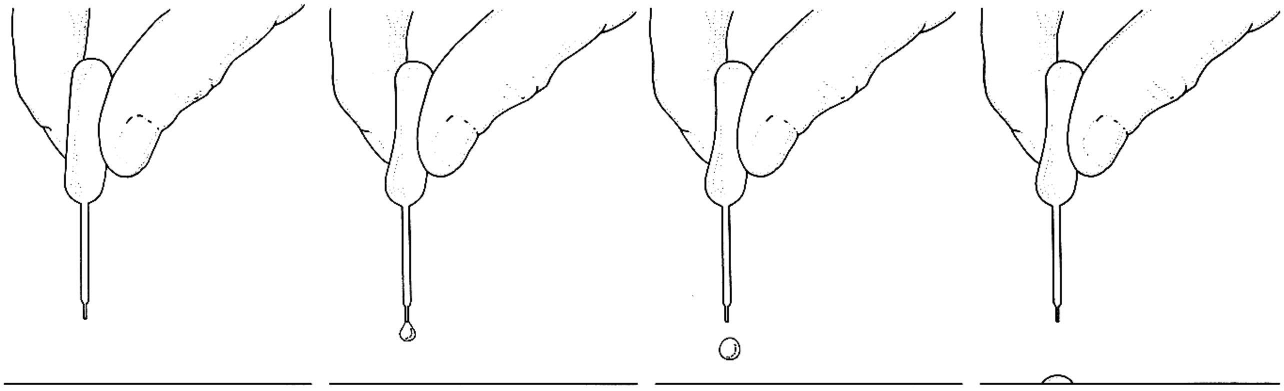
5. Holding the microburet vertically, expel to the waste cup. Press firmly to get the last drops out.

6. Now draw a little green food dye into the bulb, shake to rinse the bulb, and expel to the waste cup.

7. Finally, squeeze the bulb and draw green food dye into the microburet.

   Note: This sequence (steps 4, 5, 6, and 7) is known as *good wash, rinse, and transfer technique*.

8. Practice producing pools of various size on the labtop surface using the single drop *standard delivery technique*. To do so, hold the microburet *vertically* over the reaction surface as in the figure below, so its tip is about 1 to 2 centimeters above the reaction surface. Never touch the tip of the pipet to the reaction surface. "Touching" drops onto the reaction surface causes contamination of the contents within the microburet. Always drop the drops, never touch the drops onto the reaction surface. This technique, with the microburet held *vertically*, is called the *standard delivery technique*.



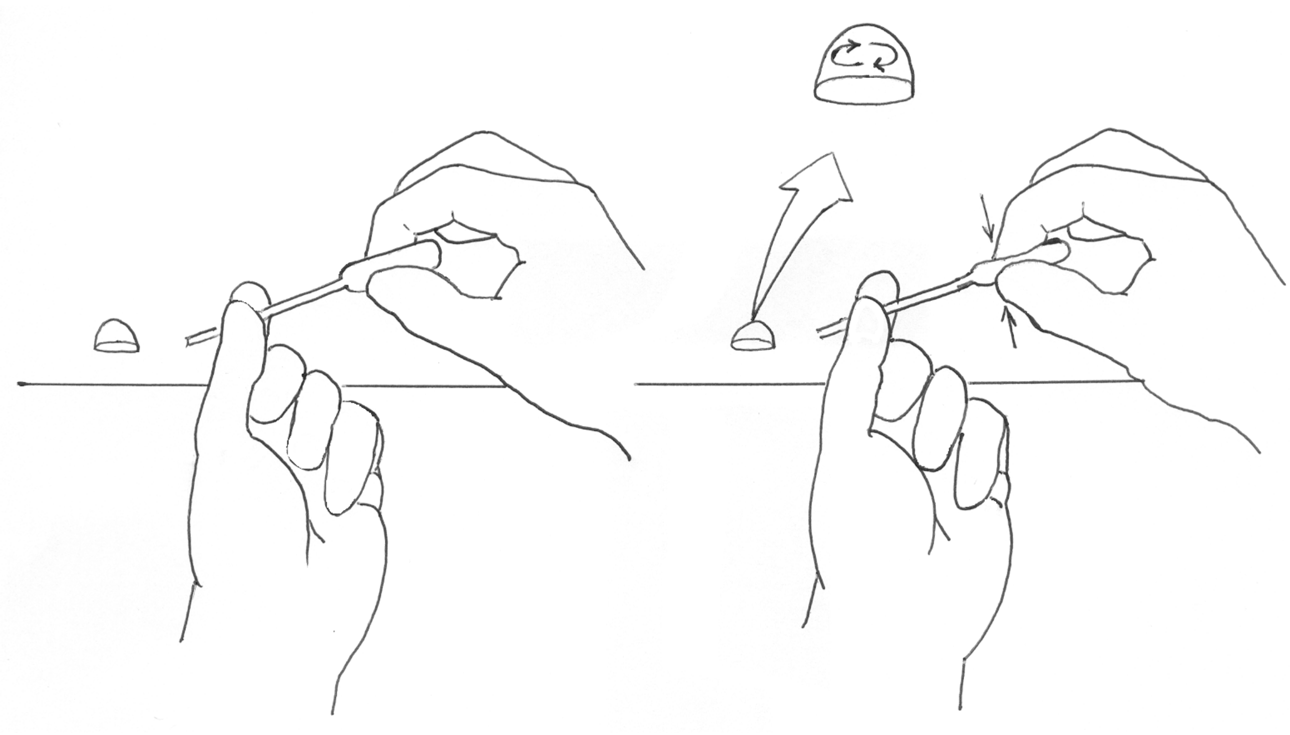
9. Make another microburet.

10. Follow the *good wash, rinse, and transfer technique* (steps 4-7) to fill the pipet with deionized water from one of the wells in the 24-well tray.

11. Drop a single drop of deionized water onto the reaction surface.

12. Now drop a drop of the green food dye solution on top of the drop of water, being careful not to touch the tip of the microburet to the deionized water drop.

13. Throughout this semester we will observe many chemical reactions between two substances conducted in this manor. While some visible change may accompany the initial mixing of two substances, often times thorough mixing is required to allow the reaction to go to completion. You can stir a hemispherical droplet on a reaction surface by gently blowing air from an empty long-stem pipet as shown below. Use one hand to steady the tip of the pipet and aim it at the side of the droplet. With the other hand gently squeeze the bulb repeatedly. The moving air will cause the droplet to swirl, mixing the contents without causing contamination.



14. Practice mixing drops of green food dye solution and deionized water until you are comfortable with the technique.

15. When finished, clean the reaction surface by first absorbing the liquid onto a small amount of microtowel. Dispose of the contaminated microtowel into the solid waste bins provided. Next wipe the surface with a paper towel and deionized water. Finally, dry the surface with a paper towel.

Microburet "Pipet" Techniques II – Affects of Drop Angle

The microburet is a simple constant-drop-volume delivery device when properly used.

1. Refill the microburet containing the green food dye solution.
2. Locate the 1 × 12 well strip in your chemistry kit. The well holds a volume of 0.40 mL when filled to the rim so that there extends a slight convex bulge above the rim of the well.
3. Count how many drops it takes to fill a single 0.40-mL well using the *standard delivery technique*. To do so, hold the microburet *vertically* (90°) over the well so its tip is about 1 to 2 centimeters above the well strip. Remember, never touch the tip of the pipet to the reaction surface. "Touching" drops onto the reaction surface causes contamination of the contents within the microburet. Always drop the drops, never touch the drops onto the reaction surface. Record your result.
4. Repeat the process to verify your result. Record your result.
5. Now tilt the microburet at 45° angle. Count how many drops it takes to fill the well with the microburet at 45°. Record your result.
6. Repeat the process to verify your result. Record your result.
7. Now tilt the microburet so that it is horizontal (0°) over the well and see how many drops are required in this position. Record your result.
8. Repeat the process to verify your result. Record your result.

The point here is that in order to obtain consistent results, you must be consistent in your delivery technique!

Using a balance

☐ Never place reagents directly on the pan of a balance.
☐ Use the same balance throughout an experiment.
☐ Record all of the numbers given in the digital output.
☐ Take into account the mass of the weighing container by:

"Weighing by difference" where you will weigh the empty container and record its mass. The mass of the object is obtained by weighing the container with the object and then subtracting the mass of the empty container, or

"Taring the balance" by placing the empty container on the pan of the balance and then pressing the "tare" button to zero the balance. You may now place the object into the container and the balance will read the mass of the object only.

Design an experiment to find the mass of one drop of water, delivered by the *standard delivery technique*, in units of milligrams. Repeat your experiment to verify your results. Be sure to record your procedure and results.

Introduction to Small Scale Chemistry

Name____________________________

Section____________ **Date**_______

Experimental Data: Record all your observations.

1. Affects of Drop Angle.

Number of drops to fill 0.40-mL well

	Trial 1	Trial 2	Trial 3
Standard Delivery Technique Vertical (90°)			
Halfway (45°)			
Horizontal (0°)			

2. Mass of a drop of water delivered by the *Standard Delivery Technique*.

Procedure:

Data:

Mass of a single drop ___________________

Chemical Changes

Goals

- ☐ Observe and record chemical changes.
- ☐ Design and carry out experiments to identify chemicals in consumer products.
- ☐ Use proper small-scale techniques to produce reproducible results.

Introduction

A chemical change involves the reaction of starting materials (reactants) to form new materials (products). Chemical changes are often visible to the naked eye in the form of color changes, a solid forming out of solution (precipitation) or gas formation. When two chemicals are mixed in the same proportions under the exact same conditions they will react in exactly the same way every time! By recording what you see when two solutions are mixed you document the chemical change. This information can be used to identify unknowns or to describe changes on a submicroscopic level.

Descriptions of liquids

Liquids and solutions have a number of properties that allow them to be distinguished from one another. Proper description of a liquid should include both the color and clarity. Water for instance is described as clear and colorless. It is clear because there are no particulates floating in the liquid and light is transmitted through it. Water is lacking in a color so it is colorless. Milk on the other hand is opaque and white.

Descriptions of precipitates

A precipitate is a solid material that results from the chemical reaction of two liquids or solutions. Precipitates are described on the basis of color, consistency and distribution. Precipitates greatly vary in appearance. Some precipitates may look like crystalline solids sitting at the bottom of a solution while others may look like muddy water (solid in suspension). To accurately describe solids we note their color followed by an appropriate adjective or two. Examples of appropriate adjectives are: milky, cloudy, sticky, clumpy, grainy, granular, free-flowing …… the list goes on.

In today's lab you will practice combining prepared solutions in a reproducible manner to observe whether a change occurs. You will study some of the chemicals in common consumer products. Based on all of your experience in mixing, you will try to make conclusions about the content of these consumer products.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

*Contact with silver nitrate ($AgNO_3$) will stain the skin.

Materials

Reagent Central chemicals include microburets filled with the following solutions:

sodium hydrogen carbonate ($NaHCO_3$)
sodium hydrogen sulfate ($NaHSO_4$)
phenolphthalein (phen)
starch
ammonia (NH_3)
sodium hypochlorite (NaOCl)
lead (II) nitrate ($Pb(NO_3)_2$)
copper (II) sulfate ($CuSO_4$)

hydrochloric acid (HCl)
FD&C No. 1 (blue dye)
potassium iodide (KI)
calcium chloride ($CaCl_2$)
sodium carbonate (Na_2CO_3)

This will damage your clothing!
silver nitrate ($AgNO_3$)
sodium hydroxide (NaOH)

Equipment: Empty pipet for stirring Lab top reaction surface

Experimental Procedure

1. Use small-scale microburets to put 2 drops of each chemical on the ×'s in the indicated spaces below. For background contrast, view the drops on black and white backgrounds provided by the ×'s. Stir each mixture by blowing air through an empty pipet (see p. 14). Record what you see in your lab notebook (see the Data Organization Suggestions section for further details). Do not clean your surface yet.

2. Test several foods for the presence of starch. If you don't know how to do this, answer the questions in the Data Analysis section. If you still don't know how to test for starch, ask your instructor.

3. Avoid contamination by cleaning up in a way that protects you and your environment. Carefully clean the small-scale reaction surface by absorbing the contents onto a small square of tissue paper or paper towel. Dispose of the paper in the appropriate waste bin. Wipe the surface with a damp towel and then dry it. Wash your hands with soap and water before leaving the lab.

Reaction Guide: Insert this page into the labtop. Mix one drop of each, using a long stem pipet to blow air past the droplet to complete the mixing.

a. $NaHCO_3$ + HCl	✕	
b. HCl + blue dye	✕	
c. NaOCl + blue dye	✕	Now add 1 drop of HCl.
d. NaOCl + KI	✕	Now add 1 drop of starch.
e. KI + $Pb(NO_3)_2$	✕	
f. $Pb(NO_3)_2$ + $CaCl_2$	✕	
g. $CaCl_2$ + $NaHSO_4$	✕	Be patient! Some chemical reactions are slow!

h. $NaHSO_4$ + Na_2CO_3	✕	
i. Na_2CO_3 + phen	✕	
j. phen + NaOH	✕	
k. NaOH + $AgNO_3$	✕	
l. $AgNO_3$ + NH_3	✕	Absorb this mixture onto a scrap of paper, expose it to sunlight, dispose of after making your observations.
m. $CuSO_4$ + NH_3	✕	
n. $CuSO_4$ + $NaHCO_3$	✕	

Data Organization Suggestions

Data is most useful when it is recorded neatly in tabular form (don’t crowed your work). This section is meant to be a guide only. You may use whatever format you like for recording data. However, the data in the laboratory notebook should be clearly organized so that anyone *knowledgeable* in chemistry can understand your observations.

Because we are looking at chemical changes it is beneficial to have initial observations and final observations so that the *change* is apparent. Try using the following format:

	Initial Observations Reactants				Final Observations Product
a.	$NaHCO_3$	+	HCl	→	Product
	clear & colorless solution		clear & colorless solution		clear & colorless solution with gas formation
b.	HCl	+	blue dye	→	Product
	clear & colorless solution		clear & blue solution		clear & green solution

For those reactions that involve three steps you can break them down as follows:

c.	blue dye	+	NaOCl	→	Product
	clear & blue solution		clear & colorless solution		clear & blue solution
	NaOCl/blue dye product	+	HCl	→	Product
	clear & blue solution		clear & colorless solution		clear & yellow solution

When testing foods for the presence of starch you must include a detailed description of your procedure along with the acquired data. Your procedure should be detailed enough so that any member of our class could reproduce your results exactly! Follow a similar data table as above to organize your observations.

Data Analysis

Answer the following questions in your laboratory notebook using complete sentences.

1. Sodium hydrogen carbonate, $NaHCO_3$, is also known as sodium bicarbonate or more commonly as baking soda. When HCl is added to sodium bicarbonate, bubbles of carbon dioxide form. Write the formula for carbon dioxide. What common consumer products contain this gas?
2. Which of the other combinations form bubbles?
3. The body uses hydrochloric acid, HCl, to help digest food. Where in the body do you think HCl is found? What color does blue food dye turn when HCl is added?
4. Sodium hypochlorite, NaOCl, is a common ingredient in household bleaches and cleansers. What happened to the color of blue dye when both HCl and NaOCl are added?
5. Potassium iodide, KI, is the source of iodine in iodized salt. What color is the KI + NaOCl mixture? What color does starch change to in the presence of KI and NaOCl?
6. A precipitate is a solid that separates upon mixing solutions. Which reaction produced a very bright yellow precipitate?
7. Which other mixings produced precipitates? Describe their colors and textures appropriately.
8. Which mixture produced a precipitate that was very slow to form?
9. Which solutions produced a distinctive brown precipitate? Describe that color.
10. Look at the scrap of paper you used to absorb the silver nitrate and ammonia mixture. What evidence do you see that indicates that silver compounds are light sensitive? Can you think of any way that this chemical change could be used for some application or useful purpose?
11. What were three observations that indicated the formation of a new substance?
12. Which foods contained starch? Is this consistent with what you would have predicted from your personal knowledge of food science? Explain.

Name____________________________

Section____________ **Date**_______

Electronic Configuration Worksheet

Scientists have learned from experiment that electrons *in atoms* occupy specific energy levels. These energy levels form an orderly pattern, and once this pattern is understood, it is rather simple to figure out, for any element, where the electrons are, i.e., which energy levels they occupy. Once we know this electron arrangement (also called the electronic configuration), we can predict a great deal about how these electrons will behave in chemical reactions. And this is the basis for chemistry--a fundamental understanding of the nature of atoms so that we can begin to understand and predict what they do.

Look at the diagram on the next page. It shows major, or principal energy levels. These are numbered consecutively 1, 2, 3, ...; the first 7 are shown on the diagram. It also shows that each of these levels is split into sublevels. The first principal energy level consists of a single sublevel; the second one has two sublevels, and so on. For levels 5, 6 and 7, not all the sublevels are shown. The sublevels are designated by letters: s, p, d, f, g, h, i, ... (the reason for choosing these letters is historical.)

Notice that as energy increases, the levels become more and more closely spaced. Although the exact spacing changes from element to element, the same general pattern exists for all elements--a large gap between levels one and two, a smaller gap between levels two and three, and so on.

Exercise:

Complete the list of sublevels For principal energy level 4:

4_s_, 4___, 4___, 4___.

List the sublevels for principal energy level 6: ___, ___, ___, ___, ___, ___

What sort of experiment indicates the existence of discrete energy levels?

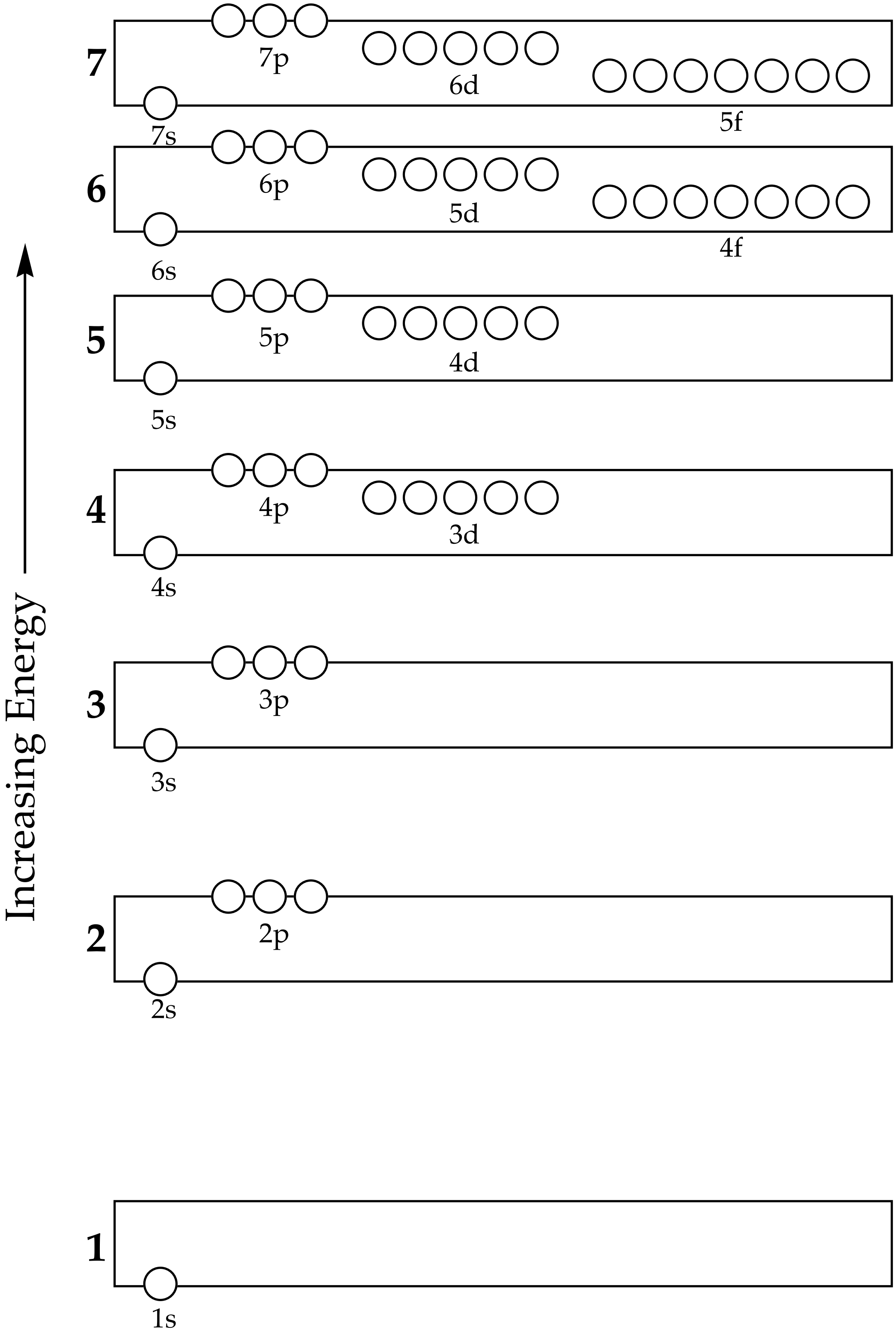
The energy levels and sublevels by themselves do not tell the whole story. Because the spacing between energy levels becomes smaller and smaller as energy increases, the sublevels of different principal energy levels overlap each other. Note on the diagram on the next page that the 3d sublevel, for instance, is higher in energy than the 4s.

One way to learn the order of the sublevels is simply to remember the arrangement on the diagram on the next page. Another is to remember the arrangement shown in Figure 11-14 on p. 467.

Exercise:

Copy the diagram from Figure 11-14
on p. 467 in the space at the right.

Electronic Energy Levels



No orbitals higher in energy than the 7p are needed to describe the unexcited electrons in every known element (as of 2003).

The next step is to find out how many electrons can fit into each sublevel. The answer to this question comes from theory, but is verified by experiment. Sublevels are divided into orbitals. An *orbital* is not the same as an *orbit*. An orbit is a well-defined path, like the earth's orbit around the sun. An orbital is a mathematically defined volume of space in an atom, molecule or ion, within which an electron can be found some specified percentage of the time. Electrons in orbitals in a given sublevel have the same energy.

Exercise:

Fill in the answers (number and diagram) for f, g, and h sublevels, below:

s sublevels have 1 orbital __
p sublevels have 3 orbitals __ __ __
d sublevels have 5 orbitals __ __ __ __ __
f sublevels have ___ orbitals
g sublevels have ___ orbitals
h sublevels have ___ orbitals

Experiments have shown us that *each orbital may hold a maximum of 2 electrons.* When two electrons occupy a single orbital, their spins are aligned in opposite directions: ↑↓

Thus an orbital may have no electrons (empty): __; one electron: ↑ , or two electrons: ↑↓.

Exercise:

For each of the following sublevels, indicate a diagram for the maximum number of electrons:

orbital	how many electrons?
2p ↑↓ ↑↓ ↑↓	6 electrons
2s	___ electrons
4p	___ electrons
3d	___ electrons
5f	___ electrons
6s	___ electrons
5g	___ electrons
6f	___ electrons

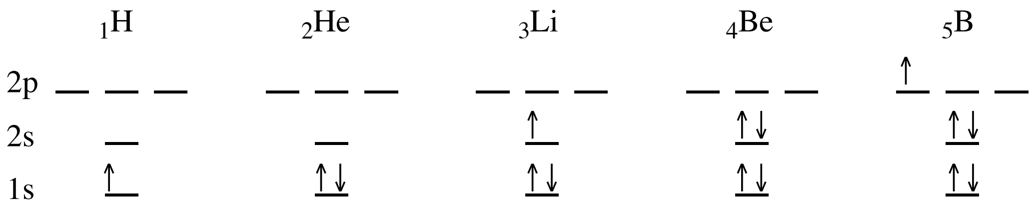
What is the maximum number of electrons in principal energy level 1 _________, principal energy level 2 _________, principal energy level 3 _________?

We have now looked at the basic rules for energy levels, sublevels, and orbitals. Now we will use this information to find the arrangement of electrons of different elements. These arrangements are called *orbital diagrams*.

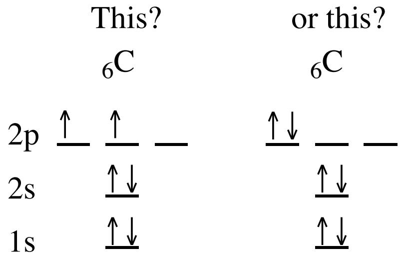
Some guidelines:

1. The energy sequence of the sublevels is the same for all elements.
2. Electrons go into the lowest energy orbital that has room available. This is the "Pauli Exclusion Principle."
3. The maximum number of electrons in an orbital is 2.

We list the first few elements and their sublevels. Electrons fill orbitals from the bottom up:



The fifth electron in boron (B) could go into any one of the three 2p orbitals. But when we come to carbon (C), we have two distinct choices for the sixth electron:

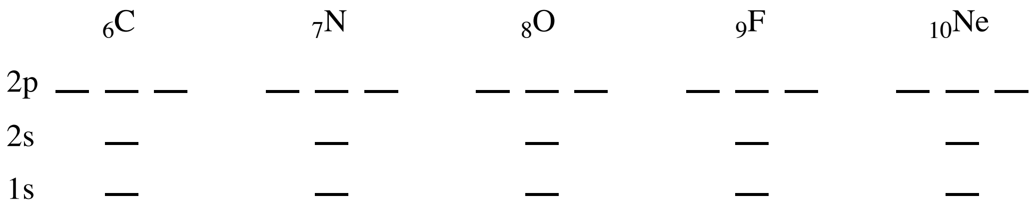


The answer again comes from experiment, and gives us Guideline #4:

4. Electrons will occupy orbitals of equal energy singly before pairing up. This statement is known as Hund's Rule of Maximum Multiplicity, or, more simply, as "Hund's Rule."

Exercise:

Fill in the electrons of the second row elements from N through Ne:



(The 2p electrons in carbon can occupy any two of the 2p orbitals.)

Exercise:

Fill in the electrons for the third-row elements:

	$_{11}$Na	$_{12}$Mg	$_{13}$Al	$_{14}$Si	
3p	— — —	— — —	— — —	— — —	3p
3s	—	—	—	—	3s
2p	— — —	— — —	— — —	— — —	2p
2s	—	—	—	—	2s
1s	—	—	—	—	1s

	$_{15}$P	$_{16}$S	$_{17}$Cl	$_{18}$Ar	
3p	— — —	— — —	— — —	— — —	3p
3s	—	—	—	—	3s
2p	— — —	— — —	— — —	— — —	2p
2s	—	—	—	—	2s
1s	—	—	—	—	1s

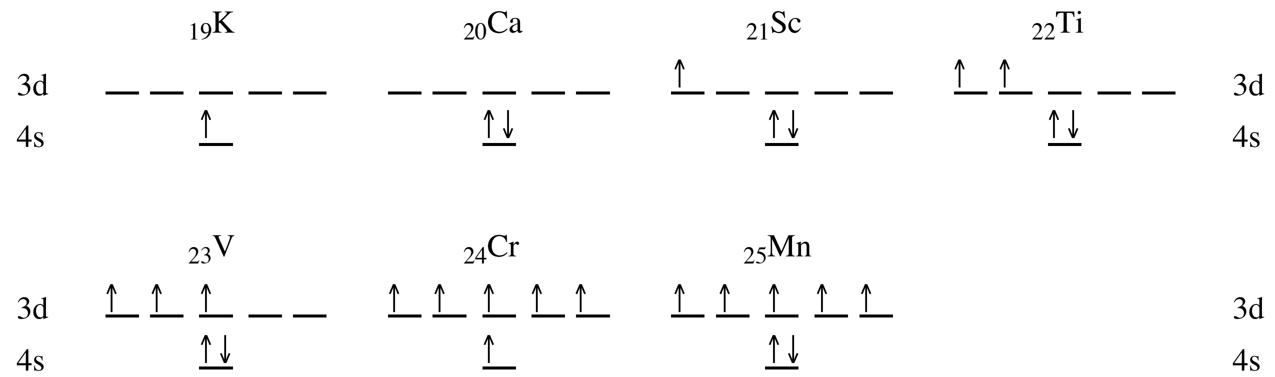
In addition to the orbital diagrams the electron arrangement can be written in shorter form called *electron configuration*:

H	He	Li	Be	B	and so on.
$1s^1$	$1s^2$	$1s^22s^1$	$1s^22s^2$	$1s^22s^22p^1$	

Exercise:

Write the short notation electron configurations for elements 6-18:

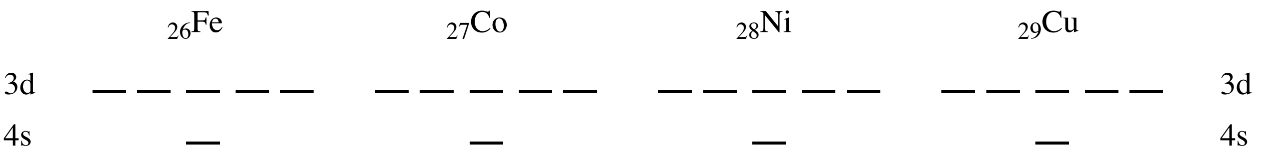
Beginning with K, the fourth period elements fill in the 4s, 3d and 4p sublevels, in that order. In the orbital diagrams below, the first 18 electrons (in the previously filled sublevels) are omitted:

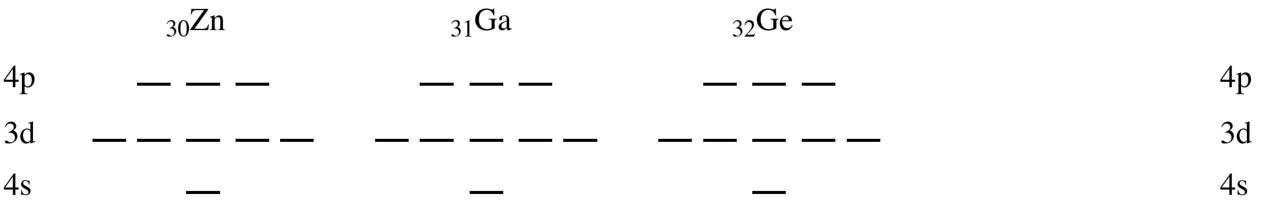


Note that Cr is unusual, with configuration $4s^1 3d^5$, and ***not*** $4s^2 3d^4$ as you would expect. The theoretical explanation is that half-filled or fully-filled sublevels are more stable than sublevels that are neither. As the energy levels get closer in energy, starting with the 4s, it takes more energy to pair the electrons in the 4s than it does to promote one to the 3d level. Why is this true? It is true because there is stability gained (energy kick-back) in producing a half-filled or fully-filled 3d sublevel. Therefore an electron will promote from the 4s to the 3d sublevel when the resulting configuration has a half-filled or fully-filled 3d sublevel. All other electron configuration inconsistencies found later in the periodic table can be explained with similar arguments, that the pairing energy is greater than the energy required to promote.

Exercise:

Continue filling in the orbital diagrams below for Fe to Ge. (NOTE: Cu is an exception, similar to Cr.) Under each orbital diagram on this page, write the short notation for the electron configuration.





Now, sharpen your pencil, and in the small space provided on the periodic chart on the next page, write the electronic configuration of the highest filled, or partially filled, energy sublevel for each element up to xenon (element 54).

1 1A																	18 8A
1 H	2 2A											13 3A	14 4A	15 5A	16 6A	17 7A	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
	Lanthanides																
	Actinides																

Which groups are the representative (main group) elements?

Which sublevels are filling as one goes across a row of representative elements?

Which sublevels are filling as one goes across a row of transition elements?

Which sublevel is being filed by the lanthanides?

Which sublevel is being filled by the actinides?

Which sublevel is being filled by elements $_{104}$Rf and $_{105}$Db?

Because the inner (noble gas core) electrons of an element are not as important as the outer most electrons, chemists often write electron configurations in an abbreviated form. The noble gas core electrons are represented with the noble gas symbol written in brackets followed by the remaining configuration. *Only the noble gases may be used for core representations.* Write the shorthand electronic configurations for the following:

Si: [Ne]$3s^2 3p^2$

Ag:[Kr]

Se:[Ar]

Br:[Ar]

Sr:[Kr]

Ar: [Ne]

Diamagnetic elements have no unpaired electrons in any orbital; paramagnetic ones have one or more unpaired electrons. Identify which of the following are diamagnetic, and which are paramagnetic (remember that all are one or the other!). Write "dia" or "para":

He: B: C: Ne: Ti: Fe: Zn: S: Sn:

Lewis Structures

To draw a Lewis structure for a molecule or an ion:

1. Write the chemical symbols for the atoms in their bonding sequence as given in the problem.
2. Draw single bonds to connect the atoms.
3. Complete octets around all atoms except H, which gets a "duet".
4. Count available valence electrons.
    a. all outer-shell electrons
    b. each "-" ionic charge represents one more electron
    c. each "+" ionic charge represents one less electron
5. If octets require more electrons than are available, remove electron pairs as needed and move other pairs to form multiple bonds (original single bond plus 1 or 2 more bonds).
    a. H, F, Cl, Br, I form 1 bond
    b. O, S, Se " 2 " bonds
    c. N, P, As " 3 "
    d. C forms " 4 "
6. If octets require more electrons than are available, but none of the atoms involved can form multiple bonds, leave the central atom with a sextet or less.
7. If octets require fewer electrons than are available, add extra electron pairs to the (large) central atom.
8. If the chemical species is ionic, put brackets around the structure and write the charge as a superscript outside the brackets.

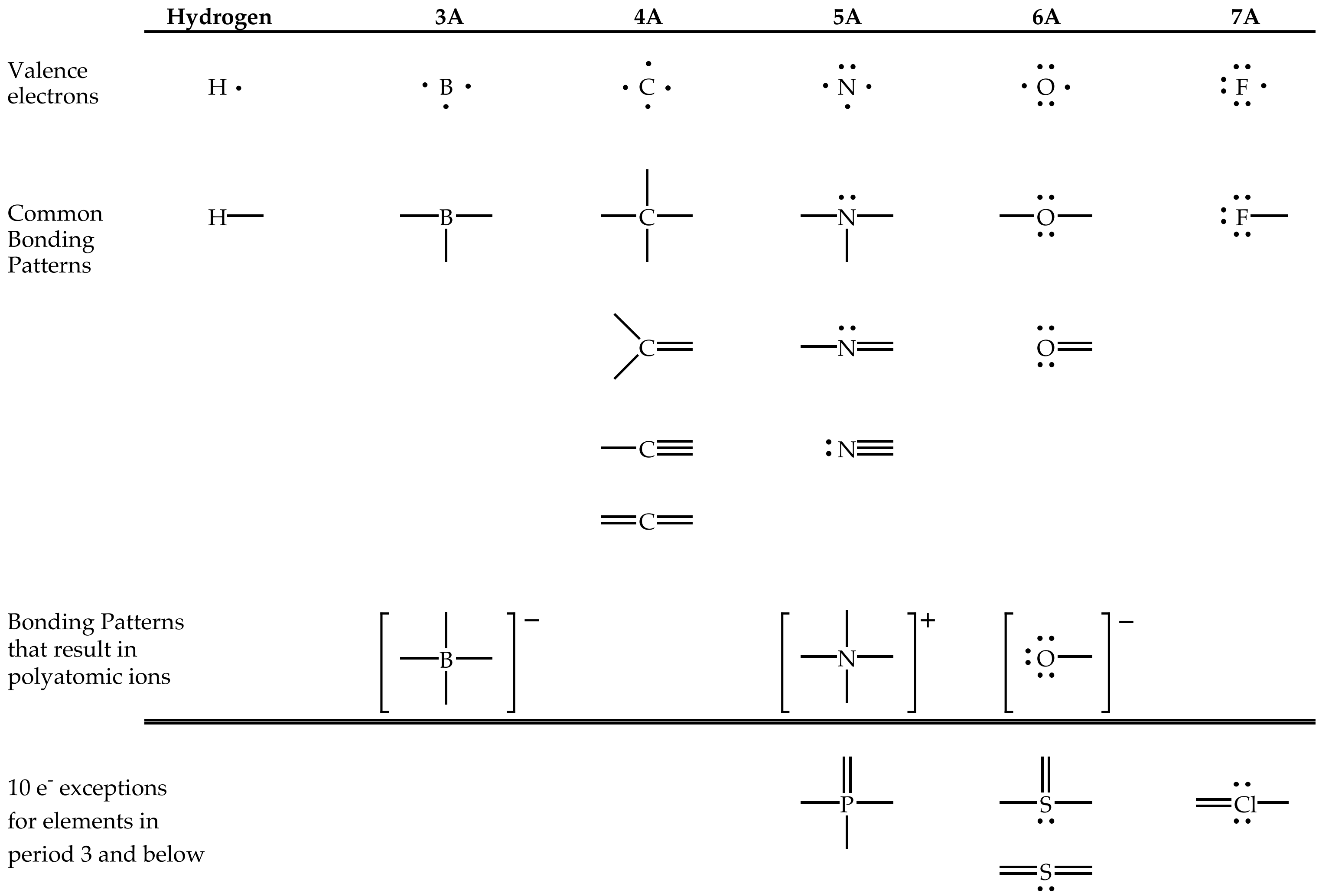
VSEPR

To predict molecular or ionic shapes using valence shell electron pair repulsion (VSEPR) theory:

1. Draw the Lewis structure (see above).
2. Identify the central atom or atoms (you may be asked to do a VSEPR analysis for more than one atom in a structure).
3. For each central atom, note the number of atoms bonded to it and the number of non-bonding electron pairs on it. Remember that the number of atoms bonded to the central atom and the number of bonds on it are not necessarily the same because of the possibility of multiple bonds.
4. Symbolize the central atom as "A", each bonded atom as "X" and each non-bonding electron pair as "E", and write a formula in the form AX_nE_m.
5. Identify the shape around the central atom from the following table, which you must memorize:

Formula	Shape
AX_2	linear
AX_3	trigonal planar
AX_2E	bent (~120°)
AX_4	tetrahedral
AX_3E	trigonal pyramidal
AX_2E_2	bent (~109°)

Lewis Structures - Common Bonding Patterns For Representative Elements



Name____________________________

Section____________ **Date**_______

Lewis Dot Structure Worksheet

Write the Lewis electron dot structures for the following molecules and ions. The bonding sequence is given where needed.

1. $\underline{N}F_3$

2. $\underline{S}Cl_2$

3. $\underline{S}OCl_2$ (central S)

4. $\underline{C}S_2$

5. $\underline{C}Cl_2F_2$ (central C)

6. C_2H_6O (C-C-O sequence)

7. C_2H_6O (C-O-C sequence)

8. $\underline{N}H_2Cl$ (central nitrogen)

9. $H\underline{Cl}O_2$ (central chlorine, hydrogen attached to oxygen)

10. ClO^-

11. $\underline{P}H_4^+$

12. CH_3-CH_2Cl (C-C-Cl sequence)

13. CN-CO-CN (C-C-C sequence)

14. CCl_3-CHO (C-C sequence)

Use the VSEPR theory to predict the shapes of the following. 1. Draw Lewis dot structures for each case. 2. Describe the shape about each underlined central atom in terms of an expression of the form AX_nE_m and in words (e.g., “linear”). Some structures have more than one indicated central atom. 3. Sketch the shape of each molecule or ion.

15. $\underline{C}OF_2$

16. $\underline{C}F_4$

17. $\underline{P}Br_3$

18. $\underline{C}OS$

19. $\underline{S}O_2$

20. $\underline{Se}Br_2$

21. $\underline{C}H_3\text{-}\underline{C}O\text{-}CH_3$ (C-C-C sequence; give shape for two centers)

22. $\underline{C}H_3\text{-}\underline{S}H$ (give shape for two centers)

23. $Cl\underline{C}H_2\text{-}\underline{C}N$ (C-C sequence; give shape for two centers)

24. $\underline{As}H_4^+$

25. $\underline{Cl}O_2^-$

26. $O\text{-}\underline{P}\text{-}\underline{O}\text{-}P\text{-}O$ (i.e., P_2O_3; give shape for two centers)

Name______________________________

Section____________ **Date**_______

Inorganic Nomenclature

Every compound has its own **CHEMICAL FORMULA** and its own **NAME**. The nomenclature (naming systems) for **IONIC** and **MOLECULAR** compounds are different.

IONIC COMPOUNDS: These consist of any positive ion (except H^+) with any negative ion. Note: if H^+ is the positive ion, it's an **ACID** – see below.

*The positive ion (**CATION**) may be a metal ion, e.g., Na^+, or a polyatomic ion, e.g., NH_4^+.

*The negative ion (**ANION**) may be a non-metal ion, e.g., Cl^-, or a polyatomic ion, such as SO_4^{2-}

Case 1. Representative Metal + Non-metal Compounds

Examples:	KBr	Potassium bromide
	$AlCl_3$	Aluminum chloride

Note: Metal always is first (name unchanged), non-metal is second in formula (name given an –IDE ending).

Note: The name does NOT indicate how many of each.

EXERCISE: Name the following:

NaF __________________

BaS __________________

SrI_2 __________________

K_2O __________________

Al_2O_3 _________________

GaN__________________

Give formulas for the following (refer to periodic table only):

Cesium phosphide____________________

Barium fluoride______________________

Lithium oxide _______________________

Strontium bromide ___________________

Barium ion _________________________

Chloride ion ________________________

Calcium iodide ______________________

Magnesium nitride ___________________

Rubidium sulfide_____________________

Sodium selenide _____________________

Nitride ion__________________________

Aluminum ion ______________________

Case 2. Transition Metal + Non-metal Compounds

In general, the ions formed by the transition metals are not predictable. **MEMORIZE** those given on the last page of handout (Index cards are very helpful).

*If the transition metal forms only 1 ion, name the compound as in Case 1.

Examples:	$ZnCl_2$	Zinc chloride
	Ag_2S	Silver sulfide

*If the transition metal can form more than one type of ion, name the compounds with on or both of two systems (each compound has 2 names).

		Name 1		**Name 2**
Examples:	FeO [formed from Fe^{2+} and O^{2-}]	Iron (II) oxide	**or**	ferrous oxide
	Fe_2O_3 [formed from Fe^{3+} and O^{2-}]	Iron (III) oxide	**or**	ferric oxide

The **–ous** ending refers to the ion with the **lower** charge state, e.g., Fe^{2+}, Cu^{+} (cuprous).
The **–ic** ending refers to the ion with the higher charge state, e.g., Fe^{3+}, Cu^{2+} (cupric).

Note: the modern names for cuprous and cupric ions would be copper (I) and copper (II), where the roman numeral in parentheses corresponds to the number of charges.

Note: For each compound you must learn both the modern (includes roman numeral) name and the older name.

Note: Lead and tin form 2+ and 4+ ions even though they are not transition metals.

EXERCISE: After you have learned the ions on the back page, name the following using both systems:

Pb^{2+} _____________	Sn^{4+} _____________	Fe^{2+} _____________	Cu^{2+} ____________
_____________	_____________	_____________	____________

Name the following compounds (two names if appropriate):

$FeBr_3$ ___________________

Cu_3N __________________

AgCl ___________________

From the three compounds above, what is the charge on Fe? on Cu?

Give formulas for the following compounds:

Ferric oxide ______________

Nickel fluoride___________

Chromic oxide____________

Cuprous sulfide __________

Ferrous iodide ____________ Plumbic chloride _________

Cases 1 and 2 (above) involve ionic compounds which consist only of a metal cation and a non-metal anion – two elements only. These are called **BINARY** compounds and consist of two **MONOATOMIC** ions. Ionic compounds can also be formed from more complex, polyatomic ions.

Case 3. Ionic Compounds with POLYATOMIC Ions

The list of polyatomic ions (names and formulas) which must be learned are on the back page of the handout.

For compounds, the cation is named first, then the anion:

Examples:	$(NH_4)_2S$	Ammonium sulfide		
	K_3PO_4	Potassium phosphate		
	$Fe_2(SO_4)_3$	Iron (III) sulfate	**or**	ferric sulfate

Question: How do you know it's Fe^{3+} and not Fe^{2+} in ferric sulfate????

Note: Parentheses ( ) are used only when there are **2 or more** polyatomic ions in the formula.

Exercise: Name the following:

$(NH_4)_2O$ ______________________________ FeC_2O_4 ______________________________

Na_2SO_3 ______________________________ $CuNO_3$ ______________________________

LiSCN ______________________________ $NaHCO_3$ ______________________________

Give formulas for the following:

Cupric nitrate ____________ Zinc phosphate ___________ Silver carbonate __________

Chromic hydroxide ________ Mercuric cyanide __________ Lead (IV) acetate _________

Potassium dichromate ______ Nickel hydroxide __________ Barium permanganate _____

Calcium carbonate ________ Potassium chlorate _________ Cadmium sulfate _________

Mercurous iodide _________ Ammonium phosphide _____ Cobaltous nitrite __________

Fill in the Formula and name each compound. For iron and copper compounds, give old and new names.

	CO_3^{2-}	HCO_3^-	CrO_4^{2-}	NO_3^-	NO_2^-	OH^-	SO_4^{2-}
Al^{3+}							
Mg^{2+}							
Cu^+							
Cu^{2+}							

	HSO_4^-	SO_3^{2-}	HSO_3^-	PO_4^{3-}	CrO_4^{2-}	MnO_4^-	$C_2H_3O_2^-$
Fe^{3+}							
Ag^+							
Fe^{2+}							

Case 4. Naming ACIDS

ACIDS (from the Latin word *acidus*, meaning “sour”) are an important class of compounds and are defined as a **substance whose molecules each yield one or more hydrogen ions (H^+) when dissolved in water**. The formula for an acid is formed by adding sufficient H^+ ions to balance the anion’s charge. The name of the acid is related to the name of the anion and includes the label **acid**. Anions whose names end in **–ide** have associated acids that have the **hydro-** prefix and an **–ic** ending.

Example:	ANION	CORRESPONDING ACID
	Cl^-	HCl (**hydro**chlor**ic acid**)

Many of the most important acids are derived form **OXYANIONS** (polyatomic ions which contain oxygen). Oxyanions whose names end in **-ite** (sulfite, nitrite, chlorite, etc.) have associated acids whose names end in **–ous**,

Examples:	SO_3^{2-} (sulfite)	H_2SO_3	sulfur**ous acid**
	ClO_2^- (chlorite)	$HClO_2$	chlor**ous acid**

Oxyanions whose names end in **–ate** (sulfate, phosphate, nitrate, chlorate, etc.) have corresponding acids whose names are given in **–ic** ending,

Examples:	SO_4^{2-} (sulfate)	H_2SO_4	sulfur**ic acid**
	ClO_3^- (chlorate)	$HClO_3$	chlor**ic acid**

Note: Chlorite and chlorate ions are not included in the attached table but are legitimate oxyanions of chlorine and quite common.

Exercise: a) Give the appropriate formula for the corresponding acids of the following anions and b) name the acid:

Sulfide a) _________ b) _____________________________

Carbonate a) _________ b) _____________________________

Oxalate a) _________ b) _____________________________

Phosphate a) _________ b) _____________________________

Acetate a) _________ b) _____________________________

Nitrite a) _________ b) _____________________________

Case 5. Naming BINARY MOLECULAR COMPOUNDS

MOLECULAR COMPOUNDS are the compounds formed when non-metal atoms **SHARE** electrons with other non-metal atoms. **BINARY** molecular compounds consist of two different atoms and should be named in the following way:

Similar to ionic compounds, the more positive ion is first and the more negative ion is second with the negative ion's name including an **–ide** ending. To determine which is the most positive or negative compare relative electronegativities.

Unlike ionic compounds, the number of each atom is specified with a prefix:

1 – mono
2 – di
3 – tri
4 – tetra
5 – penta

6 – hexa
7 – hepta
8 – octa
9 – nona
10 – deca

If there is only one of the first atom, the mono prefix is not used.

Examples:	NO	nitrogen monoxide	N_2O	dinitrogen monoxide
	NO_2	nitrogen dioxode	IF_7	iodine heptafluoride

Exercise: Name the following:

N_2O_3 ________________

N_2O_4 ________________

CO_2 ________________

P_2O_5 ________________

N_2O_5 ________________

CO ________________

Cl_2O ________________

N_2 ________________

Give the formulas for the following:

Iodine pentabromide ________________

Oxygen difluoride ________________

Sulfur hexafluoride ________________

Hydrogen ________________

Iodine heptafluoride ________________

Sulfur trioxide ________________

Chlorine trifluoride ________________

Carbon tetrafluoride ________________

Silicon dioxide ________________

Ammonia ________________

Dibromine monoxide ________________

Tetraphosphorus decaoxide ________________

Common Ions

Positive Ions (CATIONS)

+1 Charge

Ammonium (NH_4^+)*
Copper (I) or cuprous (Cu^+)*
Hydrogen (H^+) “proton”*
Hydronium (H_3O^+) “aqueous proton”*
Silver (Ag^+)*

+2 Charge

Cadmium (Cd^{2+})
Cobalt (II) or cobaltous (Co^{2+})
Copper (II) or cupric (Cu^{2+})*
Iron (II) or ferrous (Fe^{2+})*
Lead (II) or plumbous (Pb^{2+})*
Manganese (II) or manganous (Mn^{2+})*
Mercury (I) or mercurous (Hg_2^{2+})*
Mercury (II) or mercuric (Hg^{2+})*
Nickel (Ni^{2+})*
Tin (II) or stannous (Sn^{2+})*
Zinc (Zn^{2+})*

+3 Charge

Aluminum (Al^{3+})*
Chromium (III) or chromic (Cr^{3+})*
Iron (III) or ferric (Fe^{3+})*

+4 Charge

Lead (IV) or plumbic (Pb^{4+})*
Tin (IV) or stannic (Sn^{4+})*

Negative Ions (ANIONS)

-1 Charge

Acetate ($C_2H_3O_2^-$)*
Perchlorate (ClO_4^-)
Chlorate (ClO_3^-)
Chlorite (ClO_2^-)
Hypochlorite (ClO^-)*
Cyanide (CN^-)
Dihydrogen phosphate ($H_2PO_4^-$)*
Hydrogen carbonate or bicarbonate (HCO_3^-)*
Hydrogen sulfate or bisulfate (HSO_4^-)*
Hydrogen sulfite or bisulfite (HSO_3^-)*
Hydroxide (OH^-)*
Nitrate (NO_3^-)*
Nitrite (NO_2^-)*
Permanganate (MnO_4^-)*
Thiocyanate (SCN^-)

-2 Charge

Carbonate (CO_3^{2-})*
Chromate (CrO_4^{2-})
Dichromate ($Cr_2O_7^{2-}$)
Hydrogen phosphate (HPO_4^{2-})*
Oxalate ($C_2O_4^{2-}$)
Sulfate (SO_4^{2-})*
Sulfite (SO_3^{2-})*
Peroxide (O_2^{2-})*

-3 Charge

Phosphate (PO_4^{-3})*

Chemical Names and Formulas

Goals

- ☐ Write chemical names and formulas of common chemical compounds.
- ☐ Describe the colors and textures of common ionic compounds.
- ☐ Synthesize chemical compounds and write their names and formulas.

Introduction

Chemistry is the central science, a study of all that has mass and volume. An effort of this magnitude requires a clear language that communicates in a broad but consistent way. At first appearance, chemistry may appear difficult because there are common words that take on new meaning in chemistry. For example, "salt" is a term widely used to describe table salt also known as sodium chloride. In chemistry, a **salt** is simply **any compound composed of ions other than hydrogen ion, oxide ion, or hydroxide ion**. Sodium chloride is an example of a salt, as is potassium chloride, calcium carbonate and stannous fluoride. In chemistry, there is an effort to move away from using common names to identify the majority of compounds because this would require memorization of every single name. Considering the vast number of ionic compounds (over a million), a systematic method has been developed.

As a student of chemistry you will learn how to translate a chemical formula into the systematic name and vice versa. The observations and experiment in today's lab only involve compounds containing charged species – cations, anions and polyatomic ions. These combinations of oppositely charged ions are called **ionic compounds.** Their chemical formulas represent the proportion of positive ion to negative ion that results in electrical neutrality, i.e., no net charge. The correct chemical formula for sodium chloride is $NaCl$. The 1:1 ratio of sodium to chloride ions tells us that sodium ions and chloride ions must have the same charge (Note that when there is only one of an ion per formula, we do not use the number one as subscript to indicate this; i.e., we don't write Na_1Cl_1). After looking on the table on the following page, we see that sodium is a cation with a 1+ charge and chloride is an anion with a 1- charge. Knowing both the magnitude and the sign of the charge is necessary for writing the correct formulas and the correct chemical names. Sodium oxide has a formula of Na_2O. Without looking at the table of ions, what must the charge of oxide be? If you recognized that there are two sodium 1+ ions for every one oxide and deduced that oxide must have a 2- charge you are well on your way to describing ionic compounds!

In the chemical reactions of ionic compounds in solution, it is common to see **precipitates**, insoluble solids coming out of solution. A general rule is that precipitates usually do not contain sodium, potassium, acetate, or nitrate ions. This could be important!!

Table I. A Collection of Common Ions.

Name	Formula	Name	Formula	Name	Formula
sodium	Na^+	magnesium	Mg^{2+}		
potassium	K^+	calcium	Ca^{2+}		
copper (I)	Cu^+	copper (II)	Cu^{2+}		
silver	Ag^+	iron (II)	Fe^{2+}	iron (III)	Fe^{3+}
ammonium	NH_4^+	lead (II)	Pb^{2+}	lead (IV)	Pb^{4+}
		tin (II)	Sn^{2+}	tin (IV)	Sn^{4+}
fluoride	F^-	oxide	O^{2-}	nitride	N^{3-}
chloride	Cl^-	sulfide	S^{2-}	phosphate	PO_4^{3-}
bromide	Br^-	sulfate	SO_4^{2-}		
iodide	I^-	carbonate	CO_3^{2-}		
acetate	$C_2H_3O_2^-$	hydrogen phosphate	HPO_4^{2-}		
hydroxide	OH^-				
nitrate	NO_3^-				
nitrite	NO_2^-				
hydrogen carbonate (**bicarbonate**)	HCO_3^-				
dihydrogen phosphate	$H_2PO_4^-$				

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

*Contact with silver nitrate ($AgNO_3$) will stain the skin.

Materials:

Reagent Central chemicals include a variety of pure ionic compounds and aqueous solutions of ionic compounds as identified on your experimental pages.

Equipment:	Empty pipet for stirring	Lab top reaction surface

Experimental Procedure

1. View the samples of solid compounds available at Reagent Central. Record the name and formula of each substance. Write a description of the color and any other adjectives that might distinguish one compound from another. Record observations and answers in your laboratory notebook.

2. a. Record initial observations on each of the aqueous solutions (Reactants).

3. a. Insert your experimental page inside of your reaction surface.

   b. Place one drop of each solution in the indicated spaces below, taking care not to contaminate the microburets. Stir by blowing air from a dry long stem pipet. Record any observable changes, describing what happened when the two solutions were mixed.

4. Answer all the questions.

Reaction Guide: Insert this page into the labtop. Mix one drop of each, using a long stem pipet to blow air past the droplet to complete the mixing.

	$AgNO_3$	$Pb(NO_3)_2$			
$FeCl_3$	✗	✗			
KI	✗	✗	$CuSO_4$	$MgSO_4$	$FeCl_3$
NaOH	✗	✗	✗	✗	✗
Na_2CO_3	✗	✗	✗	✗	✗
Na_3PO_4	✗	✗	✗	✗	✗

Data Organization Suggestions

Part 1. Observation of common ionic compounds.

Construct a table with columns for "Compound Name," "Formula," and "Description." Leave space between each data entry row to make the information easier to read. Descriptions should be sufficient to help you identify the material within a series of unlabeled samples. Some examples follow:

Compound Name	Formula	Description
Silver Nitrate	$AgNO_3$	Brown semi-transparent flat crystals < $1mm^2$
Sodium Carbonate	Na_2CO_3	Small cylindrical rod like nodules. Not crystaline.
Calcium Chloride	$CaCl_2$	Opaque white spheres of varring sizes. Appears hygroscopic and bound in mass at bottom of tube.

Part 2. Initial observation of aqueous solutions.

Construct a table similar to that in part 1. Your descriptions should be appropriate for liquids and solutions. Review the information given on page 17 of this text regarding proper descriptive terms for liquids.

Part 3. Observation of precipitation reactions.

Construct a table with column headings for "Reaction," and "Description." Once again descriptions should be sufficient to help you identify the reaction result from a series of unknown samples. Review the information given on page 17 of this text regarding proper descriptive terms for precipitates. Some examples follow:

Reaction	Description
$KI_{(aq)}$ + $Pb(NO_3)_{(aq)}$	Fast reaction yeilding a yellow opaque suspension. Upon stirring the particles clump slightly
$NaOH_{(aq)}$ + $FeCl_{3(aq)}$	Fast reaction forming a cloudy red-brown percipitate. Upon stirring the precipitate finely divides and remains in suspension.
$Na_2CO_{3(aq)}$ + $MgSO_{4(aq)}$	Slower reaction forming small amount of suspended white precipitate. Apperance unaffected by stirring.

Data Analysis

Answer the following questions in your laboratory notebook using complete sentences.

1. Write the formulas (with charges) and names of all the cations represented in this experiment.
2. Write the formulas (with charges) and names of all the anions represented in this experiment.
3. Write some simple rules for naming ionic compounds.
4. Write some simple rules for writing chemical formulas of ionic compounds.
5. When are Roman numerals used in naming compounds?
6. What does a numerical subscript following an element in a chemical formula mean?
7. When is it correct to use parentheses in chemical formulas?
8. Any precipitates represent new compounds formed from swapping ion partners. Write the correct formulas for the two possible products. Indicate which of the two products is the precipitate by the addition of a subscripted "s" in parenthesis (ex. $PbCl_{2\,(s)}$). The precipitate will be the product that doesn't contain sodium, potassium, or nitrate ions.

Predicting the Products of Precipitation Reactions: Solubility Rules

Goals

- ☐ Observe and record precipitation reactions.
- ☐ Derive general solubility rules from the experimental data.
- ☐ Describe precipitation reactions by writing net ionic equations.
- ☐ Understand the relationship between solubility and precipitation reactions.

Introduction

Ionic compounds in solution consist of free ions surrounded by water. In the case of zinc chloride being dissolved in water to form an aqueous solution, the dissolution of the ionic compound can be depicted as follows:

$$ZnCl_{2(s)} \longrightarrow Zn^{2+}_{(aq)} + 2\ Cl^{-}_{(aq)}$$ What type of change is this?

These free aqueous ions can react with other ions. Mixing of dissolved ionic compounds can lead to precipitation reactions, a double displacement reaction. The **precipitate** is a **solid product**, a new ionic compound that is different from the reactants in both composition and solubility. Solubility is defined as the amount of substance (solute) that dissolves in a given amount of solvent. Solubility is a physical property that can be useful in predicting whether the mixing of aqueous ionic compounds will lead to a precipitation reaction. The mixing of a variety of combinations leads to the formulation of general rules of solubility. Some examples of these rules include "All sodium salts are soluble in water" or "The mixing of two ionic compounds that contain a common ion will not lead to a precipitate".

Let's look at an example to see how these solubility rules can help us. As part of the lab, aqueous solutions of silver nitrate and sodium carbonate are mixed. A foggy white precipitate is formed. To write the chemical equation:

1. *Identify the reactants and write their correct formulas*:

   $$AgNO_{3(aq)} + Na_2CO_{3(aq)} \longrightarrow$$

   silver nitrate   sodium carbonate

2. *"Swap" the cations of the reactants to form the products, two new ionic compounds:* the products will be sodium nitrate and silver carbonate.

3. *Write the* ***correct*** *formulas for the products after the arrow. Use the names of the products as an aid to get the formulas correct. Each product must have a net charge of zero. In other words the sum of the positive charges must equal the sum of the negative charges within each product.*

$$AgNO_{3(aq)} \quad + \quad Na_2CO_{3(aq)} \longrightarrow Ag_2CO_3 \quad + \quad NaNO_3$$

silver nitrate   sodium carbonate   silver carbonate   sodium nitrate

4. *Use solubility rules to determine which product is the precipitate:* "All sodium salts are soluble"; therefore silver carbonate must be the foggy, white precipitate.

$$AgNO_{3(aq)} \quad + \quad Na_2CO_{3(aq)} \longrightarrow Ag_2CO_{3(s)} \quad + \quad NaNO_{3(aq)}$$

5. ***LAST*** *but not least,* ***balance the equation*** *using whole number coefficients*:

$$2\ AgNO_{3(aq)} \quad + \quad Na_2CO_{3(aq)} \longrightarrow Ag_2CO_{3(s)} \quad + \quad 2\ NaNO_{3(aq)}$$

How do you convert a chemical equation into a correct net-ionic equation?

1. *Rewrite the correct chemical equation* ***but*** *this time write any aqueous (not solid) ionic compound as free cations, with proper charge and aqueous symbol, and anions, with proper charge and aqueous symbol. This form is referred to as the total ionic equation.* Remember that coefficients in front of a species apply to both cation and anion. Only subscripts on monatomic ions and subscripts outside of parenthesis surrounding polyatomic ions multiply against coefficients. Subscripts that are part of polyatomic ions **do not** multiply against coefficients. Look below, which subscripts changed coefficients when the aqueous compounds were written as free ions?

$$2\ Ag^{+}_{(aq)} + 2\ NO_3{}^{-}{}_{(aq)} + 2\ Na^{+}_{(aq)} + CO_3{}^{2-}{}_{(aq)} \longrightarrow 2\ Na^{+}_{(aq)} + 2\ NO_3{}^{-}{}_{(aq)} + Ag_2CO_{3(s)}$$

2. Once you are sure you have the correct total ionic equation, look at the equation again. It should be balanced and you will be able to see what ions actually changed. Any species that is exactly the same on both sides is considered to be a spectator ion. *To write the net ionic equation, eliminate the spectators and only write the species that changed*:

$$2\ Ag^{+}_{(aq)} \quad + \quad CO_3{}^{2-}{}_{(aq)} \longrightarrow Ag_2CO_{3(s)}$$

Why so many types of equations? Each equation is useful for different reasons. The net ionic is preferred in double displacement reactions because it focuses solely on the product.

In this lab you will mix a variety of solutions. By using sodium salts for at least one of the reactants you will be able to identify which product is the precipitate. To assist you in identifying the solid product, you will write chemical equations for all observable reactions. These chemical reactions will be translated into net ionic equations. After examining the chemical formulas of the precipitates, you will "conclude" by summarizing your results as a set of solubility rules.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

*Contact with silver nitrate ($AgNO_3$) will stain the skin.

Use microburets to dispense reagents in such a way that they do not make contact with other drops or the reaction surface.
Return any contaminated microburets to your instructor.

Materials:

Reagent Central chemicals include aqueous solutions of the compounds listed on your experimental page.

Equipment:	Clean, dry transfer pipet for mixing.	Labtop reaction surface

Experimental Procedure

1. Insert your experimental page under your reaction surface. Place 1 drop of each solution in the squares on your experimental page. Record what happened after mixing with the air from a clean, dry pipet. ("NVR" can indicate no visible reaction.) Please include adjectives that describe both color and texture being as specific as possible (not all white precipitates look the same).

2. After all the reactions are completed and all observations recorded, take one last look at your surface. Have any of the squares changed over time? Record any noticeable changes (there won't be many). Clean your surface by absorbing the contents onto a paper towel. Rinse the reaction surface with a damp paper towel and dry it. Dispose of paper towels in waste bin. Clean your area. Wash your hands thoroughly with soap and water.

3. Answer the questions, correctly writing both chemical and net ionic equations.

4. Draw general conclusions about the cations and anions in your experiment by formulating your own solubility rules.

5. Apply your rules to unknown combinations.

Reaction Guide: Insert this page into the labtop. Mix one drop of each, using a long stem pipet to blow air past the droplet to complete the mixing.

Sol'ns	Na_2CO_3	NaCl	NaOH	$NaNO_3$	Na_3PO_4	Na_2SO_4
$AlCl_3$	✖	✖	✖	✖	✖	✖
NH_4Cl	✖	✖	✖	✖	✖	✖
$CaCl_2$	✖	✖	✖	✖	✖	✖
$CuSO_4$	✖	✖	✖	✖	✖	✖
$FeCl_3$	✖	✖	✖	✖	✖	✖
$Pb(NO_3)_2$	✖	✖	✖	✖	✖	✖
KI	✖	✖	✖	✖	✖	✖
$AgNO_3$	✖	✖	✖	✖	✖	✖

Data Organization Suggestions

There are a number of important facets to this experiment. Learning how to predict the products of precipitation reactions, learning to write proper balanced chemical equations and net ionic equations, and observing solubility trends. To make all of these goals easier use the reaction guide on the previous page as a template for a data table. Use an entire page of your notebook for the table. Use the ruler provided in your chemistry kit to make clean lines. When a precipitation reaction is observed, indicate the color and use one or two adjectives to describe the product (milky, cloudy, marbled, granular….). Where no visible reaction is observed you can simply write "NVR" for "no visible reaction (remember to define abbreviations somewhere on the page).

Data Analysis

Answer the following questions in your laboratory notebook using complete sentences.

1. Look across the rows, what cations are always soluble (never form precipitates)?
2. Look down columns, what anions are always soluble (never form precipitates)?
3. Look down columns, what anions are usually soluble (seldom form precipitates)? Which cation(s) will form precipitates with these anions (your looking for exceptions to the general trend)?
4. Look down columns, what anions are usually insoluble (usually form precipitates)? Which cation(s) will not form precipitates with these anions (your looking for exceptions to the general trend)?
5. Given that you have answered questions two through five correctly, would you expect to see a precipitate form in the reaction of aqueous magnesium chloride with aqueous potassium phosphate? Write out the complete chemical equation for this reaction. Defend your answer by stating your reasoning.
6. Would you expect to see a precipitate form in the reaction of aqueous sodium chloride with aqueous ammonium hydroxide? Write out the complete chemical equation for this reaction. Defend your answer by stating your reasoning.
7. For all reactions in which there was an observable change (precipitate formation) write out a complete correct chemical equation to describe the reaction, followed by the net ionic equation. Refer back to the introduction if you need assistance. There should be 22 reactions in all (go back and look at the $CaCl_{2(aq)}$ with $Na_2SO_{4(aq)}$). Allow enough space between equations in your notebook so that your work is easily readable.

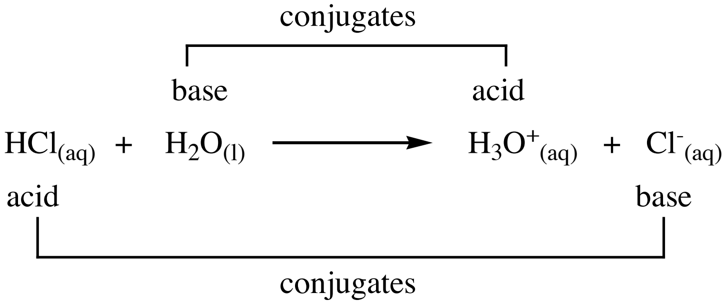
Strength of Acids and Bases: An Equilibrium Issue

Goals

- Determine the strength of acids and bases.
- Use the Brønsted-Lowry definitions of an acid and a base to identify acids and bases.
- Use the Brønsted-Lowry definition of an acid and base to identify conjugate acid-base pairs.
- Describe acid-base reactions by using hydrogen-ion (proton) transfer reactions.
- Explain the differences between strong and weak acids and bases by using equilibrium principles.

Introduction

The strength of an acid or base reflects how readily it ionizes in water. Hydrochloric acid, readily donates a proton to water to become hydronium ion and chloride ion.



$$HCl_{(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

Does this reaction go to completion? Are all the molecules converted into ions? In today's lab, you will use two indicators, bromthymol blue (BTB) and universal (UI), to probe acid-base strength.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

Materials:

Reagent Central solutions include:

Name	Formula	Name	Formula
hydrochloric acid	HCl	sodium hydroxide	$NaOH$
nitric acid	HNO_3	potassium hydroxide	KOH
sulfuric acid	H_2SO_4	calcium hydroxide	$Ca(OH)_2$
acetic acid	$HC_2H_3O_2$	ammonia	NH_3
sodium carbonate	Na_2CO_3	ammonium hydroxide	NH_4OH
sodium dihydrogen phosphate	NaH_2PO_4	sodium hydrogen carbonate	$NaHCO_3$
sodium hydrogen sulfate	$NaHSO_4$	sodium hydrogen phosphate	Na_2HPO_4
boric acid	H_3BO_3	sodium phosphate	Na_3PO_4
citric acid	$H_3C_6H_5O_7$	sodium hydrogen sulfite	$NaHSO_3$
sodium acetate	$NaC_2H_3O_2$		

Bromthymol blue (BTB) in its green form, Universal Indicator (UI)
pH Buffer solutions (pH 1 through pH 12)
Equipment: 1×12 well strip Empty pipet for stirring Lab top reaction surface

Experimental Procedure

1. Use the 1×12 well strip and the pH buffer solutions to make a pH meter. Notice on the top of the 1×12 well strip there are small numbers corresponding to 1 – 12. Start by placing 1 drop of Universal Indicator (UI) into each well. Next place 5 drops of pH 1 buffer solution into well one, followed by 5 drops of pH 2 buffer solution into well two and so on. Compare your pH meter to your neighbor's as a check.
2. Part A and B. Mix 1 drop of each of the indicated solutions. Look carefully for similarities and differences to distinguish strong acids from weak acids and strong bases from weak bases. Record your observations. Be specific in your descriptions. In the last row test each acid/base for conductivity. Be sure to rinse the leads of the conductivity apparatus between tests. Record your observations. Use your pH meter to compare with the Universal Indicator (UI) test to determine the pH of each solution. Record the pH values.
3. Part C. Mix 5 drops of each of the solutions with one drop of indicator. Record your observations. Determine which of the species is a base and which is an acid. Use your pH meter to compare with the Universal Indicator (UI) test to determine the pH of each solution. Record the pH values.

Reaction Guide: Insert this page into the labtop.

Part A. Acids and Indicators

1 drop ea	HCl	HNO_3	H_2SO_4	2 drops HCl + 9 drops H_2O	$HC_2H_3O_2$	Citric acid ($H_3C_6H_5O_7$)	Boric acid (H_3BO_3)
BTB							
UI							
Na_2CO_3							
Test for conductivity							

Part B. Bases and Indicators

1 drop ea	NaOH	KOH	$Ca(OH)_2$	2 drops NaOH + 9 drops H_2O	Na_3PO_4 (PO_4^{3-})	Na_2CO_3 (CO_3^{2-})	$NaC_2H_3O_2$ ($C_2H_3O_2^-$)
BTB							
UI							
Test for conductivity							

Part C. Acids, Bases, and Indicators

5 drops solution/1 drop indicator	$NaHSO_4$ (HSO_4^-)	Na_2HPO_4 (HPO_4^{2-})	$NaHCO_3$ (HCO_3^-)	NaH_2PO_4 ($H_2PO_4^-$)	$NaHSO_3$ (HSO_3^-)	NH_3
BTB						
UI						

Data Organization Suggestions

The actual amount of written data required for this lab is fairly small. Simple descriptions of color for tests with indicators, relative amounts of bubbles produced with sodium carbonate, brightness of the diode in the conductivity tests, and pH values are all that is required. This data will fit nicely in tables identical to those used in the reaction guide (provided that you add a row for the pH value). As always it is important not to crowd your data. Don't be afraid to turn your notebook page so that the data table runs along the long axis of the page (landscape format).

Data Analysis

Answer the following questions in your laboratory notebook using complete sentences.

Use your observations from Part A to answer the following questions:

1. What conclusions can you draw based upon the BTB test of the different acids?
2. What conclusions can you draw based upon the UI test of the different acids?
3. Describe the results (what you saw) in the sodium carbonate reaction with acids. Make a general conclusion about the reaction of sodium carbonate with acids.
4. Based on the results described in the above question, draw a conclusion about the acids involved. Do some seem "stronger"? Be specific.
5. Two columns include HCl. Which HCl column most closely resembles the acetic acid column? What does this tell you about acetic acid?
6. Why is conductivity used as a probe for acids?
7. Describe your results from the conductivity row in Part A and draw a conclusion.

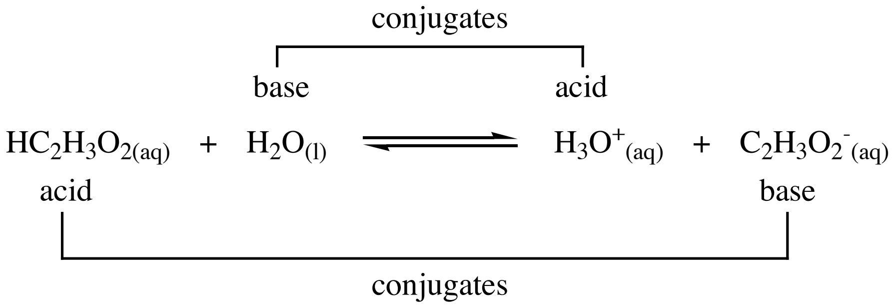
Use your observations from Part B to answer the following questions:

8. What conclusions can you draw based upon the BTB test of the different bases?

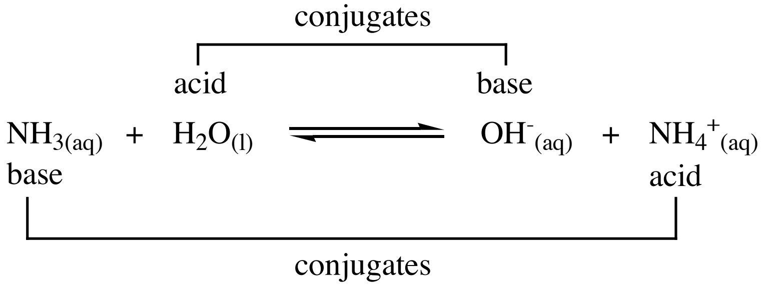
9. What conclusions can you draw based upon the UI test of the different bases?

10. The conductivity test for bases gives conflicting information. The last three compounds, Na_3PO_4, Na_2CO_3, $NaC_2H_3O_2$ are weak bases, however the conductivity tests indicate a strongly electrolytic solution. Can you explain why these weak bases would be strong electrolytes (think solubility!)?

11. For the weak acids in section A, the weak bases in section B, and all the materials in section C, write chemical equations that describe how these materials react with water. Examples of acids and bases reacting with water are given below. The sodium cations (Na^+) in each of these reactions is a spectator ion. For simplicity we are going to focus on the net ionic equations for the weak bases in section B and all materials in section C. For these materials use the formulas of the polyatomic ions given in parenthesis to react with water. Treat each of these as single proton transfers. For those species in section C, use the BTB result to decide how the species should react with water (Yellow = Acid, Blue = Base). As shown below, label each reactant and product as either an acid or base. Additionally, draw lines to link the conjugate pairs (species which differ in their chemical formula by one proton). There are 12 reactions in all.



$$HC_2H_3O_{2(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$



$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + NH_4^+{}_{(aq)}$$

Oxidation-Reduction Reactions

Goals

- ☐ To observe and record oxidation-reduction ("redox") reactions between aqueous ions/
- ☐ Describe oxidation and reduction processes by writing half-reactions.
- ☐ Write and balance redox equations by balancing half-reactions.
- ☐ Identify reactants as oxidizing agents and reducing agents.

Introduction

Oxidation-reduction reactions commonly occur when two metals react with either acids or metal ions. Redox reactions also occur between ions, with one ion being oxidized and the other being reduced. For example, permanganate ions, MnO_4^-, undergo a redox reaction with iodide ions, I^-, in the presence of acid:

$$16\,H^+ + 2\,MnO_4^- + 10\,I^- \longrightarrow 2\,Mn^{2+} + 5\,I_2 + 8\,H_2O$$

This reaction is easily seen because the permanganate has a distinctive purple color that disappears upon the formation of the nearly colorless Mn^{2+} ion. Similarly, the colorless I^- ion reacts to form the I_2 molecule, which appears yellow in solution. The reaction mixture thus turns from purple to yellow as the reaction proceeds.

In this reaction, permanganate is reduced and iodide is oxidized. In other words, manganese gains electrons and iodine loses electrons. Because MnO_4^- accepts electrons from I_- and causes I^- to be oxidized, MnO_4^- is also called the oxidizing agent. Similarly, because I_- donated electrons to MnO_4^- and causes MnO_4^- to be reduced, I^- (the oxidized reactant) is also called a reducing agent.

In this lab, you will carry out a number of redox reactions involving ions in aqueous solution and will write balanced equations to describe the reactions. You will also identify the oxidizing and reducing agents in each case and become familiar with which substances usually act as oxidizing agents and which are most often reducing agents.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

Materials:

Reagent Central solutions include:

Name	Formula	Name	Formula
potassium iodide	KI	sodium nitrite	$NaNO_2$
sodium hypochlorite	NaOCl	copper II sulfate	$CuSO_4$
potassium permanganate	$KMnO_4$	Iron III chloride	$FeCl_3$
hydrogen peroxide	H_2O_2	sodium hydrogen sulfite	$NaHSO_3$
potassium dichromate	$K_2Cr_2O_7$	sodium thiosulfate	Na_2S_2O

Equipment: Empty pipet for stirring Lab top reaction surface

Experimental Procedure

1. Start by making initial observations on all starting materials.
2. Mix 1 drop of each of the indicated solutions in the spaces below (a – h). If a visible reaction is observed record the observation in your laboratory notebook, otherwise record "NVR" for "no visible reaction." Proceed to step three.
3. Proceed by adding a drop of starch solution, an indicator of I_2, to each of the reaction mixtures. If you observe the formation of a black opaque suspension, a positive test for I_2, record the observation in your laboratory notebook. Go on to the next reaction and start the process over again with step two. **If, and only if**, there is no visible reaction record "NVR" for "no visible reaction" and proceed to step four.
4. If there was no visible reaction after the addition of starch, add a drop of hydrochloric acid (HCl). Record your observation. Go on to the next reaction and start the process over again with step two.
5. Once you have completed steps two through four for all the reactions, go back and add drops of **$NaHSO_3$** to each of the mixtures in **a – d**. Add drops of $NaHSO_3$, keeping track of the number of drops, until the black opaque suspension disappears. Record the number of drops and your observations in your data table.
   The reaction is: $HSO_3^- + I_2 \longrightarrow SO_4^{2-} + 2\,I^-$
6. Add drops of **$Na_2S_2O_3$** to each of the mixtures in **e – h**. Add drops of $Na_2S_2O_3$, keeping track of the number of drops, until the black opaque suspension disappears. Record the number of drops and your observations in your data table.
   The reaction is: $S_2O_3^{2-} + I_2 \longrightarrow S_4O_6^{2-} + 2\,I^-$

Reaction Guide: Insert this page into the labtop.

	KI (I^-)	
a. NaOCl (OCl^-)		$OCl^- + 2\,I^- \longrightarrow Cl^- + I_2$
b. $KMnO_4$ (MnO_4^-)		$MnO_4^- + 2\,I^- \longrightarrow Mn^{2+} + I_2$
c. H_2O_2		$H_2O_2 + 2\,I^- \longrightarrow H_2O + I_2$
d. KIO_3 (IO_3^-)		$IO^- + 2\,I^- \longrightarrow I^- + I_2$

e. $K_2Cr_2O_7$ ($Cr_2O_7^{2-}$)		$Cr_2O_7^{2-} + 2\,I^- \longrightarrow 2\,Cr^{3+} + I_2$
f. $NaNO_2$ (NO_2^-)		$NO^{2-} + 2\,I^- \longrightarrow NO_{(g)} + I_2$
g. $CuSO_4$ (Cu^{2+})		$Cu^{2+} + 2\,I^- \longrightarrow CuI_{(s)} + I_2$
h. $FeCl_3$ (Fe^{3+})		$Fe^{3+} + 2\,I^- \longrightarrow Fe^{2+} + I_2$

Data Organization Suggestions

For the initial observations construct a table with columns for "Compound Name," "Formula," and "Description." Leave space between each data entry row to make the information easier to read.

For this experiment the table given in the reaction guide is not sufficient for the amount of data you will be collecting. Try building your own table. Make column headings for additions of KI, starch, HCl, $NaHSO_3$, and $Na_2S_2O_3$. The row heading can be the same as they are in the reaction guide. Once again, your table should be easy to read. Don't crowd the data into too small a space.

Data Analysis

Answer the following questions in your laboratory notebook using complete sentences.

1. What color is aqueous I_2? What color is I_2 in the presence of starch?
2. What color did the mixtures turn upon the addition of $NaHSO_3$ or $Na_2S_2O_3$? What does this indicate?

Analysis steps 3, 4, 5, and 6 can be done separately as individual exercises. Alternatively you can take reactions a – h, and those in question 4, and individually run through step 3, 5, and 6 for each reaction.

3. Listed below are the reactants and products for each mixing. Assign oxidation numbers to each chemical species that changes oxidation number and then write the half-reactions for reach reaction.

   a. $OCl^- + 2\ I^- \rightarrow Cl^- + I_2$

   b. $MnO_4^- + 2\ I^- \rightarrow Mn^{2+} + I_2$

   c. $H_2O_2 + I^- \rightarrow H_2O + I_2$

   d. $IO_3^- + 2\ I^- \rightarrow I^- + I_2$

   e. $Cr_2O_7^{2-} + 2\ I^- \rightarrow 2\ Cr^{3+} + I_2$

   f. $NO_2^- + 2\ I^- \rightarrow NO_{(g)} + I_2$

   g. $Cu^{2+} + 3\ I^- \rightarrow CuI_{(s)} + I_2$

   h. $Fe^{3+} + 2\ I^- \rightarrow Fe^{2+} + I_2$

4. Given the reactants and products for Steps 5 and 6, write the half-reactions for each.

$HSO_3^- + I_2 \rightarrow SO_4^{2-} + 2\,I^-$ $S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2\,I^-$

5. Complete and balance the redox reactions for each reaction you observed. Multiply each half-reaction by coefficients that balance the electrons lost or gained. Then, add the half-reactions. For example,

$$\begin{array}{rcll} 2\left[MnO_4^- + 5\,e^-\right. & \longrightarrow & \left.Mn^{2+}\right] & \text{reduction} \\ 5\left[2\,I^-\right. & \longrightarrow & \left.I_2 + 2\,e^-\right] & \text{oxidation} \\ \hline 2\,MnO_4^- + 10\,I^- & \longrightarrow & 2\,Mn^{2+} + 5\,I_2 & \end{array}$$

6. Add water to one side to balance the oxygens, and add H^+ to the other side to balance the hydrogens:

$$16\,H^+ + 2\,MnO_4^- + 10\,I^- \longrightarrow 2\,Mn^{2+} + 5\,I_2 + 8\,H_2O$$

Check to see that all atoms are balanced and total charge is balanced.

7. Make a list of the oxidizing agents and a list of the reducing agents used in this experiment.

Density

Goals

- ☐ The purpose of this experiment is to investigate the topic of density by determining the densities of some materials.

Introduction

There is a lesson in the old bromide: “Which is heavier, a pound of feathers or a bound of lead?” Neither is heavier since a pound of anything is still a pound. It is not possible to compare masses of objects unless it is done on the basis of the same portion of each object. For instance, it is possible to say that one person is heavier than another person since the comparison is on a pound per person basis. Since all materials have mass and occupy space, it is possible to compare materials by stating the amount of mass contained in a specific volume of each material. Density is defined as the mass per unit volume of a material, where mass is usually expressed in grams and volume in cubic centimeters or milliliters. Density is determined by measuring the mass and volume of a sample of a material and then dividing the mass by the volume:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \qquad \text{or} \qquad D = \frac{m}{V}$$

The density of any material is a characteristic property of the material and can help identify the material. Furthermore, density is used to relate the mass of a sample of the material to the volume or the volume to the mass. That is, the density is used as a property conversion factor to find the volume given the mass,

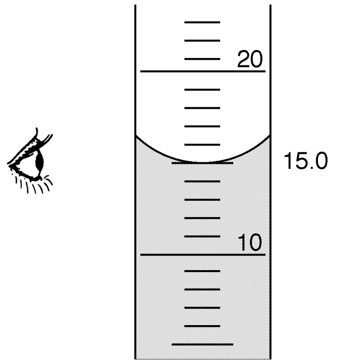
$$V = \frac{m}{D}$$

or to find the mass given the volume,

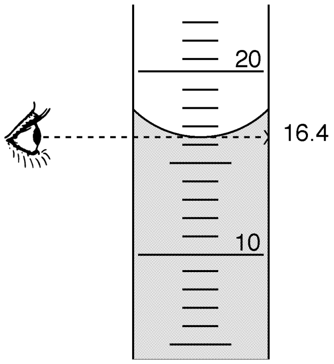
$$m = V \times D$$

In this experiment, the densities of several materials are determined and densities are used as conversion factors.

The volumes of liquids are often determined in the laboratory by use of graduated cylinders. These cylinders are calibrated in the factory so that they will contain certain specific volumes. Lines are etched on the outside surface of the cylinder to indicate the positions corresponding to various volumes. Distances between subsequent lines on the cylinder correspond to specific volumes. Volumes of liquids are measured by pouring a liquid sample into the cylinder and observing the position of the surface of the liquid. Some liquids, such as water, do not form a flat surface called a meniscus. Consequently, glass graduated cylinders are calibrated so that the volumes are read by observing the lowest portion of the curved meniscus. In other words, read the bottom of the meniscus. The volume of the liquid is measured by observing the position of the meniscus with respect to the calibration lines on the cylinder. Consider an example of the reading of the volume of a liquid in a graduated cylinder. In Fig. 2-1b, the position of meniscus coincides with the 15-mL mark and the volume of the liquid would be read as line but falls between lines. The estimation of the position of the meniscus illustrates a very important method involved in the measurement called interpolation. See the discussion at the beginning of Appendix 1. Interpolation is accomplished by visually splitting the distance between two lines into equal parts and then approximating the position at which the meniscus lies between the two lines. In the figure, if the distance between lines is split into 10 parts then the meniscus appears to lie at the 16.4-mL position. The interpolation method is very important in volume measurements. The number of digits to be read by interpolation depends on how the volume measuring device is calibrated. When in doubt about how to read a cylinder, consult you instructor.



(a)



(b)

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

Materials:

Reagent Central solutions include:
Unknown solution

Equipment: Graduated cylinder pennies aluminum foil paper samples

Experimental Procedure

Measure all masses to the nearest 0.01 g. Be careful to use the balances correctly. Your instructor will demonstrate the proper use of the laboratory balances. Balances are delicate instruments, so use them with care and respect. It is a good idea to use the same balance for subsequent weighings. The volumes of liquids are measured using graduated cylinders or a calibrated syringe. Since the number of significant digits in the volumes will dictate the number of digits in your calculated densities, be sure to read the volumes carefully.

1. DENSITY OF WATER

Dry your 10-mL graduated cylinder and find its mass to the nearest 0.01 g. Pour about 9 mL of water into the cylinder and carefully read the volume to the nearest 0.1 mL. Now weigh the cylinder plus the water to the nearest 0.01 g. Be sure to record your data in your notebook. Repeat the experiment three times.

2. DENSITY OF AN UNKNOWN LIQUID

Obtain a sample of unknown liquid and find its density using the same method you used for water.

3. THE VOLUME OF A DROP OF WATER

You need a small container and a disposable pipet, and a measuring syringe or an eye dropper. To make a convenient pipet, cut off half of the tip of a disposable pipet to leave a tip about 5 cm long. Weigh the container to the nearest 0.01 g.

Draw some water into the pipet, syringe, or dropper and practice dispensing drops until you make drops of the same size. Add 30 drops of water to the container and weigh it. You may add a few more or less than 30 drops but keep an accurate count on the number of drops you add. Repeat the exercise using another 30 drop sample.

4. DENSITY OF PENNIES

You need 15 post 1983 pennies, a small container, a balance, and a plastic metric rule. You are going to determine the density of the penny sample by measuring the mass of the sample and finding the volume by stacking the pennies to make a cylinder.
Assuming that the stack of pennies is a cylinder, the volume is calculated by the formula

$$V = \pi (d/2)^2 h$$

where V is the volume, π is 3.142, d is the diameter, and h is the height.

Weigh an empty container to the nearest 0.01 g. Place the 15 pennies into the container and weigh it again to the nearest 0.01 g.

Use the metric rule to measure the diameter of a penny to the nearest 1 mm. Stack the pennies into a cylinder and devise a way to measure the height of the cylinder to the nearest 1 mm.

5. THICKNESS OF METAL FOIL

To directly measure the thickness of a sample of aluminum foil a special instrument is needed. However, the thickness can be simulated indirectly by viewing a sheet of foil as a thin rectangular solid having a length, width, and height. The height is the thickness. Since density relates mass to volume, the volume of a rectangular sample can be found from its mass. If we know the length, width, and volume of a rectangular solid it is possible to calculate the height.

Obtain a sheet of aluminum foil having a regular shape. Smooth the sheet on your laboratory book and, using a metric ruler, carefully measure the length of the edge of the sheet to the nearest 0.01 cm. If the sheet is not square, measure the length of adjacent edges. Use the balance to find the mass of the sheet. Do not crease or crumple the sheet. Record your data.

Data Organization Suggestions

Create tables for your numerical data as shown below. When taking numerical data be sure to record the value and the unit. Every number should have a unit written after it! Separate these small experiments 1 per page, allowing enough room for calculations below your data.

1. Density of Water	Trial 1	Trial 2
Mass of cylinder + water		
Mass of cylinder		
Mass of water		
Volume of water		

Data Analysis

1. Density of water

Calculate the mass of the water sample and use the mass and volume to find the density of the water for each of the three trials. Report the average density.

2. Density of unknown liquid

Calculate the mass of the liquid sample and use the data to find the density of the liquid for eachof the three trials. Report the average density.

3. Volume of a drop of water

Use your data to find the mass of the 30 drops and calculate the mass per drop. Using a density for water of 1.00 g/mL, calculate the average volume of one of your drops of water.

Calculate the average volume of a drop that you dispense from your pipet, syringe, or dropper. Also, calculate the number of your drops per 1.0 mL of water.

4. Density of pennies

Calculate the mass of the sample and the volume of the cylinder. Calculate the density of the pennies by dividing the mass of the sample by the volume. A table of the densities of some common metals, in grams per cubic centimeter, is given below. Which metal has a density comparable to your calculated density?

Aluminum 2.7	Magnesium 1.7	Copper 8.9	Nickel 8.9	Silver 10.5
iron 7.9	Zinc 7.1	Titanium 4.5	Gold 19.3	Tungsten 19.1

5. Thickness of aluminum foil

Calculate the area, A, of the sheet from the length of the edge or edges. The area is the product of the length, l, and the width, w, or the square of the length of the edge if the sheet is a square:

$$A = l \times w \text{ or } A = l^2$$

Calculate the approximate thickness of the foil as follows. Assuming that the foil is a rectangular solid, the volume is given by the area of the face times the height (thickness). This can be expressed as:

$$V = At$$

Where V is the volume, A is the area, and t is the thickness. If the density of aluminum is 2.70 g/cm^3, the volume of the foil sample is determined from the mass and the density. Once the volume is known, it is used along with the area to find the thickness of the foil.

$$t = V/A$$

Calculate the thickness of the aluminum foil in centimeters. Express the thickness in terms of millimeters.

Questions for Analysis

Answer the following questions in your laboratory notebook:

1. Mercury is a liquid metal having a density of 13.6 g/mL. What is the volume of 5.00 lb of mercury metal? (453.6 g/1 lb)

2. A sample of aluminum has a mass of 14.24 g and a volume of 5.27 cm^3. What is the mass of a 49.7 cm^3 sample of aluminum?

3. A piece of magnesium is in the shape of a cylinder with a height of 7.13 cm and a diameter of 1.63 cm. If the magnesium sample has a mass of 26.4 g, what is the density of the sample? (See part 4 of the experiment)

4. A sample of aluminum foil has the dimensions of 12.1 cm × 16.2 cm. If the mass of this sample is measured to be 0.12 g and the density is known to be 2.70 g/cm^3, what is the thickness of the foil in millimeters to the correct number of significant figures?

Measuring Mass: A Means of Counting

Goals

- ☐ Properly use a top loading balance to determine the mass of a sample .
- ☐ Use molar masses to connect the measured mass of a sample to the number of particles in that sample.
- ☐ Use safe and good lab practice to characterize matter.
- ☐ Apply principles to samples of both pure substances and mixtures.
- ☐ Apply dimensional analysis techniques to count small particles such as atoms and molecules.

Introduction

Our world contains groupings of objects everywhere, a dozen eggs, a pair of socks, a gross of pencils. These collections are convenient "packets" of individual pieces. The individual pieces of pure substances can be described by chemical formulas, e.g., H_2O is the chemical formula for water. This formula indicates that each molecule of water consists of two atoms of hydrogen combined with one atom of oxygen. The mass of this molecule is the sum of the masses of the atoms combined to form this compound. We cannot directly measure the mass of one molecule of water but we can recognize its relative mass and use a convenient "packet" of molecules to describe real world quantities. The **mole** is the chemist's standard collection of particles and is defined **as the amount of substance in a sample that contains as many units as there are atoms in exactly 12 grams of carbon-12**. That number of carbon-12 atoms is **6.022×10^{23}** and is known as **Avogadro's number**.

1 mole carbon atoms = 12.0 g C = 6.02×10^{23} atoms C

1 mole H_2O = 2(1.01 g H) + 1(16.00 g O) = 18.02 g H_2O = 6.022×10^{23} molecules of water

Using these relationships, any mass of water can be converted into a number of molecules:

$$100.00 \text{ g } H_2O \left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right)\left(\frac{6.022 \times 10^{23}}{1 \text{ mol}}\right) = 3.34 \times 10^{24} \text{ molecules } H_2O$$

In this lab you will measure amounts of substances. You will then calculate the number of particles contained in the sample, numbers that cannot be counted-- only calculated.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

Materials:

Reagent Central solutions include:

sucrose ($C_{12}H_{22}O_{11}$), sodium chloride (NaCl), chalk (calcium carbonate)

✓**Check out** a sample containing:
Glass slides (assumed to be pure silicon dioxide), polystyrene peanuts, sulfur, fluorite, hematite (or other minerals as provided by stockroom)

Equipment:	Balance	Plastic spoons

Experimental Procedure

1. Using a weighing paper or boat and balance, "weigh" one level teaspoon of sodium chloride and record its mass in your laboratory notebook. Using the same balance as before measure the mass of one teaspoon of water and one of sucrose. Record your measurements.
2. "Weigh" a glass slide, and record its mass in your laboratory notebook. Repeat for the piece of chalk and a polystyrene peanut.
3. "Weigh" a piece of sulfur, and record its mass in your laboratory notebook. Repeat for a piece of fluorite and a piece of hematite.
4. A nickel coin is a mixture of metals called an alloy. It consists of 75% copper and 25% nickel. Design and carry out an experiment to find out how many nickel atoms there are in one 5-cent piece. Record your experiment procedure in your laboratory notebook. Show all your calculations and give your final answer with the correct number of significant figures and in scientific notation.

Data Analysis

For each of the masses recorded perform the following calculations. Report your answers with the proper number of significant figures.
Use the formula to determine the molar mass in units of g/mol.
Use the molar mass to determine the number of moles in the sample.
Use the number of moles of the substance and molar ratios to calculate the moles of each element within the sample.
Use the moles of each element along with Avogadro's number to calculate the number of atoms of each element within the sample.

Acid-Base Titration

Goals

- ☐ To practice the titration method of analysis and to use the method to analyze a vinegar solution.

Terms to Know

Titration - The process of adding a measured amount of a solution of known concentration to a sample of another solution for purposes of determining the concentration of the solution or the amount of some species in the solution.

Titrant - The solution measured from the buret in a titration.

Buret - A narrow, cylindrical-shaped, precisely calibrated piece of glassware. A device used to measure the volume of titrant delivered in a titration.

End Point - The stage of a titration when just enough titrant has been added to make the indicator change.

Indicator - A chemical, added to a titration mixture, which changes color at the end point of the titration.

Introduction

A titration is the process of adding a measured volume of a solution of one chemical species to a sample of a solution containing another species for the purpose of determining the concentration of the dissolved species in one of the solutions. A species in the solution of known concentration reacts with another species in the unknown solution. The addition and measurement of the volume of the added solution is done by use of a buret. (See Fig. 29-1.) A titration is usually carried out by placing a measured sample of the unknown solution in a flask, filling the buret with the known solution (called the titrant), and then slowly delivering the titrant to the flask until the necessary amount has been added to the unknown solution.

The point at which the necessary amount has been added is called the end point of the titration. The end point is often detected by placing a small amount of a chemical called an indicator in the reaction flask. The indicator is chosen so that it will react with the titrant when the end point is reached. The reaction of the indicator produces a colored product; the appearance of the color signals the end point of the titration. Some indicators are colored to begin with and react at the end point to produce a different colored product, so the change in color indicates the end point. Other indicators change from a colorless form to a colored form at the end point.

Once the end point has been found, the volume of titrant used can be determined from the buret. Using this volume, the concentration of the known solution and the stoichiometric factor from the balanced equation, we can deduce the number of moles of species in the solution being analyzed. If the molarity of the unknown solution is to be calculated, it is necessary to measure the volume of the original sample of unknown solution before it was titrated. Then, the molarity can be found by dividing the calculated number of moles by this volume. As an example, consider the following case. Determine the molarity of a hydrochloric acid solution if 30.21 mL of a 0.200 *M* sodium hydroxide solution is needed to titrate a 25.00-mL sample of the acid solution. First, the chemical reaction involved is:

$$OH^-_{(aq)} + H_3O^+_{(aq)} \rightleftharpoons 2\,H_2O$$

Note that one mole of acid reacts with one mole of base. The number of moles of hydroxide ion needed to react can be found from the volume used and the molarity. The titration required 30.21 mL of 0.200 *M* NaOH to react with the acid. The number of moles of hydroxide ion is found by multiplying the volume of sodium hydroxide solution in liters by the molarity. The calculations are:

First the number of moles of OH^- used are found as the product of the volume and molarity: $V_b M_b$

$$\frac{30.21\ \text{mL} \times 0.200\ \text{mol OH}^-}{1000\ \text{mL}} = 0.006042\ \text{mol OH}^-$$

Next the number of moles of hydronium ion is found by multiplying by the molar ratio obtained from the equation for the titration reaction.

$$0.006042\ \text{mol OH}^- \times \frac{1\ \text{mol H}_3\text{O}^+}{1\text{mol OH}^-} = 0.006042\ \text{mol H}_3\text{O}^+$$

Finally, the molarity of the acid solution can be found by dividing the number of moles of hydronium ion by the original volume of the acid solution in liters.

$$\frac{0.006042\ \text{mol H}_3\text{O}^+}{25.00\ \text{mL}} \times \frac{1000\ \text{mL}}{1\ \text{L}} = 0.242\ M\ \text{H}_3\text{O}^+$$

If the acid and base in a titration react in a one-to-one molar ratio, as is the case in the above example, the calculations can be simplified by using the equation:

$$M_a = \frac{M_b\, V_b}{V_a}$$

where V_b is the volume of the base used in the titration, V_a is the volume of the original acid solution, M_b is the molarity of the base solution, and M_a is the molarity of the acid solution. Using this simple equation for the example given above, the molarity of the acid solution is:

$$M_a = \frac{0.200\ M_b \times 30.21\ \text{mL}}{25.00\ \text{mL}} = 0.242\ M$$

How to use the buret

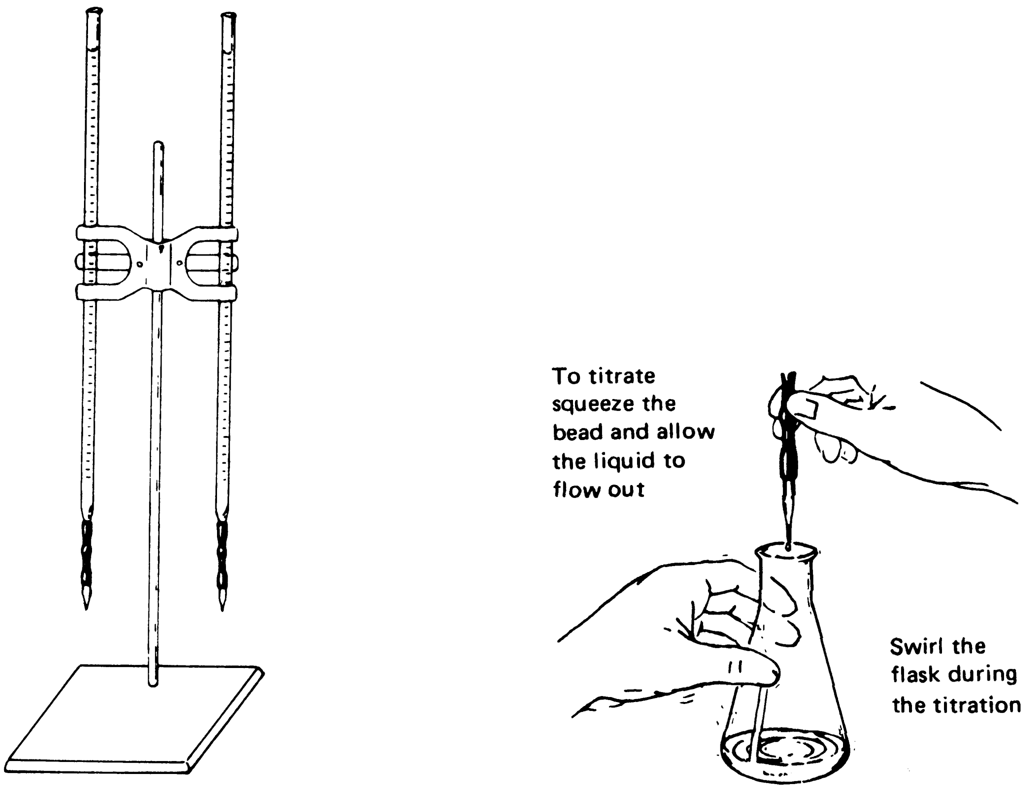
A buret is a piece of glass tubing calibrated to deliver measured volumes of solution. A 50-mL buret is calibrated to deliver between 0 and 50 mL. Each etched line on the buret corresponds to a 0.1-mL increment of volume. By interpolation, the buret can be read to the nearest 0.01 mL. These instructions apply to a Mohr buret as pictured in Figure 29-1. If you use a buret with a glass or plastic stopcock, disregard the reference to the rubber tubing and bead but follow the other instructions.

1. Cleaning

Place some water in the buret and allow it to run out through the tip by squeezing the rubber tube around the glass bead (or opening the stopcock). If you notice water drops adhering to the sides of the buret as it is drained, the buret needs cleaning. Clean the buret using tap water and a small amount of detergent. Use a buret brush to scrub the inside of the buret. Rinse the buret four or five times with tap water and allow some water to run out of the tip each time. Finally, rinse the buret three or four times with 10 mL portions of deionized water. By rotating the buret, allow the water to rinse the entire buret and be sure to rinse the tip by allowing some water to pass through.

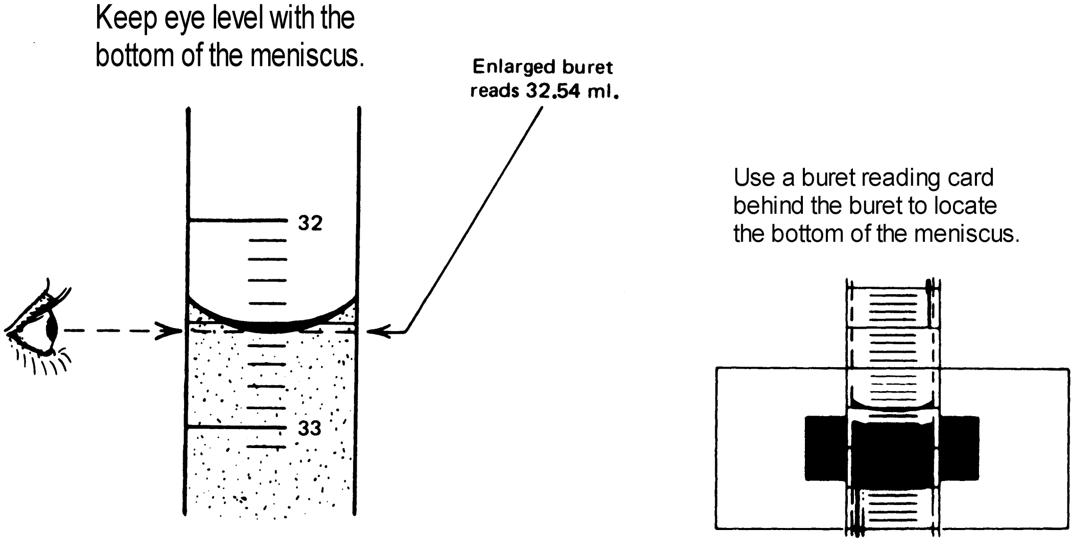
2. Filling

When filling a buret, try not to splash or spill any titrant. Clean up any spills. Use a paper towel to dry the outside of the buret. Rinse the clean buret two or three times with less than 5 mL portions of the solution with which it is to be filled. Be sure to rinse the tip each time. Fill the buret to a level just above the 0.00-mL mark. Fill the tip by bending the rubber to point the tip upward and then gently squeeze the tube to allow the liquid to displace the air (or open the stopcock and remove any air bubbles). Adjust the meniscus to a position somewhat below the 0.00-mL mark.



(a) (b)

Figure 29-1 Burets and the titration process



(a) (b)

Figure 29-2 Reading a buret

3. Reading

The position of the bottom of the meniscus is read to the nearest 0.01 mL. The buret can be read directly to the nearest 0.1 mL, but you must interpolate to read to the nearest 0.01 mL. To interpolate, you imagine that the distance between lines is made up of 10 equally spaced parts. Then decide with which part the bottom of the meniscus coincides. A buret reading card can be used to aid you in reading the position of the meniscus. (See Fig. 29-2.) When reading the buret, make sure your eye is level with the bottom of the meniscus.

4. Titration Method

Be sure to record the initial and final buret readings when you do a titration. Place the sample of solution to be titrated in a flask. After recording the initial volume of the buret, allow the titrant to flow into the flask by pinching the rubber tube around the glass bead (or opening the stopcock). Let the titrant flow rapidly at first, and then add smaller and smaller volume increments as the end point is approached. When you are close to the end point, add one drop or less at a time. A fraction of a drop can be added by allowing a portion of a drop to form on the tip, touching the tip to the inside of the flask, and then washing down the sides of the flask with a small amount of deionized water from a wash bottle. During the titration, mix the solutions in the flask by swirling but do not splash the solutions out of the flask. (See Fig. 29-1b.)

5. Cleaning Up

Drain the solution from the buret and rinse thoroughly with tap water. Remember to rinse the tip. Try not to spill any of the solutions on the desk or your clothing. If you do spill the solutions, clean them up. If any solutions spill on your clothing, tell your instructor. Burets can be stored corked and filled with deionized water.

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

Materials:

Reagent Central solutions include:

Experimental Procedure

1. Analysis of Acetic Acid in Vinegar

In this experiment you will titrate a vinegar solution using a standardized sodium hydroxide solution. You need a clean, dry 100-mL beaker, a 250-mL beaker, a 250-mL erlenmeyer flask, a 50-mL buret, a buret clamp, and a ring stand. Set up a buret on a ring stand as shown in Fig. 29-1 but use only one buret.

Obtain about a little less than 100 mL of a standardized sodium hydroxide solution in the 100-mL beaker. (DANGER: NaOH is very caustic.) Write down the molarity of this solution. Be very careful not to splash or spill this solution. Sodium hydroxide solutions are especially dangerous if splashed in your eyes. Always pour such solutions in a buret by removing the buret from the clamp and holding the buret in a piece of paper towel well below your eye level. Never raise the beaker of solution above your eye level. Do not put a funnel in the buret to fill it.

Rinse and fill your buret with the sodium hydroxide solution according to the instructions given in the discussion section.

You will find vinegar solution in a bottle fitted with a precision 10.0-mL pump. As demonstrated by your instructor, pump a 10.0-mL sample of vinegar into a clean 250-mL flask. Add 2 or 3 drops of phenolphthalein indicator to the flask and dilute to about 50 mL volume using deionized water.
Take the initial volume reading of the buret. Titrate the acid solution sample with the sodium hydroxide solution until the end point is reached. As you add base, you will notice a slight pink color at the point at which the base solution enters the acid solution. As you approach the end point, you may see the entire solution momentarily become pink as you swirl it. When this occurs, slow down the rate of addition of the base and carefully approach the end point. When the addition of a single drop of base makes the solution in the flask turn a very light pink and remain so for one minute or longer, you have reached the end point. The sides of the flask may be washed down with small amounts of deionized water at any time. Once the end point has been reached, take the final volume reading.

Rinse the 250-mL flask with deionized water. Pump another 10.0 mL of acid into the flask and titrate the acid with the base. You may have to refill the buret with some sodium hydroxide solution. Record the titration data in your laboratory notebook. **Repeat the titration procedure a total of three times.**

Rinse the flask and carry out a third titration. When you are finished be sure to rinse out your buret with water and remember to rinse the tip of the buret so that it does not contain base solution. Test the tip using red litmus paper. If the paper turns blue rinse your buret and tip with more water.

Use the volumes of the acid and base solutions used from the data table for the calculations. The acid/base reaction involved is:

$$HC_2H_3O_{2(aq)} + OH^-_{(aq)} \rightleftharpoons H_2O_{(l)} + C_2H_3O_2^-{}_{(aq)}$$

Calculate the molarity of acetic acid in vinegar for each titration. If your results are within 10% of one another, calculate the average molarity. If your results do not seem close enough, consult your instructor.

2. The Percent By Mass Acetic Acid In Vinegar

From density of vinegar, which is 1.00 g/mL, it is possible to calculate the percent by mass of acetic acid in vinegar. This is done by carrying out the following sequence of calculations. (AA is an abbreviation for acetic acid, $HC_2H_3O_2$)

$$\frac{\text{moles AA} \xrightarrow{\text{molar mass AA}} \text{grams AA}}{\text{L vinegar} \longrightarrow \text{mL vinegar} \xrightarrow[\substack{\text{density}\\ \text{water}}]{} \text{g vinegar}} \times 100 = \%\text{AA}$$

Using your experimental value for the molarity of acetic acid in vinegar, calculate the percent by mass acetic acid in vinegar.

Questions for Analysis

Answer the following questions in your laboratory notebook:

1. Explain the function and purpose of an indicator in a titration.
2. If you are titrating an acid with a base, tell how each of the following factors will affect the calculated molarity of the acid. That is, explain whether the calculated molarity would be greater than it should be, less than it should be, or not changed, and explain why.
    a. The end point is exceeded by adding too much base.
    b. The volume of acid is measured incorrectly so that it is smaller than the value used in the calculations.
    c. Deionized water is used to wash down the sides of the flask during the titration.
3. A sample of vinegar is titrated with a sodium hydroxide solution to find the molarity of acetic acid. If 18.82 mL of a 0.430 *M* NaOH solution is required to titrate 10.00 mL of vinegar solution, what is the molarity of acetic acid in the vinegar?

The Molar Mass of a Gas

Goals

- The purpose of this experiment is to determine the number of grams per mole of a gas by measuring the pressure, volume, temperature, and mass of a sample.

Terms to Know

Molar Mass - The number of grams per mole of a substance.

Ideal Gas Law - The relation between pressure, volume, number of moles, and temperature of a gas: $PV = nRT$

Vapor Pressure of Water - The pressure exerted by water vapor as it saturates the environment above a sample of liquid water. Water vapor pressures depend upon the temperature, and tabulated values are found in reference tables.

Introduction

The molar mass of a gaseous compound can be determined by experiment even though the formula or composition of the compound is not known. In other words, it is possible to find the molar mass of a gas even if we do not know its identity. The molar mass is determined by using the fact that the number of moles of a gas sample relates to the pressure, volume, and temperature by the gas laws. The mass, pressure, volume and temperature of a gas sample are measured experimentally and, used to calculate the molar mass.

To determine the molar mass of a gaseous substance from the mass, volume, temperature, and pressure of a sample, the number of moles in the sample is first calculated using the ideal gas law, $PV = nRT$. Recall that n represents the number of moles. The number of moles of a gas is found from the pressure, volume, and temperature of a sample: $n = PV/RT$ where R is the universal gas constant.

As an example, consider a 0.508 g sample of a gas that occupies 522 mL volume at 100 oC and 0.960 atmospheres. First summarize the data as:

mass	P	V	T
0.508 g	0.960 atm	0.522 L	373 K (100 + 273)

The number of moles of gas is found using the ideal gas law. (Fill in the following blanks and calculate the number of moles.)

$$n = \frac{\quad \text{atm} \times \quad \text{L}}{0.0821 \frac{\text{L atm}}{\text{mole K}} \times \quad \text{K}} = \qquad \text{mole}$$

The molar mass of the gas is found from the ratio of the mass of the sample to the moles in the sample. Divide the mass of the sample by the number of moles using the value calculated above.

$$\frac{0.508\ \text{g}}{\qquad \text{mole}} = \frac{\qquad \text{g}}{1\ \text{mole}}$$

In this experiment, a sample of gas is collected by water displacement. That is, the gas bubbles into a container of water and as the gas accumulates, it displaces the water. The volume of the sample is the volume the gas occupies in the container. The temperature is found by measuring the temperature of the water in contact with the gas. The mass of the gas sample is measured by weighing a small tank or cylinder of gas, delivering the sample, and reweighing the cylinder.

The pressure of the sample is found by making sure that the pressure of the gas sample is the same as the atmospheric pressure. The atmospheric pressure is measured with a barometer. However, when a gas is collected by water displacement, it becomes saturated with water vapor. This means that once the gas sample is collected, it will be a mixture of the gas and water vapor. This does not affect the volume or temperature of the gas but the measured pressure is the pressure of the mixture. That is, the measured pressure is the total pressure of the gas and the water vapor. To determine the pressure of the gas sample, the pressure of the water vapor can be subtracted from the total pressure. Appendix 2 lists the vapor pressure of water at various temperatures. To find the pressure of the gas sample in the experiment, look up the vapor pressure of water at the measured temperature and subtract this pressure from the measured atmospheric pressure:

$$P_{\text{GAS}} = P_{\text{ATM}} - P_{\text{WATER}}$$

Safety

Act in accordance with the laboratory safety rules of Cabrillo College.
Wear safety glasses at all times.
Avoid contact* with all chemical reagents and dispose of reactions using appropriate waste container.

Materials:

Reagent Central solutions include:

Experimental Procedure

For this experiment you will need a 1-L beaker, a 250-mL Erlenmeyer flask, and a thermometer.

Place a piece of magic tape or a length of a gummed label along the neck of the flask near the top. Fill the flask to the brim with tap water and put about 500 mL of water into the beaker. Allow the water to stand for a while to be sure that it is at room temperature. Set up the apparatus as shown in Figure 19-1 below.

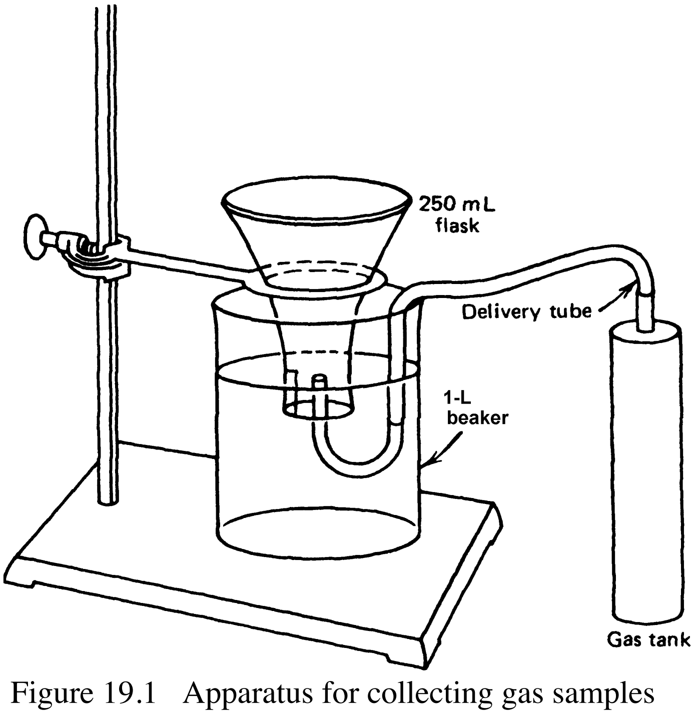


Figure 19.1 Apparatus for collecting gas samples

Gently stopper the flask with a rubber stopper and do not allow any air to become trapped in the flask under the stopper. Invert the flask through the iron ring which will hold it in position. Lower the ring with the flask so that the mouth of the flask is below the water level in the beaker. Secure the ring to support the flask. Use a spatula to remove the rubber stopper from the mouth of the flask. The stopper can remain in the beaker since it will not interfere with the experiment.

Obtain a small tank of gas. Use a paper towel to gently wipe the tank to make sure that it is clean. Weigh the tank to 0.01 g.

Attach a plastic delivery tube to the valve on the tank and insert the glass tube on the other end into the mouth of the flask in the beaker of water. Push the valve on the tank straight down to deliver gas to the flask. Continue to deliver gas until the flask is nearly full of gas. That is, deliver gas until the gas level is near the neck of the flask. Do not allow any gas to escape from the flask. If any gas escapes, you will have to start the experiment over.

Carefully remove the delivery tube from the gas tank, gently clean and dry the tank with a paper towel and weigh the tank to 0.01 g. Place a thermometer in the beaker.

Move the flask of gas so that the level of water in the flask is the same as the water level in the beaker. This will adjust the pressure in the flask to atmospheric pressure. Use a pen or pencil to place a mark on the tape to indicate the water level.

Remove the flask from the water and fill it with water to the level corresponding to the mark on the tape. This volume of water will correspond to the volume of the gas sample. Carefully pour the water from the flask into a graduated cylinder and measure the volume to the nearest 1 mL. Record this as the volume of the gas sample.

Record the temperature of the water in the beaker. Assume that the temperature of the gas sample is the same as the water.

Record the atmospheric pressure and look up the vapor pressure of water at the temperature of the gas in the table at the end of this experiment.

Repeat the experiment to obtain a second set of data. Calculate the molar mass of the gas using each set of data. If the results are within 10 percent of one another, express your answer as the average of the two results. If your two results seem to be too far apart, consult your instructor.

Vapor Pressure of Water

Temperature °C	Vapor Pressure torr	Temperature °C	Vapor Pressure torr
0	4.6	25	23.8
5	6.5	26	25.2
10	9.2	27	26.7
15	12.8	28	28.3
16	13.6	29	30.0
17	14.5	30	31.8
18	15.5	40	55.3
19	16.5	50	92.5
20	17.5	60	149.4
21	18.6	70	233.7
22	19.8	80	355.1
23	21.2	90	525.8
24	22.4	100	760.0

Questions for Analysis

Answer the following questions in your laboratory notebook:

1. Referring to the experimental determination of the molar mass, explain how and why each of the following factors would affect your calculated molar mass. That is, would the calculated value be greater than it should be, less than it should be, or not changed? Note that the relationships between the various factors involved in the calculation are:

   $$\text{molar mass} = \text{g/mol} = \text{g}/(PV/RT) = \frac{gRT}{PV}$$

   a. The measured temperature is a lower value than the actual temperature.
   b. The measured volume is a higher value than the actual volume.
   c. Some of the gas sample escapes from the tank before it reaches the flask.

2. The molar mass of a gas is determined by collecting a gas sample by water displacement.

   a. Using the following data, calculate the molar mass of the gas: Sample volume, 163 mL; temperature, 21.0 °C; mass, 0.281 g; total (i.e., barometric) pressure, 752 torr.
   b. The gas in part (a) contains 85.5% C and 14.5% H. First, determine the empirical formula, and then use the result from part (a) to determine the actual formula.

Appendix

Chemical Changes

9/10/03

Purpose: To observe and record chemical changes.

Procedure: The procedure on page 18 in the manual was followed without modification.

Data:

Chemical Mixings	observations initial	 final
a) $NaHCO_3 + HCl$	Clear and colorless liquid + Clear and colorless liquid	clear and colorless liquid. with gas formed
b) HCl + Blue Dye	clear and colorless liquid	Clear and green liquid
c) $NaOCl$ + Blue Dye ∗	clear and colorless liquid + Clear and Blue liquid	Clear and blue liquid. same color blue as blue Dye.
d) $NaOCl + KI$	clear and colorless liquid + Clear and colorless liquid	clear and yellow liquid
d) —— + starch	clear and yellow liquid	~~solid~~ opaque black liquid
∗ $NaOCl$ + Blue Dye + HCl	clear and Blue liquid Same color as blue dye	Clear and pale yellow liquid
e) $KI + Pb(NO_3)_2$	clear and colorless liquid + clear and colorless liquid	~~clear liq~~ colorless liquid with yellow PPT** formed
f) $Pb(NO_3)_2 + CaCl$	clear and colorless liquid + clear and colorless liquid	colorless liquid with white PPT formed
g) $CaCl_2 + NaHSO_4$	clear and colorless liquid + clear and colorless liquid	colorless liquid with white PPT formed
h) $NaHSO_4 + Na_2CO_3$	clear and colorless liquid + clear and colorless liquid	clear and colorless liquid with gas formed
i) Na_2CO_3 + phen	clear and colorless liquid + clear and colorless liquid	clear Hot pink liquid spreads all over page
j) phen + $NaOH$	clear and colorless liquid + clear and colorless liquid	Clear hot pink liquid

** PPT = precipitate

9/10/03

Chemical Changes

9/10/03

Data:

Chemical Mixings	observations initial	 final
k) $NaOH + AgNO_3$	clear and colorless liquid + clear and colorless liquid	~~solid~~ brown liquid with brown PPT formed
l) $AgNO_3 + NH_3$	clear and colorless liquid + clear and colorless liquid	cloudy mixture +
—— + sunlight	cloudy mixture	Turned brown
m) $CuSO_4 + NH_3$	Clear and pale blue liquid + clear and colorless liquid	~~Clear +~~ colorless liquid with pale blue PPT formed
n) $CuSO_4 + NaHCO_3$	clear and pale blue liquid + clear and colorless liquid	colorless liquid with pale blue PPT formed

9/10/03

Testing Food For starch

9/10/03

Purpose: To test different foods for starch

Procedure: place 2 drops of NaOCl on test surface. Add 2 drops of KI. Mix using air blown by an empty pipet. Add a small amount of cereal. record whether or not the cereal turns black. Repeat with chip, cracker, salt and baking soda.

Data:

Food	Observation	Result (+/-)
Cereal	liquid turned brown, cereal crumbs, black	+
chip	liquid stayed yellow. Chip piece turned black	+
cracker	liquid stayed yellow, cracker piece turned black	+
salt	nothing happened	−
Baking soda	liquid turned a little darker yellow.	−

9/10/03

Chemical Names and Formulas

10/01/03

Purpose: To write chemical names and formulas of common chemical compounds. To describe the colors and textures of common Ionic compounds. To synthesize chemical compounds and write their names and formulas.

Procedure:
① View samples and describe them
② insert experimental page inside reaction surface.
③ place one drop of each solution in the marked boxes.
④ stir with an empty pipet.
⑤ Record changes. ~~and~~
⑥ write the formula for both new products
⑦ figure out which product is the precipitate.

Data:

compound	Formula	Description
• Sodium Acetate or Sodium ethanoate	$NaC_2H_3O_2$	white, clumps together, small and large chunks
• Copper(II) Sulfate	$CuSO_4$	electric blue, small crystals, a little clumpy, easy to break apart.
• Potassium Bromide	KBr	white, large crystals, doesn't stick together at all
• Tin(IV) Chloride Pentahydrate	$SnCl_4$	off white, one large chunk, ppt in test tube
• Sodium dihydrogen phosphate	NaH_2PO_4	White, clumps together, small and large chunks, Mostly powdery.
• Magnesium Sulfate	$MgSO_4$	white, large crystals, not clumpy

JR 10/01/03

10/01/03

Product #1	Product #2	Precipitate
$Fe(NO_3)_3$	$AgCl$	$AgCl$ Silver chloride
$Fe(NO_3)_3$	$PbCl_2$	$PbCl_2$ Lead(II) chloride
KNO_3	~~$PbCl_2$~~ AgI	AgI Silver iodide
[redacted]	[redacted]	[redacted]
Na_2SO_4	$Cu_3(PO_4)_2$	$Cu_3(PO_4)_2$ copper(II) phosphate
Na_2SO_4	$Mg_3(PO_4)_2$	$Mg_3(PO_4)_2$ Magnesium phosphate
$NaCl$	$FePO_4$	$FePO_4$ Iron(III) phosphate

J 10/01/03

Predicting the Products of Precipitating Reactions
Solubility rules

10/8/03

Purpose: To observe and record precipitation reactions. To derive general solubility rules from the experimental data. To describe precipitation reactions by writing net ionic equations. To understand the relationships between solubility and precipitation reactions.

Procedure: The Procedure in the Lab Manual was followed without modification.

Data:

Solutions	Na_2CO_3	$NaCl$	$NaOH$	$NaNO_3$	Na_3PO_4	Na_2SO_4
$AlCl_3$	white suspended ppt	NVR*	white cloudy ppt	NVR	white cloudy ppt	NVR
NH_4Cl	NVR	[redacted]	[redacted]	[redacted]	[redacted]	[redacted]
$CaCl_2$	white suspended ppt	[redacted]	[redacted]	[redacted]	[redacted]	[redacted]
$CuSO_4$	blue suspended ppt	[redacted]	[redacted]	[redacted]	[redacted]	[redacted]
$NaCl$	NVR	[redacted]	[redacted]	[redacted]	[redacted]	[redacted]
$FeCl_3$	brown suspended ppt	[redacted]	[redacted]	[redacted]	[redacted]	[redacted]

*NVR = No visible reaction

R 10/8/03

* $AlCl_3 + Na_3PO_4 \rightarrow AlPO_4 + 3NaCl$

Al	1	Al	1
Cl	3	Cl	~~1~~ 3
Na	3	Na	~~1~~ 3
PO_4	1	PO_4	1

* $CuSO_4 + Na_2CO_3 \rightarrow CuCO_3 + Na_2SO_4$

Cu	1	Cu	1
SO_4	1	SO_4	1
Na	2	Na	2
CO_3	1	CO_3	1

* $CuSO_4 + 2NaOH \rightarrow Cu(OH)_2 + Na_2SO_4$

Cu	1	Cu	1
SO_4	1	SO_4	1
Na	~~1~~ 2	Na	2
OH	~~1~~ 2	OH	2

* $3CuSO_4 + 2Na_3PO_4 \rightarrow Cu_3(PO_4)_2 + 3Na_2SO_4$

Cu	~~1~~ 3	Cu	3
SO_4	~~1~~ 3	SO_4	~~1~~ 3
Na	~~3~~ 6	Na	~~2~~ 6
PO_4	~~1~~ 2	PO_4	2

JR 10/8/03

Chemical and Ionic equations:

- $2AlCl_3 + 3Na_2CO_3 \rightarrow Al_2(CO_3)_3 + 6NaCl$
- $2Al^{3+}_{(aq)} + \underline{6Cl^-_{(aq)}} + \underline{6Na^+_{(aq)}} + 3CO_3^{2-}{}_{(aq)} \rightarrow Al_2(CO_3)_{3(s)} + \underline{6Na^+_{(aq)}} + \underline{6Cl^-_{(aq)}}$
- $2Al^{3+}_{(aq)} + 3CO_3^{2-}{}_{(aq)} \rightarrow Al_2(CO_3)_{(s)}$

- $AlCl_3 + 3NaOH \rightarrow Al(OH)_3 + 3NaCl$
- $Al^{3+}_{(aq)} + 3Cl^-_{(aq)} + 3Na^+_{(aq)} + 3OH^-_{(aq)} \rightarrow Al(OH)_{3(s)} + 3Na^+_{(aq)} + 3Cl^-_{(aq)}$
- $Al^{3+}_{(aq)} + 3OH^-_{(aq)} \rightarrow Al(OH)_{3(s)}$

- [redacted]

- [redacted]

- [redacted]

- $3CaCl_2 + 2Na_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6NaCl$
- $3Ca^{2+}_{(aq)} + 6Cl^-_{(aq)} + 6Na^+_{(aq)} + 2PO_4^{3-}{}_{(aq)} \rightarrow Ca_3(PO_4)_{2(s)} + 6Na^+_{(aq)} + 6Cl^-_{(aq)}$
- $3Ca^{2+}_{(aq)} + 2PO_4^{3-}{}_{(aq)} \rightarrow Ca_3(PO_4)_{2(s)}$

- [redacted]

10/8/03