SCH3U Chemistry 11

Course Notes

2015-2016

York Mills CI

Name: ________________________________

"Every aspect of the world today – even politics & international relations – is affected by chemistry."

LINUS PAULING, (1901 – 1994)

Compiled by Erik Lindala, York Mills CI

Contents

Course Overview

SCH3U - Chemistry Grade 11, University Preparation
York Mills CI
Mr. Lindala (erik.lindala@tdsb.on.ca) room 125, (416) 395-3340 x 20125

Extra help: every day at 12:00 in room 121/123/125

Welcome to the Chemistry 11 course. We will explore the subject of chemistry through class activities, lab work & applications to the real world. A course manual supplements your study with carefully chosen exercises to test & further your knowledge. The course emphasizes collaborative group-based learning. Students will learn to communicate their knowledge effectively.
The prerequisite for this course is grade 10 Academic Science (SNC2D). We strongly recommend a final mark of **75%** or above in grade 10 Science to give you the necessary skills to be successful in a senior science course.

Mark Breakdown:

Knowledge & Understanding	28%	Tests
Thinking	14%	Group work - 5% Weekly quizzes - 5% Projects - 4%
Communication	14%	Tests - 6% Study notes - 4% Assignments - 4%
Application	14%	Labs - 14%
Culminating Exam	30%	Covers the entire course

Course Website:

You'll find full solutions to homework problems, links to videos & other materials used in the course. Please sign into the website as soon as possible & share your login information with your parent/guardian.

1. Go to the website: http://abelmoodle.abel.yorku.ca
2. On the right side of the page click on the **"Create new account"** button below the login button.
3. Use the initial of your first name followed by your last name for your user name (for example jsmith for John Smith). Choose any appropriate password. Remember your password!!Follow the rest of the instructions for logging in.
4. A message will be sent to your e-mail address. Follow the instructions in this message to validate your account.
5. Go back to the website: http://abelmoodle.abel.yorku.ca and login with your username and password.
6. Once you have access to the list of courses click on the Toronto District School Board science category and your course is called: **York Mills – SCH3U**. Click on the course name.
7. Type your enrollment key in the text box. The enrollment key to get in is: **chem11@ym**

Chem 11: Learning Log

One of the most important educational skills you can develop is how to monitor and track your own learning. Either at the end of class or at home, you will complete a daily entry in your learning log.

Written Work: Use our marking scheme for daily class work (out of 5) to assess your written work. What mark do you think your work would receive if it was collected today?
Group Work: Use a marking scheme out of 5 to assess your contribution to the group's work and discussions. Remember that valuable contributions come in many forms: sharing ideas, asking questions, organizing the group, and more. If your teacher was observing your group for the full period, what mark do you think you would get?
Learning Goals / Difficulties: Record the chemistry ideas you feel you learned well today. If you had difficulties with any chemistry ideas make a note so you will remember to get help with them.
Homework: Check off the homework column once you complete the lesson's homework. If you have a problem with a question or get stuck, write down the question number.
Got it: Record the result of your effort to resolve any difficulties you had with the lesson or homework. For example, you might write: help from friend, saw teacher, figured it out, etc. If you had no difficulties or problems to take care of, check off this column indicating that you feel very confident.

Date	Lesson	Written Work	Group Work	Learning Goals / Difficulties	Home-work	Got it
Sept. 10	Atoms Isotopes	4	3	- Didn't understand difference between ions and istopes	Pg 15	Asked group

Toronto District School Board

Science Risk and Safety Notice for Parents/Guardians

School: ______________________ Science Class/Course: ______________

Student's Name: ______________________ Teacher's Name: ______________

An active science program presents some hazards to both staff and students. All attempts will be made however, to identify hazards and manage risks so that they become minimal. Before each activity, instructions will be given to reduce any risks. Teachers will assess the readiness level of students to protect everyone in the class. If a student is considered unready, he or she will not be able to participate in the activity. If no other opportunity to participate can be arranged, then the student's development of hands-on skills may be affected. This may affect the student's achievement.

During an activity, students are expected to follow safe procedures and wear appropriate personal protective equipment (PPE). Student compliance with safe techniques and practices is a part of a teacher's assessment of laboratory work.

The most basic piece of personal protective equipment is a pair of goggles, and these will *always* be made available to students. Like a calculator for mathematics, and running shoes for physical education goggles are personal pieces of equipment best owned by students. When students own their own goggles, they share responsibility for safety in the classroom. Parents can also appreciate that student-owned goggles will result in greater hygiene. Before purchasing their own from an outside source, students should enquire from their teacher on the type of eyewear that is most appropriate for their class.

For the parent/guardian to complete:

1. Does your child/ward have any known allergies to chemicals? ☐ YES ☐ NO
   If yes, please state the allergies: ______________________

2. Does your child/ward wear contact lenses or prescription glasses? ☐ YES ☐ NO
   Please note:
   - Contact lenses can be a problem in the event of an eye injury. Students with contact lenses should wear the same eye protection as the rest of the class.
   - Students wearing prescription glasses are expected to wear eye protection that fits over their eyewear.

3. Are you aware of anything else that the science teacher should know to help maintain a safe classroom? ☐ YES ☐ NO

   If yes, please explain: ______________________

I, ______________________ understand and agree to follow the attached TDSB Science Lab Safety Rules to this contract.

Student's Signature: ______________________**Date:** ______________

I, the parent/guardian of ______________________ have read, understood and discussed the attached TDSB Science Lab Safety Rules with my child.

Parent's Signature: ______________________ **Date:** ______________

Science Lab Safety Rules

1. Come prepared…Stay Focussed

- ✓ Complete all assigned pre-lab tasks.
- ✓ Wear personal protective equipment (PPE) when instructed to do so.
- ✓ Wear closed-toe shoes during lab activities.
- ✓ Make sure long hair is tied or pined back and loose clothing tucked in.
- ✓ Remove all dangling materials from your body (e.g. ear buds, jewellery).
- ✓ Stand when doing lab work and avoid "childish" play.

2. Keep your teacher informed…*Many eyes make safe science*

- ✓ Report all accidents, spills and breakages at once. Never attempt to handle broken glass, sharp metal pieces, or spills on your own. Alert the teacher immediately.
- ✓ If any chemical gets on your hands or any body part, inform teacher and follow his or her direction immediately.
- ✓ Report any damaged or defective equipment immediately to teacher.

3. Follow proper safety etiquettes...at all times

- ✓ Never eat or drink in the science lab.
- ✓ Ensure all equipment and chemicals are put away or disposed of appropriately at the end of lab.
- ✓ Ensure all lab work stations are clean and tidy.
- ✓ Ensure all hot plates/Bunsen burners, gas valves, water and electrical equipment are turned off when not in use.
- ✓ Wash your hands after practical work.

If you don't know or understand…Please Ask!

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

SCH3U – Introduction to Grade 11 Chemistry

In this course, you'll be engaged in inquiry lessons as part of a group. To structure our groups, everyone has a role to play. To start, randomly assign each member one of the following roles: Manager, Speaker & Recorder. Enter your names at the top of the page.

Each group needs a white board, markers & a molecular model kit.

1. The student in the role of **manager** says the following:
   "Today we will review grade 10 chemistry. Using the model kit & the colours listed inside the lid, construct the molecule C_3H_8. Make sure each hole in the atom is filled with a bond."

2. When you've completed the model, the speaker will call over the teacher to check your work.

3. Manager: "Now we need to represent our 3-D model on the whiteboard. We need to create at least three representations of the molecule C_3H_8. One of the representations must be able to be sent in a text message. First we'll brainstorm the drawings & finalize the three versions we'll present to the class".

4. Recorder: draw the final good copy versions on the whiteboard large enough so that the entire class can see it. Ask follow up questions like "Is this how it's drawn?" to make sure all group members agree with the answers.

5. Speaker: You'll present & explain the three drawings to the class.

6. After the class discussion, write your final drawings below. Circle the best representation for C_3H_8.

7. Manager: "Now we'll build & draw the following molecules. Once we agree on the drawing we'll write our good copy below":
   a. CH_3OH b. NH_3 c. CO_2 d. H_2O

8. Summary: on the page below, write down what you think the responsibility for each role in the group is:

   Manager: Speaker: Recorder:

9. Match the role to the action (use M, S, R or all):

   a. Keeps the group on-task & working well: ______

   b. Summarizes the groups ideas: ______

   c. Contributes ideas to the group: ______

   d. Prepares the final copy of the work for marking by the teacher: ______

   e. Presents group ideas to the class: ______

   f. Makes sure all members understand before moving on to new tasks: ______

10. Each class, groups will be chosen at random to submit their work for evaluation. Look at the marking scheme at the front of the room. Note that your final mark is the LOWEST of any of the categories.

11. What mark would you give your group based on this activity? Justify the mark in a few sentences.

12. This activity helps us visualize chemicals that are invisible to the naked eye. Watch the video to see many chemical reactions we will investigate in this course (visit http://beautifulchemistry.net/).

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Measurement & Precision

Group work – many students have had bad experiences with group work. Complete the following on the whiteboard with your group:

Problems with group work (list three):	How to avoid this problem:

Today we'll explore how to make precise measurements in chemistry & how to record them. Remember that your group may be randomly selected to hand in your work at the end of the period as part of your group work mark for the course!

1. Look at the glassware on your desk. Sketch them below & label each.

2. Rank them in order of increasing precision. Explain your order.

3. What is the smallest increment you could measure using your glassware?

4. If you needed to measure exactly 9 mL of water, which glassware would you use? Why?

5. In a lab you made the following measurements: 8.99 mL, 7.8 mL, 9.0 mL & 7.75 mL.
   Using the rules listed in your manual, determine the average volume. Explain how you decided how many significant digits are in your final answer.

6. Record your measurement of the volume of water in the burette at the front of the room. ________

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Thickness of Aluminum Foil

Your group's task today is to determine the thickness of a small piece of aluminum foil. The goals are to create a lab procedure & to measure precisely.

1. On your whiteboard, brainstorm how you could determine the thickness of aluminum foil only using the materials provided. You will share your ideas with the class.

2. Based on the class discussion, write your method flowchart in the space provided. Use your manual as a guide. Remember your method flowchart must be detailed & accurate enough so a grade 11 chemistry student from another class could perform the lab.

Materials	Action

3. Create a simple data table to record your results. Include a table number & descriptive title. You only include measurements (with attention to significant digits) & no calculations in your table. Include units. Use a ruler to create a neat table.

4. When your flowchart & data table are complete, ask the recorder to show it to your teacher for approval before starting the lab.

5. Calculate the aluminum foil thickness below using the problem solving format described in the course manual. Report your final answer with the correct number of significant digits. Explain briefly how you decided on the number of significant digits.

6. Calculate your group's percent error using the equation & value on the board.

7. Every lab has some uncertainty. Assuming your procedure was followed carefully, identify two steps that limited the precision in your final answer. Be clear whether each step decreased *or* increased the final value calculated for the aluminum foil thickness.

8. Think of any assumptions that were made about the aluminum foil.

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Atoms & Their Isotopes

Atoms & isotopes are identified by the numbers of protons, neutrons & electrons that they contain. The isotopic notation for an atom includes the following information:

- symbol of the element;
- the element's atomic number (Z) which specifies the number of protons in the nucleus; &
- the mass number (A) which indicates the number of protons plus neutrons in the nucleus

The number of electrons in a neutral atom is equal to the number of protons in the nucleus of the atom. The mass contributed by the electrons in an atom is very small, so it is not included when calculating the mass number.

Atomic Symbol Notation:

$$^{A}_{Z}X$$

A = Mass Number
X = Atomic Symbol
Z = Atomic Number

Subatomic Particles:

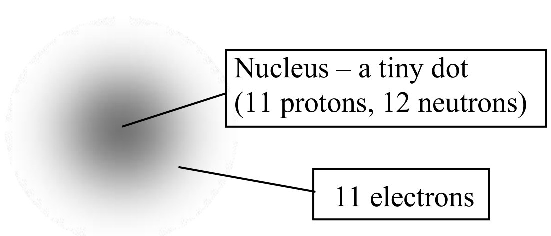
Particle	Symbol	Relative Charge	Absolute Mass	Relative Mass
electron	e^-	-1	9.109×10^{-31} kg	0
proton	p^+	+1	1.673×10^{-27} kg	1
neutron	n^0	0	1.675×10^{-27} kg	1

Note: the diameter of an atom is about 10,000 times larger than the diameter of the atomic nucleus so the relative sizes of the atom & the nucleus are not accurately depicted in the following diagrams.

Model: Two Isotopes of Sodium. The diagrams below show representations of sodium isotopes.

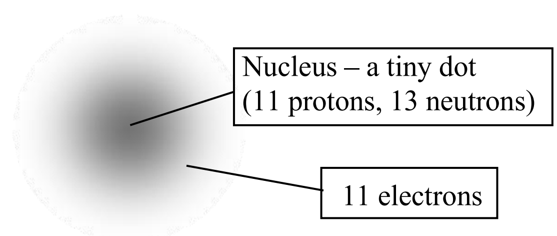
Isotope 1

$^{23}_{11}Na$



Isotope 2

$^{24}_{11}Na$



Key Questions

1. What information is provided by the atomic number, Z?

2. What information is provided by the mass number, A?

3. What is the relationship between the number of protons & the number of electrons in an atom?

4. Because of the relationship between the number of protons & number of electrons in an atom, what is the electrical charge of an atom?

5. What kind of ion is formed if the number of protons is greater than the number of electrons?

6. How many protons, electrons and neutrons are in the O^{2-} ion?

7. What do the two sodium isotopes shown in the model have in common with each other?

8. How do the two sodium isotopes shown in the model differ from each other?

9. What distinguishes an atom of one element from an atom of another element?

Exercises

1. Describe the similarities between chlorine-35 & chlorine-37.

2. Describe the differences between chlorine-35 & chlorine-37.

3. Write the atomic symbols for two isotopes of carbon, C, one with 6 neutrons & the other with 7 neutrons.

4. Use a periodic table to fill in the missing information in the following table.

Name	Symbol	Atomic Number Z	Mass Number A	Number of Neutrons	Number of Electrons
oxygen	^{16}O	8	16	8	8
		7		7	
		1		1	
		1	3		
		12	24		
		12	25		
			238		92

Problems

1. The radius of a Cℓ nucleus is 4.0 fm, & the radius of a Cℓ atom is 100 pm. How many times larger is the diameter of the chlorine atom than the diameter of the chlorine nucleus?
   ($1 \text{ fm} = 1 \times 10^{-15}$ m; $1 \text{ pm} = 1 \times 10^{-12}$ m)

2. How many times larger is the volume of the atom than the volume of the nucleus?

Authored by: Dr. Stephen Prilliman; Revised by: Josephine Parlagreco, Lizabeth Tumminello

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Isotopic Abundance

1. State in your own words what an isotope is:

2. Are the mass numbers (A) of isotopes of an atom the same? Explain.

3. What unit is used to measure the mass of an isotope?

Some isotopes are more common than others. For example, 19.78% of the boron on earth is the boron-10 form and the remaining 80.22% is boron-11.

Isotopic abundance states how commonly an isotope occurs on Earth.

Example: Calculating Relative Atomic Mass:

4. Use the AAM formula to calculate the average atomic mass of boron if the percentage abundances of boron-10 is 19.78% & boron-11 is 80.22%.

Weighted Average Formula:

Average Atomic Mass = Σ (mass of isotope) (% abundance of isotope in decimal form)

5. Notice that the atomic mass of boron found on the periodic table is not a whole number, but a decimal. The value on the table is called the atomic weight. Explain the difference between the mass number and atomic weight.

6. What is the Average Atomic Mass (AAM) of carbon? Check the periodic table. If carbon exists in three isotopes (carbon-12, carbon-13 & carbon-14), which isotope is the most common? Why?

7. If we average the mass of the isotopes of carbon (12 + 13 + 14 / 3), we get 13 u. What is misleading about this number?

<u>Example: Calculating Isotopic Abundance</u>

Sometimes we need grade 10 algebra to answer these questions. Create a "let x =" statement for the following problem. Express the other isotopes in terms of x & solve the following:

8. If the average atomic mass of magnesium is 24.31 u, calculate the % abundances of magnesium-24 & magnesium-25.

9. The average atomic mass of lithium is 6.94 u. Suppose there are three isotopes of lithium: lithium-5, lithium-6 & lithium-7. If lithium-5 is half as abundant as lithium-6, what are the percentage abundances of each isotope of lithium?

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Periodic Trends

The periodic table (PT) list elements in order of increasing ________________ number. The elements are placed in vertical columns called ________________ & horizontal rows called ________________.

Elements in the same group share similar physical & chemical properties.

Trend #1: Atomic Radius (Size)

1. Draw the Bohr-Rutherford (B-R) diagrams for the elements listed below.
2. Label the number of protons & **valence** electrons for each element.

Li	Be	B	C	N	O	F	Ne
p^+ = e^- =	p^+ = e^- =	p^+ = e^- =	p^+ = e^- =	p^+ = e^- =	p^+ = e^- =	p^+ = e^- =	p^+ = e^- =
Na p^+ = e^- =							
K p^+ = e^- =							

3. Look at the B-R diagrams down a group. Does the atomic size get larger or smaller as you go down a group? Explain.

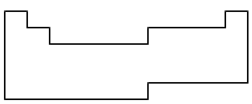
4. Look at the number of protons & valence electrons as you go left to right across a period. What is the general trend?

5. How does the group number correspond to the number of valence electrons?

The attraction of the valence electrons by the protons in the nucleus is called the effective nuclear charge. The effective nuclear charge is equal to the group number.

6. State the horizontal trend for atomic size based on effective nuclear charge.

7. Summarize the atomic size trend on the PT below. Draw arrows in the direction of **increasing** size.



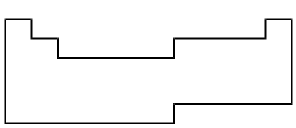
Trend #2: Ionic Radius

1. Draw the B-R diagram for potassium (K). Draw the B-R diagram for the potassium ion (K^+). Which is bigger? Explain.

2. Draw the B-R diagram for oxygen (O). Draw the B-R diagram for the oxygen ion (O^{2-}). Which is bigger? Explain.

3. Now compare the size of the K^+ & Ca^{2+} ion. Also compare the size of the N^{3-} & O^{2-} ion.

4. Draw the arrows for ionic radius. (NOTE: with ionic radius we can only compare metal ions to metal ions & non-metal ions to non-metal ions)



Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Trend #3 Ionization Energy (IE)

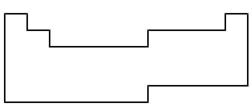
1. For each ion, list the number of electrons gained or lost to be isoelectronic (same number of electrons) with a noble gas.

Group1	Group 2	Group 6	Group 7	Group 8
Li^{+} ______ e- lost/gained	Be^{2+}	O^{2-}	F^{-}	Ne
Na^{+}				
K^{+}				

2. Ionization Energy (IE) is defined as the energy required to remove an electron. It is measured in joules (J).

3. As you go down a group the atomic size ___________________. If the distance from the nucleus to the valence electrons increases, what happens to ionization energy? Explain.

4. Metal atoms (on the left side of the PT) tend to lose electrons, non-metals (on the right side) tend to gain electrons. Using the definition of ionization energy, predict the horizontal & vertical trend on the periodic table.

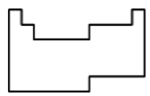
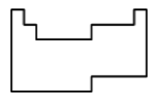
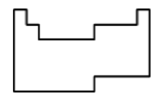
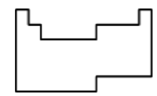
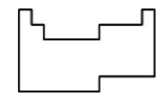


Trend #4 Electron Affinity (EA)

EA is the *opposite* of IE. Replace the words *required* & *remove* from the IE definition in step 2 & fill in the EA definition below.

Electron affinity is the energy _________________ when an atom ________________ an electron.

5. Is the trend for electron affinity the same or different from the trend for ionization energy? Explain.

Trend	Atomic Radius	Ionic Radius	Ionization Energy	Electron Affinity	Electronegativity
Description/ Description					
Horizontal & Vertical Trend Directions					
Horizontal Trend Explanation					
Vertical Trend Explanation					
Significance (Why It is Important)					

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Electronegativity Differences, Types of Bonding & Types of Compounds

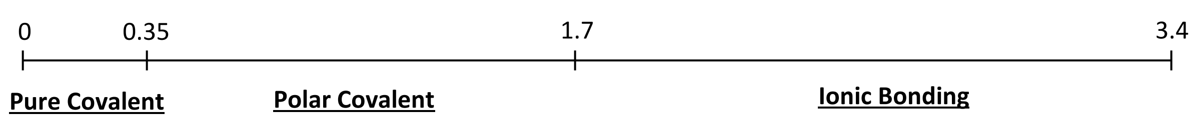
Atoms form bonds with other atoms to become more stable. They can transfer or share electrons to become ________________ with the nearest noble gas.

1. What are the two types of chemical bonds called? Which one involves sharing electrons & which involves a transfer of electrons?

2. In grade 10 you learned some simple guidelines to determine if a bond was ionic or covalent. In grade 11 we use the electronegativity scale to more accurately determine the bond type. In your own words define electronegativity.

3. Calculating the electronegativity difference between two atoms helps us classify the bond type. If two atoms have a large electronegativity difference (ΔEN), will they form an ionic or covalent bond? Explain briefly.

<u>Electronegativity Differences (ΔEN):</u>



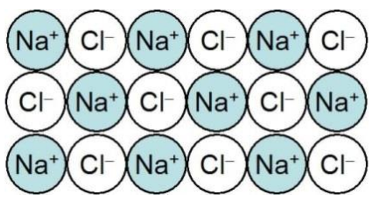
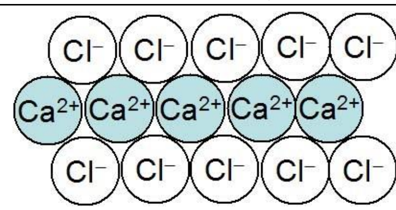
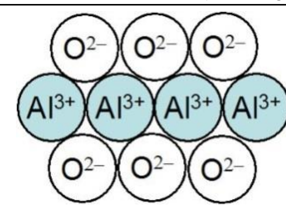
Properties of Ionic & Covalent/Molecular Compounds:

Property	**Ionic Compound**	**Covalent/Molecular Compound**
Examples		
State at SATP		
Melting Point		
Electrical Conductivity as a liquid		
Solubility in water		
Conducts electricity when dissolved in water		

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Chemical Formulas & Names of Ionic Compounds

The diagrams below represent some ionic compounds at the atomic level. Redraw the diagram on the white board to properly represent the relative size of the ions.

<table>
<tr><td>sodium chloride, $NaC\ell$</td><td>calcium chloride, $CaC\ell_2$</td><td>aluminum oxide, $A\ell_2O_3$</td></tr>
<tr><td>
</td><td>
</td><td>
</td></tr>
</table>

Key Questions

1. What are the names of the elements found in the compounds above?

2. How does the name of the elements in each compound differ from that of the free elements?

3. How many sodium ions are there in the sodium chloride sample shown above? How many chloride ions? What is the ratio between the two?

4. How many calcium ions are there in the calcium chloride sample shown above? How many chloride ions? What is the ratio between the two?

5. How many aluminum ions are there in the aluminum oxide sample shown above? How many oxide ions? What is the ratio between the two?

6. What is the relationship between the chemical formula for the compounds above & the ratio of the ions in them?

7. What is the charge of a sodium ion? What is the charge of a chloride ion?

8. What is the charge of the calcium ion? What is the charge of the chloride ion?

9. What is the charge of the aluminum ion? What is the charge of the oxide ion?

10. All samples of sodium chloride have a ratio of one sodium ion for one chloride ion. What must be true of the total (net) charge for any sample of sodium chloride?

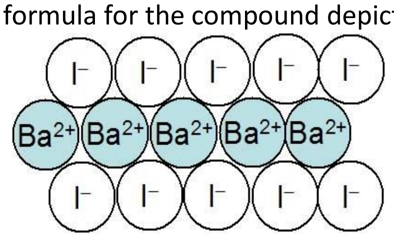
11. All samples of calcium chloride have a ratio of one calcium ion for two chloride ions. What must be true of the total (net) charge for any sample of calcium chloride?

12. All samples of aluminum oxide have an atomic ratio of two aluminums for three oxide ions. What must be true of the total (net) charge for any sample of aluminum oxide?

13. From the pattern seen in the last three questions, what is the rule for the total charge for a compound?

Exercises

1. Write the name & the chemical formula for the compound depicted below.



Model 2: Ionic Charges

Many ions have the same charge whenever they are found in a compound. Some of these ions are listed in the table below.

Group	**1**	**2**	**3**	**15**	**16**	**17**
Valence	1+	2+	3+	3-	2-	1-
	lithium Li^{+}			nitride N^{3-}	oxide O^{2-}	fluoride F^{-}
	sodium Na^{+}	magnesium Mg^{2+}	aluminum $A\ell^{3+}$	phosphide P^{3-}	sulphide S^{2-}	chloride $C\ell^{-}$
	potassium K^{+}	calcium Ca^{2+}			selenide Se^{2-}	bromide Br^{-}
	rubidium Rb^{+}	strontium Sr^{2+}				iodide I^{-}
	cesium Cs^{+}	barium Ba^{2+}				

What patterns do you notice about the charges of the ions with respect to their positions in the periodic table (or their Group number in the periodic table)?

2. Following the rule you established in the last key question, write correct chemical formula for each of the following compounds.

Compound	Formula
lithium chloride	
magnesium iodide	
strontium selenide	
rubidium fluoride	
lithium oxide	
sodium sulphide	
potassium chloride	
calcium phosphide	
barium oxide	
aluminum sulphide	

3. Use your answers to the Key Questions & the Exercise Questions to draw a conclusion about the ratio of ions in two compounds if the elements in the compounds are from the same groups (example: aluminum oxide & aluminum sulphide; lithium chloride & potassium chloride).

Drawing Lewis dot diagrams

Lewis diagrams illustrate only the valence electrons around an element. They are useful for ionic bonding.

1. Draw the following Lewis dot diagrams:
   a. H   b. F   c. Ar

2. Lewis diagrams for negative ions include a square bracket & the charge.
   a. S^{2-} b. N^{3-}

3. Draw the Lewis dot diagrams for magnesium & bromine atoms. Use the diagrams to explain the bonding between the atoms. List electrons lost & electrons gained. Calculate the net charge.

4. Draw the Lewis diagrams for aluminum & fluorine atoms. Use the diagrams to explain the bonding between the atoms. List electrons lost & electrons gained. Calculate the net charge.

Drawing Structural Diagrams for Molecules

Molecules are a diverse category of chemical compounds. The shape of the molecule affects its physical properties. Below is the method we use to draw a 2-D diagram of molecules.

1. Calculate the total number of valence electrons.
2. Arrange the atoms around the central atom.
3. Draw a line to represent single bonds between all atoms & the central atom.
4. Distribute electrons in pairs to atoms surrounding the central atom(s) to fill their valence shells.
5. Distribute the remaining electrons in pairs around the central atom(s).
6. When all the electrons are distributed, count those around the central atom. If there are more or less than 8 electrons, check the following:
    a. A central atom from group IIA or IIIA will have only 4 or 6 electrons respectively. A central atom from group VA, VIA, VIIA or VIIIA may have more than 10 electrons, but it must be an even number.
    b. If the central atom has fewer than 8 electrons & no exception applies, a multiple bond is needed.
7. If there is more than one way to rearrange the electrons for the multiple bond, you may have a resonance structure.
8. For polyatomic ions, square brackets & the charge must be added for the formula to be complete.

a. SO_3

b.) $C\ell_4$

c.) CO_2

d.) CO_3^{2-}

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

The Shape of molecules

Covalently bonded molecules have a variety of physical properties. Some are liquid or solid or gas at room temperature. The shape of the molecule has a big impact on their physical & chemical properties, so we need a way to predict, name & draw shapes.

Be proud to be Canadian – VSEPR theory was largely created at McMaster University in Hamilton, Ontario. It stands for **V**alence **S**hell **E**lectron **P**air **R**epulsion. The basic idea is that pairs of electrons in bonding or lone pairs have the same charge & repel each other. The molecule takes on a shape so that electron pairs are as far away from each other as possible.

VSEPR theory also has its own notation AB_xE_y

(A = central atom, B = outer atoms, E = lone pairs on central atom)

Each group needs 5 styrofoam balls (taken from failed grade 12 projects) & four toothpicks. The balls represent atoms (1 central & 4 outer atoms) & the toothpicks represent electron pairs.

1. Fill in the table with the correct VSEPR notation.
2. Build the molecule using the Styrofoam & toothpicks (arrange each pair of electrons as far apart as possible).
3. Draw the 3D shape using the convention explained in class.
4. Complete the shape name & try to think of an example molecule:

# of Outer Atoms	# of LP	VSEPR Notation	3-D Shape	Shape Name	Example
2	0	AB_2			
3	0				
		AB_2E_1			
1	2				

4	0				
		AB_3E_1			
		AB_2E_2			
		AB_1E_3			

Note: VSEPR only applies to molecules, NOT ionic compounds.

There are many more shapes! You'll learn several more in Chem 12.

Questions:

1. List three VSEPR notations for the shape "linear".

2. Go back to your table & label all the bond angles in each 3-D shape.

3. Is the VSEPR notation $A_0B_2E_1$ possible? Explain

4. Think of at least three molecules with an angular shape.

5. What shape (or shapes) is produced in molecules with group 4 elements as the central atom with no LP?

6. How do chemists simplify the notation AB_1E_1?

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Polar Molecules

Some molecules have a clear + & – charged ends. You can think of them as magnets with a north & south pole. Molecule polarity has a big impact on physical properties & water solubility.

1. For the water molecule:
   a. Draw the structural diagram for H_2O
   b. Draw the 3D shape (based on VSEPR)
   c. Indicate polar covalent bonds with the δ^- & δ^+ symbols

2. Imagine that a water molecule was placed in a magnetic field. Which (A, B or C) best represents the orientation of the water molecule. Explain your choice.

3. Copy down the simplified rule for determining if a molecule is polar or non-polar:

NOTE: The simplified rule assumes all the outer atoms bonded to the central atom are the same.

4. Draw the structural diagram & 3D shape of the following molecules. Label any polar covalent bonds (with δ^- & δ^+). Identify if it's a polar or non-polar molecule. Explain why.
   a. CF_4 b. NH_3 c. CO_2

Properties of Polar & Nonpolar Molecules:

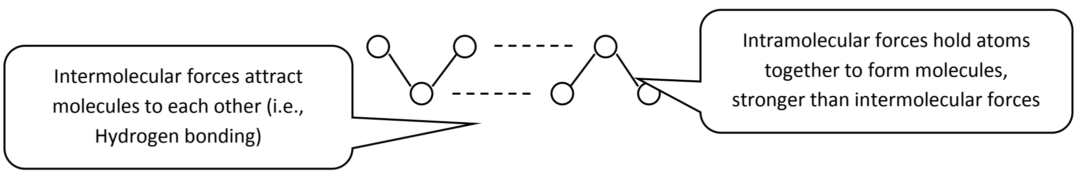
Polar Molecules:

- are water soluble
- have high melting & boiling points because they stick together like magnets

Nonpolar Molecules:

- do not dissolve in water
- have low melting & boiling points because they have little or few intermolecular forces

5. Intermolecular Forces – forces holding two molecules together. Intermolecular forces are much weaker than intramolecular bonds (ie. ionic & covalent bonds)



6. <u>Types of Intermolecular Forces:</u>

Type of Force	Characteristics
London Force	
Dipole-dipole Force	
Hydrogen Bonding (H-bonding)	

7. Indicate the type of intermolecular force between the following molecules:

a. CO

c. OF_2

e. MgH_2

b. H_2O

d. $SiBr_4$

Chemical Nomenclature

Trivial/Common Name	IUPAC Name
Quicklime	calcium oxide
Laughing gas	dinitrogen monoxide
Saltpetre	sodium nitrate
Potash	potassium carbonate
Muriatic acid	hydrochloric acid
Rust	iron (III) oxide
Baking soda	sodium hydrogen carbonate (sodium bicarbonate)
Cream of tartar	potassium hydrogen tartrate

Valence Numbers & Bonding Capacity:

Group #	I	II	III	IV	V	VI	VII	VIII
Valence	1+	2+	3+	4	3-	2-	1-	0

- Groups I → III form cations while Groups V → VII form anions

Multiple Valences:

- some elements (most are transition metals) have more than 1 valence
- common ones include:

Element	Cu	Hg	Au	Fe	Co	Sn	Pb	Sb	P
Latin Root Name	cupr-	mercur-	aur-	ferr-	cobalt-	stann-	plumb-	stibn-	phosphor-
Lower Valence	1+	1+	1+	2+	2+	2+	2+	3+	+3
Higher Valence	2+	2+	3+	3+	3+	4+	4+	5+	+5

Polyatomic Ions:

- some atoms travel together
- atoms are covalently bonded inside the polyatomic ion
- common ones include (there are many more!):

Polyatomic Ion	BrO_3^-	$C\ell O_3^-$	MnO_3^-	NO_3^-	HCO_3^-	CO_3^{2-}	CrO_4^{2-}	$Cr_2O_7^{2-}$
Valence	1-	1-	1-	1-	1-	2-	2-	2-
Name of Ion	bromate	chlorate	manganate	nitrate	bicarbonate	carbonate	chromate	dichromate

Polyatomic Ion	SO_4^{2-}	PO_4^{3-}	NH_4^+	OH^-	CN^-	CH_3COO^-	IO_3^-
Valence	2-	3-	1+	1-	1-	1-	1-
Name of Ion	sulphate	phosphate	ammonium	hydroxide	cyanide	acetate	iodate

Writing Chemical Formulas Using Valences:

- criss-cross rule—valence of one element becomes the subscript of the other

Chemical Nomenclature Rules:

1. Binary Compounds: (2 elements only)
   a. Metal-Non-metal Compounds:
      - (metal name) + (non-metal root) + "ide"
      - e.g., $BaC\ell_2$ → barium chloride & $FeC\ell_2$ → iron (II) chloride

b. Non-metal-Non-metal Compounds:
   - use Greek prefixes: **mono, di, tri, tetra, penta, hexa, hepta, octa, nona, deca**
   - (prefix & name of 1st non-metal) + (prefix & root of 2nd non-metal) + "ide"
   - e.g., $NO_2 \rightarrow$ (mono)nitrogen dioxide

2. Latin Names:
   - used for multivalent elements only
   - higher valence $\rightarrow$ "ic", lower valence $\rightarrow$ "ous"
   - (latin root of metal) + "ic/ous" + (root of 2nd element) + "ide"
   - e.g., $CuC\ell \rightarrow$ cuprous chloride & $CuC\ell_2 \rightarrow$ cupric chloride

3. Polyatomic Compounds:
   - many ways to modify a polyatomic
   - (name of 1st element) + (name of ion)
   - e.g., $Na_3PO_3 \rightarrow$ sodium phosphite

	Special Notes	**Name of Ion**	**New Ion**
add 1 oxygen	Valence stays the same	per___ate	SO_5^{2-}
parent polyatomic	Valence stays the same	___ate	SO_4^{2-}
subtract 1 oxygen	Valence stays the same	___ite	SO_3^{2-}
subtract 2 oxygens	Valence stays the same	hypo___ite	SO_2^{2-}
subtract 1 oxygen, add 1 sulphur	Valence stays the same	thio___ate	$S_2O_3^{2-}$
add 1 hydrogen	Add +1 to the valence	hydrogen ___ate	HSO_4^-

4. Hydrates:
   - molecules that have water attached to them
   - (name of anhydrous salt) + prefix + "hydrate"
   - e.g., $CuSO_4 \cdot 5H_2O \rightarrow$ copper (II) sulphate pentahydrate

5. Compounds Containing Hydrogen:
   a. Hydrides:
      - valence of H is 1- (H is anion)
      - (name of 1st element) + "hydride"
      - e.g., $LiH \rightarrow$ lithium hydride

   b. Acids:
      i. Binary Acids:
         - H is the cation (charge of 1+)
         - "hydro" + (root of element) + "ic acid"
         - e.g., $HC\ell \rightarrow$ hydrochloric acid

      ii. Oxyacids
         - H + polyatomic ion
         - (root name of polyatomic ion) + ate becomes `` ic acid``
           ite becomes ``ous acid"
         - e.g., $H_2SO_5 \rightarrow$ persulphuric acid & $H_2SO_2 \rightarrow$ hyposulphurous acid

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Nomenclature (Naming chemicals)

1. With your group write down the word for “water” in as many languages as you know. Put your answers on the white board

2. With your group, think of two reasons why “water” is not the official name for the chemical H_2O. Summarize your answer on the white board.

The International Union of Pure & Applied Chemists (IUPAC) have created a naming system that does two things:

i.

ii.

Note: the naming rules are adjusted every few years to make them clearer.

3. How can we tell if a compound is ionic?

4. Are all the bonds in an ionic compound ionic (ΔEN>1.7)?

5. Look at your naming sheet & summarize the rule for **naming binary ionic compounds** below:

6. Use your naming rule to name the following. Show your teacher your answers.

Formula	Compound Name
$MgC\ell_2$	
BaI_2	
ZnH_2	
AgBr	
SrO	
BeF_2	

7. Where are the transition elements found on the periodic table?

8. What makes transition metals different from group 1 & group 2 metals. List at least two differences.

9. What are the two ways to name compounds with a transition metal? Name CuF_2 using two different ways.

NOTE: we always use the roman numeral system, not the Latin!

10. Using your naming sheet as a reference, name the following & show your teacher.

Formula	Compound Name
CuI_2	
Sn_3N_2	
$SnC\ell_4$	
FeO	

11. What is wrong with the following names? Explain what's wrong & write the correct name:

Incorrect Name	Error	Possible Correct Name(s)
copper chloride		
tin (III) sulphide		
magnesium hydrate		
beryllium (II) iodide		
lithium carbonide		

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Nomenclature Part Two

Polyatomic ions

1. Some ions travel in groups. Consider the polyatomic ion PO_4^{3-}
   a. What does the 3- mean?

   b. What are the intramolecular bonds in PO_4^{3-}?

   c. Write & name the compound formed between Na^+ & PO_4^{3-}

2. What's the difference between sodium phosphate & sodium phosphide?

3. The PO_4^{3-} form of the ion is called the parent. There exist alternate versions of the ions. Using your table, name these ions:
   a. PO_5^{3-} b. PO_3^{3-}

4. Explain what the prefix "per" mean? What does the prefix "hypo" mean?

5. You need to memorize the polyatomic ions on your sheet (this is a short list – there are hundreds!) Look at the following memory aid:

   "Nick the camel ate a clam supper in Phoenix"

   a. The word "Phoenix" helps us remember the charge & number of oxygen atoms in the phosphate ion. How does it work? (hint: look at the number of vowels & consonents)

   b. Now explain the rest of the memory aid.

*****The best way to remember the polyatomic ions is to do lots of naming practise******

6. Write the chemical formula:

Compound Name	Formula
copper (I) sulphate	
tin (II) carbonate	
mercury (II) hydroxide	
magnesium bromate	
mercury (II) perchlorate	
manganese (II) sulphite	

7. Some acids contain polyatomic ions. These acids are called oxyacids. Why?

8. Read your naming sheet & summarize the rule for naming oxyacids in your own words.

9. Use your rule to name the following:

Formula	Compound Name
$HNO_{3\,(aq)}$	
$H_3PO_{3\,(aq)}$	
$HBrO_{3\,(aq)}$	
$H_2Cr_2O_{7\,(aq)}$	

10. The last category of acids are called binary acids. Read the naming rules & write the formula for hydrosulphuric acid.

11. Compare the formula for hydrosulphuric acid with sulphuric acid.

12. Note: The IUPAC name for acids is changing. Currently $HNO_{3\,(aq)}$ is called aqueous hydrogen nitrate. Give two reasons why this name is superior to nitric acid.

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Nomenclature Part Three

Naming bases - In grade 11 the bases we use are those containing OH^- (plus the base NH_3)

1. What is the OH^- ion called?

2. Name the following:
   a. $Mg(OH)_2$
   b. $LiOH$

3. $Mg(OH)_2$ is referred to as dibasic, but $LiOH$ is not.
   a. Explain what dibasic means.
   b. What kind of base is $A\ell(OH)_3$?

Covalent compounds (molecules)

4. What are the first 10 prefixes used in naming molecules?

5. Is the prefix "mono" used before the name of the first element in a compound? What's the rule for using "mono"?

6. What is the IUPAC name for the H_2O molecule?

7. Peroxides are a special class of molecules which contain a special type of covalent bond. If H_2O_2 is named hydrogen peroxide, name the following:
   a. Li_2O_2
   b. Na_2O_2

8. Answer questions #35-50 from "Chemical Nomenclature Review". Try to remember the polyatomic ions first, then check your page to make sure you got them correct.

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

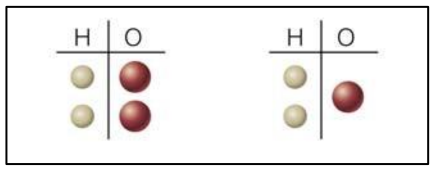
Balancing Chemical Equations

Atoms are neither created nor destroyed in a chemical reaction, they are just rearranged. In other words, in a chemical reaction, what goes into the reaction must come out of the reaction. Using this knowledge & some bookkeeping skills, all unbalanced chemical equations can be balanced.

Model 1

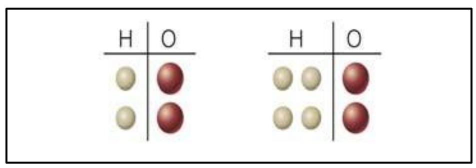
The following figures show the combination of hydrogen & oxygen to produce water. Illustrations from: http://wps.prenhall.com/wps/media/objects/439/449969/Media_Portfolio/ch10.html

Figure 1



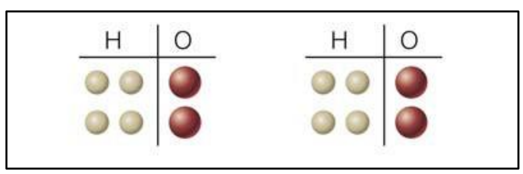
1 molecule of hydrogen + 1 molecule of oxygen → 1 molecule of water

Figure 2



2 molecules of reactants → 2 molecules of product

Figure 3



4 H atoms in reactants → 4 H atoms in products 2 O atoms in reactants → 2 O atoms in products

Key Questions

1. In Figure 1 there is one molecule of H_2 & one molecule of O_2 on the left side of the equation & one molecule of H_2O on the right. Even though there is 1 of everything, why is this reaction not balanced?

2. In Figure 2 there are two molecules on the left & two molecules on the right. Even though there are 2 on the left & 2 on the right, why is this reaction not balanced?

3. In Figure 3, how many reactant molecules & product molecules are shown in the model?

4. Does Figure 3 represent a balanced equation? Explain your answer.

5. What condition must be met in order for there to be a balance between reactants & products?

6. Write the balanced equation to show the reaction between hydrogen gas & oxygen gas to form water. (Hint: look at the model for guidance.)

7. Use the model kit & build the chemicals in your equation. **Show your teacher**.

8. Identify whether the following is a balanced chemical equation. Explain why or why not. If not, write the balanced equation.

$$H_2O_2 \rightarrow H_2O + O_2$$

9. If mercury (Hg) & oxygen (O_2) were reacted to form mercury (II) oxide (HgO), how many molecules of each reactant & product would be needed to balance the equation?

10. Write a balanced reaction equation for hydrogen reacting with nitrogen to produce ammonia. Build the compounds using the model kit & show your teacher. Name each chemical

11. Using the smallest whole number coefficients, balance the following reactions. Write the name of each chemical.

    a. ____ HgO → ____ Hg + ____ O_2

b. ____ Fe + ____ $O_2 \rightarrow$ ____ Fe_2O_3

c. ____ $KC\ell O_3 \rightarrow$ ____ $KC\ell$ + ____ O_2

d. ____ $Ca(OH)_2$ + ____ $H_2SO_4 \rightarrow$ ____ HOH + ____ $CaSO_4$

e. ____ Cu + ____ $AgNO_3 \rightarrow$ ____ $Cu(NO_3)_2$ + ____ Ag

f. ____ C_2H_6 + ____ $O_2 \rightarrow$ ____ CO_2 + ____ H_2O

Problems

Write the formulas for the components in each reaction and, using the smallest whole number coefficients, balance each equation.

1. Zinc metal reacts with hydrochloric acid to produce hydrogen gas & aqueous zinc chloride.

2. Solid carbon reacts with oxygen gas to produce carbon dioxide gas.

3. Solid sodium chloride is broken down into its elements.

Written by Bryan Horan
Edited by Linda Padwa & David Hanson, Stony Brook University

Chemical Reactions

We'll investigate a series of reactions & predict the products. Your teacher will perform the demos behind the blast shield. Follow all safety instructions.

1. What's the difference between a chemical reaction & a physical change?

2. What can we observe that indicates a chemical reaction has occurred? List 5.

Part A - $Cu_{(s)}$ & $AgNO_{3\ (aq)}$

1. Write the IUPAC name for $AgNO_3$(aq) ____

2. Write the predicted chemical reaction. Include states.

3. Sketch the beaker & wire before & after reacting. Label your diagram.

4. Classify this reaction. Explain.

Part B – Lithium solid in water.

1. List the physical & chemical properties of lithium.

2. Why can't your teacher hold the lithium in their hand?

3. Predict & balance the reaction between $Li_{(s)}$ & $H_2O_{(\ell)}$:

4. What evidence are you expecting to observe that a chemical reaction took place?

5. Explain the colour change of the water in the beaker.

6. Classify the reaction. Explain.

Part C – Sodium metal in water.

1. Will sodium produce a more or less vigorous reaction? Why?

Part D – Calcium metal in water

1. Write the balanced chemical reaction between $Ca_{(s)}$ & $H_2O_{(\ell)}$. What must you check before writing the equation?

2. How can we determine experimentally if the reaction produces H_2 gas?

Part E – Magnesium metal & oxygen.

1. Predict the reaction between $Mg_{(s)}$ & $O_{2(g)}$

2. Based on the colour of the flame, how much energy does this reaction release?

3. Classify the reaction. Explain.

*****Be sure to take observations of the Cu & $AgNO_3$ before the class ends! ****************

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Shall we dance? **Classifying types of chemical reactions**.

Complete the following table using what you remember of chemical reactions from grade 10. Complete the table in pencil in case you make a mistake. Use chemical symbols A, B, C, D, etc for the example reaction. As you complete the activity you may add or change the information in your table.

Type of reaction	Short form	Explanation	Example (including state) (s, ℓ, g, aq)
		Elements or less complex compounds come together to form a single, more complex compound.	
		a compound breaks apart into either elements or less complex compounds	
		A single element replaces another one in a compound	
		Ions in a compound switch places with ions in another compound to form two new compounds	

The dance…

Adam and Barbara were both single. No one was talking about "Adam and Barbara" being together before the dance. They both go to the dance alone.

However, they meet at just the perfect time when a song they both adore is playing. They end up holding hands the entire dance. After that fateful meeting no one ever sees Adam without Barbara, they are forever referred to as "Adam and Barbara".

1. Represent the drama of Adam and Barbara as a chemical equation? Use A to represent Adam and B to represent Barbara.

2. If A and B represent elements can you describe what is happening?

3. How would you classify A and B?

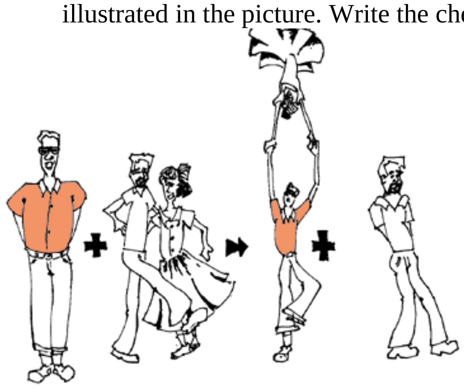
The dance continues…Later that same evening Xavier and Yasmine, who have been ‘the couple’ forever, have a heated quarrel and break up.

4. Represent the drama of Xavier and Yasmine as a chemical equation? Use X to represent Xavier and Y to represent Yasmine.

5. If X and Y represent elements can you describe what is happening?

6. How would you classify X and Y using the words from the first page of this activity?

7. What type of reaction is represented in the picture below? Write your own analogy for the reaction illustrated in the picture. Write the chemical reaction.



The dance continues…

In their blissful state, Adam and Barbara (AB) try to help Xavier and Yasmine (XY) reconcile their differences. After everyone agrees to stop quarreling, Adam asks Yasmine to dance. Xavier and Barbara decided that they will dance together as well.

8. Represent Adam and Barbara's attempt to reconcile Xavier and Yasmine's differences as a chemical equation.

9. How would you summarize this reaction?

10. What type of reaction does this scenario represent?

Model 2: Types of Reactions (aq = aqueous, g = gas, s = solid, l = liquid)

Classification (Type) of Reaction	**Example Reactions**
S	$2H_2 (g) + O_2 (g) \rightarrow 2H_2O(\ell)$
D	$2H_2O (\ell) \rightarrow 2H_2(g) + O_2(g)$
SD	$2A\ell(s) + 3Cu(NO_3)_2(aq) \rightarrow 2A\ell(NO_3)_3(aq) + 3Cu(s)$
DD	$Pb(NO_3)_2 (aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$

1. As shown in the model, list the number of reactants and the number of products found in the synthesis reaction?

2. As shown in the model, list the number of reactants and the number of products found in a decomposition reaction?

3. How would you compare a single displacement reaction to a double displacement reaction?

4. Identify the type of reaction shown in each of the following chemical equations:

- $KC\ell O_3(s) \longrightarrow 2KC\ell(s) + 3O_2(g)$ _____
- b. $AgNO_3(aq) + KC\ell(aq) \longrightarrow AgC\ell(s) + KNO_3(aq)$ _____
- c. $2 H_2(g) + O_2(g) \longrightarrow 2 H_2O(g)$ _____
- d. $2 AgNO_3(aq) + Cu(s) \longrightarrow Cu(NO_3)_2(aq) + 2 Ag(s)$ _____
- e. $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$ _____
- f. $NaI(aq) + C\ell_2(g) \longrightarrow 2NaC\ell(aq) + I_2(s)$ _____

5. For the reaction between zinc metal and hydrochloric acid (aqueous) producing zinc chloride (aqueous) and hydrogen gas

- write an equation for this reaction

- balance the equation from Part a

- classify the reaction

6. For the reaction between aqueous calcium nitrate and aqueous sodium hydroxide producing the precipitate (solid) calcium hydroxide and aqueous sodium nitrate

- write a balanced equation for this reaction

- classify the reaction and explain how you came to this conclusion

7. Nitrogen molecules and hydrogen molecules react to form ammonia gas.

- write a balanced equation for this reaction

- classify the reaction and explain how you came to this conclusion

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Classifying Reactions (part 2):

- Several types, the most basic of which include: (C)combustion, (S) synthesis, (D) decomposition, (SD) single displacement & (DD) double displacement

Combustion Reactions:

- Adding oxygen & heat
- 3 types

i. Combustion of Nonmetals:
- Produces non-metal oxides
- Example: $S_{8\,(s)} + 8\,O_{2\,(g)} \rightarrow 8\,SO_{2\,(g)}$
- Note: non-metal oxide in water forms acid solution: $SO_2 + H_2O \rightarrow H_2SO_3$

ii. Combustion of Metals:
- Produces metal oxides
- Example: $4\,Fe_{(s)} + 3\,O_{2\,(g)} \rightarrow 2\,Fe_2O_{3\,(s)}$
- Note: metal oxide in water forms a base: $Fe_2O_3 + H_2O \rightarrow 2\,Fe(OH)_3$

iii. Combustion of Hydrocarbons (compounds made of C and H):

(a) Complete Combustion:
- $CO_{2\,(g)}$ & $H_2O_{(l)}$ are always the products
- Example: $C_6H_{12}O_{6\,(s)} + 6\,O_{2\,(g)} \rightarrow 6\,CO_{2\,(g)} + 6\,H_2O_{(l)}$

(b) Incomplete Combustion:
- $C_{(s)}$, $CO_{(g)}$, $CO_{2\,(g)}$, $H_2O_{(l)}$ are possible products
- Hard to predict products

1. Give an example of a reaction that is both a synthesis and a combustion reaction.

2. How can you tell if a reaction is combustion?

3. a.) Is the following (unbalanced) reaction a combustion? Explain

$$NaHCO_3 + \text{heat} \rightarrow Na + H_2 + C + O_2$$

b.) How could this reaction be classified?

c.) Balance the reaction

Manager: ___________________ Speaker: __________________ Recorder: ________________ 1 2 3 4 5

Single Displacement Reactions

The general form for a single displacement reaction can be written as:

$$AB + C \rightarrow CB + A$$

1. Classify A, B, and C as metals, or non-metals explain.

4. What state (s, ℓ, g or aq) are the chemicals A, B and C

Another single displacement could also be written as:

$$XY + Z \rightarrow XZ + Y$$

5. Classify X, Y and Z as metals, or non-metals explain.

6. Label the states of chemicals X, Y and Z

7. Based on your answers to #1 and #2, which kind of elements can be displaced in a single displacement reaction?

   How do we know if a reaction will take place? We need the metal activity series (which was found by doing a series of reactions of metals and ionic solutions). The metal activity series will be provided on tests and quizzes.

Li	Linda
K	Please
Na	Send
Ca	Charlie
Mg	McCarthy
Al	A
Zn	Zebra
Fe	In
Sn	The
Pb	Large
H	Heavy
Cu	Crate
Hg	Marked
Ag	Striped
Pt	Perishable
Au	Goods

<u>Halogen Activity Series:</u>

F > Cℓ > Br > I

<u>Note:</u>

- When comparing any 2 elements (metal or halogen), the higher (& more reactive) element will want to be part of a molecule

8. What do you notice about the metals at the bottom of the activity series? Explain.

9. Why do we need a halogen activity series as well?

10. What pattern do you notice in the halogen activity series?

Practise problems. If no reaction occurs write NR

1. $Li\ (s) + H_2O\ (\ell) \rightarrow$
2. $A\ell\ (s) + FeO\ (s) \rightarrow$
3. $Cu\ (s) + H_2O\ (\ell) \rightarrow$
4. $Mg\ (s) + CaCO_3\ (s) \rightarrow$
5. $Ag\ (s) + HC\ell\ (aq) \rightarrow$
6. $Fe\ (s) + HC\ell O_3\ (aq) \rightarrow$
7. $Na\ (s) + H_2O\ (\ell) \rightarrow$
8. $Zn\ (s) + PbI_4\ (s) \rightarrow$
9. $Pt\ (s) + H_2SO_4\ (aq) \rightarrow$
10. $Ni\ (s) + MgCO_3\ (s) \rightarrow$
11. $Ba\ (s) + H_2O\ (\ell) \rightarrow$
12. $Sn\ (s) + HgS\ (s) \rightarrow$
13. $A\ell\ (s) + HNO_3\ (aq) \rightarrow$
14. $Ca\ (s) + FeC\ell_2\ (aq) \rightarrow$
15. $Pb\ (s) + H_2O\ (\ell) \rightarrow$
16. $Ni\ (s) + HC\ell\ (aq) \rightarrow$
17. $A\ell\ (s) + H_3PO_4\ (aq) \rightarrow$
18. $Ca\ (s) + H_2O\ (\ell) \rightarrow$
19. $Zn\ (s) + Pb(NO_3)_2\ (aq) \rightarrow$
20. $Ag\ (s) + CuSO_4\ (aq) \rightarrow$

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Solubility & Net Ionic Equations (NIE)

The following table (called a Solubility Table) was created experimentally (empirically) by placing ionic compounds in water (at SATP conditions):

	Anions									
Cations		$C\ell^-$, Br^-, I^-	S^{2-}	OH^-	SO_4^{2-}	CO_3^{2-}, PO_4^{3-}, SO_3^{2-}	CH_3COO^-	NO_3^-	$C\ell O_3^-$	O^{2-}
	High Solubility (>0.1 mol/L at SATP)	most	Group 1, NH_4^+, Group 2	Group 1, NH_4^+, Sr^{2+}, Ba^{2+}, $T\ell^+$	most	Group 1, NH_4^+	most	all	Most	Group 1, NH_4^+, Ba^{2+}
	Low Solubility (<0.1 mol/L at SATP)	Ag^+, Pb^{2+}, $T\ell^+$, Hg_2^{2+}, (Hg^+), Cu^+	most	most	Ag^+, Pb^{2+}, Ca^{2+}, Ba^{2+}, Sr^{2+}, Ra^{2+}	most	Ag^+	none	Ca^{2+}	most

All Group 1 compounds, including acids & all ammonium compounds are assumed to have high solubility in water

1. Is calcium carbonate low or high solubility?

2. Is silver bromide low or high solubility?

3. Is iron (III) sulphate low or high solubility?

4. For what cations (positively charged) & anions (negatively charged) are the compounds always high solubility in water?

5. For anions:
   a. For what anions are most of the compounds usually highly soluble?

   b. For the anions that are usually soluble, list the cations that form low solubility compounds.

6. For what anions are compounds usually low solubility?

7. What patterns can be found in your answers for questions 4 & 5. Consider the location of the elements in the periodic table.

Double displacement reactions (DD):

8. How can we tell if a chemical reaction occurred? (list 5 things)

In a DD reaction one of the products must be a low solubility ionic compound (or liquid or gas)
If both products are soluble, then no reaction will occur.

9. Use your solubility table to determine if this reaction will occur:

   Potassium iodide solution is combined with copper (II) nitrate solution

10. Lead (II) nitrate solution is mixed with potassium iodide.
    a. Write the balanced chemical reaction. Include states.
    b. Draw a beaker with all the ions labelled BEFORE the reaction occurs.
    c. Draw a beaker with the chemicals AFTER the reaction occurs.

Ions that do not take part in the chemical reaction are called "Spectator ions"

   d. Cancel out any ions present in the beaker before & after the reaction. List all of these spectator ions.

e. Write the chemical reaction for the ions that actually reacted. This is called the Net Ionic Equation (NIE)

11. For the reaction between silver nitrate & potassium chloride solution:
    a. Write the balanced chemical reaction including state (phase labels).
    b. Write all the ions on the reactant & product sides (called the Ionic Equation or IE).
    c. Cancel any spectator ions & write the Net Ionic Equation (NIE).

12. Now consider the reaction between an acid (eg. $HC\ell$) and a base (eg. NaOH)
    a.) Write the balanced chemical reaction including state
    b.) Is a precipitate formed? Did a chemical reaction occur? Explain.
    c.) Write the net ionic equation (NIE) for ANY acid base neutralization.

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Unit 3: Quantities in Chemical Reactions: Mole Concept

1. We commonly use counters to simplify expressing quantities of objects. Express the following quantities another way:

2 dozen = _________________ 3 pairs = __________________ 1 gross = __________________

2. What about chemistry? Estimate with your group how many atoms of carbon are in 12 g of carbon solid. Think about the order of magnitude (100s, 1000s, etc)

3. There is a convenient counter used in chemistry. It is called the mole (or mol for short) where 1 mol = 6.02×10^{23} objects. This number is referred to as Avogadro's constant, named after the famous Italian scientist (it is written as $N_A = 6.02 \times 10^{23}$objects/mol). Write Avogadro's constant with all zeroes.

4. 1 mol is a huge number. Let's try to imagine just how big. Consider the following:

You just won the Chemistry Lottery! The prize is $6.02 x 1023. Since you are so rich you can spend $1 billion every second. How long until you run out of money?

   a. Discuss the problem with your group & estimate how long it will take to spend the money. (First decide if it will take days, years, etc). Write down your estimate.

   b. In a well-organized solution (given information, calculations with headings, therefore statement) calculate how long it will take to spend the money.

   c. Compare your result to the estimate you made in part (a).

5. In previous units we calculated the average atomic mass (AAM). We didn't really discuss the units but we will now. All masses in the periodic table are given in units of **g/mol** & we use the symbol **M** to represent molar mass. For example, use your periodic table & write the molar mass for the following elements:

   M_{Cu} = M_{Li}=

6. Which element on the periodic table has the lowest mass? Which has the highest?

7. We can also calculate the molar mass of compounds. Simply add the molar masses of all the elements in the compound. Find the molar mass of CO_2. Use the symbol M_{CO2}.

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Mole Calculations

Now that we understand what a mole is, we can start doing some calculations. Up to this point we have looked at chemical reactions but haven't calculated quantities of reactants & products.

1. Converting Between Mass & Moles:

$$m = nM$$

(m = mass (in grams), n = moles (in mol), M = molar mass (in g/mol)

   a. Find the number of moles of 40.0 grams of H_2O. (2.22 mol)

   b. Find the mass of 2.80 moles of CH_3OH. (89.7 g)

2. Converting to Number of Entities
   (HINT: when doing mole calculations you can always replace the word "mol" with the number 6.02×10^{23})

   a. How many molecules of H_2 are there in 1.50 moles of H_2? (9.03×10^{23})

   b. How many H atoms are there in 1.5 moles of H_2? (1.8×10^{24} atoms)

c. How many moles of $Zn(NO_3)_2$ can be produced if 4.46 x 10^{24} oxygen atoms are available?
(1.23 moles)

Advanced Mole Calculations

3. Mass to Molecules:
   How many molecules of carbon dioxide gas, CO_2, are in 50.0 g of gas? (6.84 x 10^{23} molecules)

4. Molecules to Mass:
   Suppose you had 2.50 x 10^{39} oxygen atoms. What mass of sulphuric acid, H_2SO_4 can be produced?
   (1.02 x 10^{17} g)

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Percent Composition

Congratulations! On your last chemistry test you answered 38 of 40 questions correctly. You achieved a grade of 95%! The percentage grade on your exam indicates the part of the exam you answered correctly. The part of the exam that you answered incorrectly is only 5%, or 2 questions.

Percentage is also a useful tool in chemistry. The percent composition by mass of a compound represents the percent that each element in a compound contributes to the total mass of the compound.

Using the students present in class today, complete the data table in Model 1

Item	Total # of Students	Total # of Girls	Total # of Boys	Percent Girls	Percent Boys
Students					

1. Describe how one could calculate the percent of boys present today using the data in the chart.

2. What is the percent of boys present today, based on the data in the chart?

When you chew a piece of gum, mass is lost as the sweetener dissolves. After the sweetness is gone, only the 'gum' remains. Given a piece of gum to chew, each student will collect data to complete the data table in Model 2.

As a group, propose & describe below an experimental procedure that will permit you to collect the data needed to determine the percent of sweetener in the gum. Do not proceed with the experiment until your teacher has approved your procedure.

Item	Total Mass	Mass of sweetener	Mass of gum	Percent sweetener	Percent gum
Chewing Gum					

<u>Key Questions</u>

1. Show the set-up used to determine the percent by mass of sweetener in the chewing gum.

2. What is the percent of sweetener in the chewing gum?

3. At the end of this experiment the gum is wet with saliva. How does the presence of saliva influence the percentage reported in Question 2?

4. What method could be used to eliminate the error introduced by the presence of the saliva? How does this affect the results?

Item	Total Mass	Mass of Carbon	Mass of Hydrogen	Percent Carbon	Percent Hydrogen
CH_4 (methane)	16.0 g	12.0 g	4.0 g	75.0%	25.0%

5. What is the mass of carbon in 16.0 g of methane?

6. Determine the percent by mass of carbon in methane .

7. What information do you need in order to determine the percent composition by mass of sodium in sodium chloride?

<u>Exercises</u>

1. Write the general mathematical formula that can be used to calculate the percent composition by mass of any substance?

2. Use the atomic masses of the elements to determine the percent composition by mass of sodium & chlorine in sodium chloride.

3. Use the atomic masses of the elements to find the percent composition of hydrogen in H_2O.

4. Determine the percent composition of oxygen in potassium chlorate, $KC\ell O_3$

5. Determine the percent composition of phosphorus in calcium phosphate, $Ca_3(PO_4)_2$.

6. Copper (II) sulphate pentahydrate, with the formula $CuSO_4 \cdot 5H_2O$, has a deep blue colour. When heated to remove the water, the crystals crumble & turn white. What is the percentage, by mass, of water in copper (II) sulphate pentahydrate crystals?

Problems

1. The percent of oxygen in a colourless liquid is determined to be 94.1%. Is this liquid water or hydrogen peroxide (H_2O_2)?

2. What is the percent composition of a compound that contains 8.1 grams of nickel & 2.2 grams of oxygen in a 10.3 gram sample?

3. Nicotine, the addictive drug in cigarettes, contains 74.0% carbon, 8.6% hydrogen, & 17.3% nitrogen. What mass of each element can be recovered from a 55.0- gram sample of nicotine?

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Simplest Formulas

Simplest Formula (SF) (also called empirical formula):

- SF is the lowest whole number ratio of all atoms in a compound
- Example: HO is the simplest formula of the molecule H_2O_2 (hydrogen peroxide)

Example:

What is the simplest formula for the following?

$C_6H_4Cl_2$ _________ H_2O _________ N_2O_4 _________ P_2O_3 _________

Note:

- SF is NOT always a proper chemical formula
- many different chemical compounds have the same simplest formula but are different compounds, each with different physical & chemical properties:

Molecular Formula	Name of Compound	Simplest Formula
$C_6H_{12}O_6$	glucose	CH_2O
$C_5H_{10}O_5$	ribose	CH_2O
$C_4H_8O_4$	erythrose	CH_2O
$C_3H_6O_3$	lactic acid	CH_2O
$C_2H_4O_2$	acetic acid	CH_2O
CH_2O	formaldehyde	CH_2O

Finding Simplest Formula:

Use the following chart:

Element	% Composition	Molar Mass	# Moles	Molar Ratio	Whole # Ratio

Example #1:

A compound consists of 79.85% carbon & the remainder is hydrogen. What is the simplest formula of the compound?

Element	% Composition	Molar Mass	# Moles	Molar Ratio	Whole # Ratio

Therefore the simplest formula of the compound is ________________.

Example #2:

25 grams of a binary acid contains 24.30 grams of chlorine. What is the simplest formula of the compound?

Element	% Composition	Molar Mass	# Moles	Molar Ratio	Whole # Ratio

Therefore the simplest formula of the compound is ________________.

Example #3:

A 4.000 g sample contains 3.748 g of carbon & the remainder is hydrogen. Determine the simplest formula for this compound.

Element	% Composition	Molar Mass	# Moles	Molar Ratio	Whole # Ratio

Therefore the simplest formula of the compound is ________________.

What is the rule when determining the whole # ratio?

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Molecular Formulas & Hydrates

Molecular Formula (MF):

MF is the actual formula of a compound.

Molecular Formula	MFW (g/mol)	SFW (g/mol)	$\frac{MFW}{SFW}$	Simplest Formula
$C_6H_{12}O_6$	180.18	30.03	6	$C_6H_{12}O_6 = 6(CH_2O)$
$C_5H_{10}O_5$	150.15	30.03	5	$C_5H_{10}O_5 = 5(CH_2O)$
$C_4H_8O_4$	120.12	30.03	4	$C_4H_8O_4 = 4(CH_2O)$
$C_3H_6O_3$	90.09	30.03	3	$C_3H_6O_3 = 3(CH_2O)$
$C_2H_4O_2$	60.06	30.03	2	$C_2H_4O_2 = 2(CH_2O)$
CH_2O	30.03	30.03	1	$CH_2O = 1(CH_2O)$

Finding Molecular Formula:

$$MF = \frac{MFW}{SFW}(SF)$$

MF = molecular formula
MFW = molar mass of molecular formula
SFW = molar mass of simplest formula
SF = simplest formula

Example 1: Finding the Molecular Formula of a Compound

A compound contains 54.52% carbon, 9.17% hydrogen & the remainder is oxygen. If the molecular mass of the compound is 88.11 g/mol, what is the molecular formula?

Element	% Composition	Molar Mass	# Moles	Molar Ratio	Whole # Ratio

SF = MF =

SFW =

MFW =

Example 2: Finding the Molecular Formula of a Compound

Analysis of a sample of oxalic acid shows it is 26.68% carbon, 71.08% oxygen, & the rest is hydrogen. In addition, 3.50 mol of oxalic acid has a mass of 316 g. What is the molecular formula of oxalic acid?

Hydrates:

- an ionic compound that has a specific number of water molecules associated with it
- of the form: ionic compound · x H_2O
- the ionic compound is often referred to as the "anhydrous salt"
- finding the anhydrous salt to water ratio is easy to find because the only thing that is needed is the simplest formula chart

Example: Finding the Molecular Formula of a Hydrate

4.52 grams of a lithium chloride hydrate was heated & 2.84 grams of water was released. What is the molecular formula of the hydrate?

Compound	**% Composition**	**Molar Mass**	**# Moles**	**Molar Ratio**	**Whole # Ratio**

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Stoichiometry

Let's investigate this concept with a car analogy.

1. Write the balanced chemical reaction for producing one car from tires & window wipers:

2. What is the ratio for tires:wipers:cars?

3. Give two ways the production of cars could stop.

4. Why would car manufacturer need to know the formula for producing cars?

5. Make up your own analogy like the car analogy. Be prepared to share it with the class.

*In chemistry we use **Stoichiometry** to calculate quantities of reactants & products in a chemical reaction. We need the mole concept & the calculations we've been practising to solve larger problems.*

6. Balance the following reaction:

$$\_\_\_ N_2 + \_\_\_ H_2 \rightarrow \_\_\_ NH_3$$

7. What do the coefficients represent? What are their units?

8. A student was overheard saying the following:

   "This reaction is not balanced because there are 4 molecules total on the reactant side & only 2 molecules on the product side."

   Do you agree or disagree with this student?

9. Is it possible to balance the reaction like this? Why or why not?

$$\tfrac{1}{2} N_2 + {}^{3}/_{2} H_2 \rightarrow NH_3$$

10. Let's look at another chemical reaction & practise good problem solving technique:

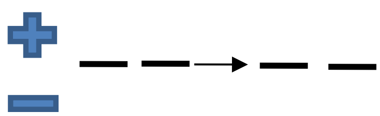
Example: Mass-Mass Stoichiometry Calculations

Methane gas (CH_4) is added to 7.00 g of oxygen gas in a combustion reaction. What mass of carbon dioxide will be produced?

11. Write the balanced chemical reaction. Include phase labels.

12. Which chemicals will decrease in amount & which will increase as the reaction progresses?

13. Complete the following bar chart to illustrate your answer to question #12. Include the molar ratio of increase & decrease.



14. We want to visualize what is happening in our chemical reaction. Which chemicals must hit each other at the same time for a reaction to occur? Include amounts of each. Draw a diagram illustrating the collision.

15. What if only one molecule of O_2 collides with one CH_4 molecule? Explain.

When the correct number of reactant molecules collide with enough energy & correct orientation, an ***effective collision*** *occurs.*

16. Now it's time to solve the Methane Gas problem (from question #10). We will use the same format every time we do stoichiometric calculations: (Hint: write the information in table form).
    a. write the balanced reaction
    b. write any given masses (m) under the chemical formula
    c. calculate the molar mass (M) for oxygen & carbon dioxide
    d. calculate the number of moles (n) of oxygen gas

balanced reaction:

mass (m):

molar mass (M):

number of moles (n):

17. Now we're ready to use stoichiometry to find the number of moles of $CO_{2\,(g)}$. For every two moles of O_2 gas consumed, how many moles of CO_2 are produced?

18. Use the following relationship to find moles of CO_2 produced:

$$\frac{coefficient\ of\ O_2}{coefficient\ of\ CO_2} = \frac{n_{O_2}}{n_{CO_2}}$$

Cross multiply to find the number of moles of CO_2. Put your answer in the table.

19. What's the last step? Do your final calculation & write your therefore statement. Be sure to include units & consider significant digits (sig digs).

Every time we do stoichiometry, set up a table & perform any calculations under the table.
You don't need to show simple calculations like finding molar mass.

20. Can we use mass in stoichiometry instead of moles? Check your answer to the above problem.

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Stoichiometric calculations

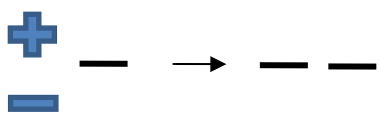
Use this format for calculations on quizzes & tests

Example 1: Mole to Mass Problem:

What mass of potassium is produced when 0.15 mol of potassium oxide decomposes?

A. Write the balanced chemical equation. Include phase labels.

B. Complete the bar chart to indicate which chemical increases & decreases in amount & by how much.



C. Draw a diagram to show the reactant molecules colliding. Describe the reaction in words.

D. Solve the problem using the table format. Include units & a final therefore statement with correct sig. digs.

balanced reaction:

mass (m):

molar mass (M):

number of moles (n):

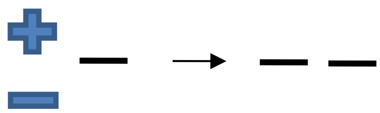
E. Discuss your final answer & explain if it is reasonable.

Example 2: Mass to Mass Problem:

An automotive airbag is inflated with nitrogen produced from the decomposition of sodium azide, $NaN_{3\,(s)}$ into $Na_{(s)}$ & $N_{2\,(g)}$. If the mass of the gas inside a fully inflated airbag is 87.5 g, what mass of sodium azide is needed?

A. Write the balanced chemical equation. Include phase labels.

B. Complete the bar chart to indicate which chemical increases & decreases in amount & by how much.



C. Draw a diagram to show the reactant molecules colliding. Describe the reaction in words.

D. Solve the problem using the table format. Include units & a final therefore statement with correct sig. digs.

balanced reaction:

mass (m):

molar mass (M):

number of moles (n):

E. Discuss your final answer & explain if it is reasonable.

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Limiting Reagent Stoichiometry Calculations

Sometimes we run out of one of the reactants in a chemical reaction. This reactant is called the **Limiting reagent** (LR) because it limits the amount of products that can be produced.

Example: Real-Life Scenario

1. In an assembly line, 4 tires & 2 wiper blades are needed to equip a car. Suppose in the warehouse 140 tires & 74 wiper blades are available. How many cars can be equipped?

2. Are the tires or the wiper blades the limiting reagent? Explain.

Example: LR Stoichiometry Calculation:

Complete **Section D** of a full solution problem for the following:

2.00 grams of methane gas (CH_4) is added to 7.00 grams of oxygen gas for a combustion reaction. What mass of carbon dioxide will be produced?

balanced reaction:

mass (m):

molar mass (M):

number of moles (n):

Summary: Steps to Solve a LR Stoichiometry Problem

- write out the balanced chemical equation
- write out the given information
- convert all given information to moles (may require finding the molar mass first)
- determine which of the reactants is the LR (find ratio)
- use mole ratio with the LR to find the amount of moles of the compound needed
- convert moles to the desired units

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% Yield & % Purity

No chemical reaction is perfect. We can predict the amount of product in a perfect world using stoichiometry, but always some chemical product is lost in the laboratory.

List the three main ways chemical product is lost in a reaction:

Chemists always strive for no waste in a chemical reaction. To analyse a reaction they calculate the % yield:

$$\%Yield = \frac{actual yield}{theoretical yield} x100\%$$

Definitions:
Actual Yield = the amount measured during an experiment
Theoretical Yield = the amount calculated using stoichiometry

Example: Percentage Yield Calculation:
In an experiment, suppose 10.0 g of potassium iodide was used with excess lead (II) nitrate & 9.50 g of precipitate was collected. Calculate the percentage yield of the reaction.

Solution:

$$2\,KI \quad + \quad Pb(NO_3)_2 \quad \rightarrow \quad 2\,KNO_3 \quad + \quad PbI_2$$

m

M

n

Percentage Purity:

When purchasing chemicals, the price depends on how pure the chemical is. These are referred to as reagent grade (most pure), laboratory grade, & technical grade (least pure). Technical grade chemicals are advertised to be 90% pure. What makes up the other 10%? Nobody knows (not even the chemical supply company). We hope that the 10% is chemically inert & won't interfere with our reaction.

1. Would a percent purity <100% increase or decrease our predicted yield?

2. How can we factor the percent purity into our calculations to get a more realistic percentage yield?

Example: Calculations with Percentage Purity

In the previous example, 10.0 g of potassium iodide was reacted with lead (II) nitrate. What if the potassium iodide was only 85.0% pure? How much precipitate would we expect then?

$$2\ KI \quad + \quad Pb(NO_3)_2 \quad \rightarrow \quad 2\ KNO_3 \quad + \quad PbI_2$$

m

M

n

6. Compare the mass of precipitate in the Percentage Purity example to the Percentage Yield example.

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Solution Terminology & Properties of Water

In this unit, we will look at the chemistry of aqueous solutions, usually ionic compounds & covalent molecules dissolved in water.

solubility	solute	miscible
saturated solution	dilute solution	immiscible
electrolyte	concentrated solution	alloy
solution	aqueous solution	unsaturated solution
solvent	precipitate	insoluble

1. Draw a picture of solid NaCℓ being added to a beaker of water. Use the word bank to describe the water, the NaCℓ & what is formed when the two are added.

2. We keep adding more NaCℓ to the beaker. Use the word bank to describe this solution compared to the one created in step 1.

3. Eventually we add so much NaCℓ that undissolved salt settles to the bottom of the container. Which word describes this solution? How does it differ from a solution that can dissolve more salt?

4. When have we used the term *precipitate* before in the course?

5. Is the solid salt settling to the bottom of the beaker in step 3 a precipitate? Explain.

6. Will $CuCℓ_{(s)}$ dissolve in water? How do you know?

7. Which word describes a solid that cannot dissolve? __________________________

8. Which word describes a liquid that does not dissolve in water? Give an example.

9. Are all solutions liquids?

10. What is an aqueous solution that conducts electricity called? _______________________

11. The properties of Water video: List the main properties of water explained in the video.

12. Why is water called the "Universal solvent"?

13. Distinguish between cohesion & adhesion.

14. Why can water absorb a large amount of heat?

15. Explain surface tension.

16. Why is ice less dense than liquid water? Sketch the water molecules in a liquid compared to ice.

17. Which property of water explains the following:
    a. water erodes a mountain. _________________________
    b. water spreads through a paper towel. _________________________
    c. a rock skims across the surface of a lake. _________________________
    d. water heaters are one of the biggest energy uses in the home. _________________________

Manager: _______________ Speaker: _______________ Recorder: _______________ 0 1 2 3 4 5

Methods of Dissolving

We want to model & understand how compounds dissolve in water. There are three methods:

- ionic dissociation
- simple dissolving
- covalent ionization

Ionic Dissociation:

1. What is another name for an ionic compound? ___________________

2. How can we tell if a salt is highly soluble?

3. Draw the 3-D diagram for a water molecule. Label the positive & negative end.

4. Let's look at what happens when NaCℓ is dissolved in water:

Equation (with states)	
Diagram	

5. Will $NaC\ell_{(aq)}$ conduct electricity? Explain.

Simple Dissolving:

There is no solubility table for molecules. The only rule we have is "**Like dissolves Like**"

Polar solutes dissolve in polar solvents & *Non-polar solutes dissolve in non-polar solvents*

6. What is the simplified rule you learned earlier in the course for determining if a molecule is polar or non-polar?

7. Water is a polar solvent. Which of the following will dissolve in water?
   a. $NC\ell_3$ b. CH_4 c. CO_2 d. SO_2

8. Will molecules (covalently bonded compounds) dissociate into ions like salts do? What is holding atoms together in a molecule?

9. Let's look at what happens when sugar ($C_{12}H_{22}O_{11}$) is put in water:

Equation (with states)	
Diagram	

10. Will a solution of NF_3 conduct electricity in water? Explain.

Covalent Ionization

Molecules react with solvent to form ions. Applies to all acids & ammonia

11. Use electronegativity values to determine if $HC\ell$ is ionic or covalent.

12. Write the chemical reaction between hydrochloric acid & water.

13. Will the solution in #12 conduct electricity? Explain.

14. Write the chemical reaction between ammonia (NH_3) & water that produces OH^- ions.

15. Summary – with your group & a whiteboard create a flowchart to determine the method of dissolving for any compound.

16. Use your flowchart to determine the process by which the following dissolve:
    a. K_2O b. $MgC\ell_2$ c. HNO_3 d. CH_3OH

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Concentration Formulas

As a chemist you will need to communicate with many audiences: other scientists, politicians, engineers, the general public, etc.

1. If chemists are talking to other chemists, which unit of concentration will they prefer?

2. Why wouldn't the average person understand the unit of mol/L?

3. How is the alcohol content of wine expressed? _______________
   Volume by volume percent:

   v/v% = _________________________________ x 100

4. 4.5mL of acetic acid is dissolved in 85.5 mL of water.
   a. Find the v/v%.

   b. What common mistake do some students make when solving question #1?

   c. We can measure the amount of solute & solution using mass or volume or both. This leads to other % concentration units:
      i. mass by volume percent:
      ii. mass by mass percent:

5. 3.57mL of HCℓ is dissolved in 72.0g of water.
   a. Find the m/m%.

b. What information is missing in the above question? How do you know?

c. Solve the problem:

d. Would the v/v% = m/m%. Explain

6. Environmental chemists need to explain the risk even small amounts of chemicals can pose to the general public. What unit do atmospheric chemists use when discussing global warming & CO_2 concentration? Why do they use this unit?

   ppm = ppb=

7. Find the ppm & ppb of 0.250 mg of NO_3 in 125 mL of solution.

8. Summarize all the ways chemists calculate concentration.

9. Since we can report the same concentration multiple ways, we need to be able to convert between units: Find the m/v% of 2.05 mol of NaOH when it is dissolved in 1 L of solution.

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Molarity

Observe the two 1000mL volumetric flasks, each filled with a different concentration of $CuSO_{4\ (aq)}$. With your group, discuss the concept of concentration using the terms dilute & concentrated as they relate to the two solutions.

1. Which solution has the darker colour?

2. What might be responsible for the darker colour in one of the solutions?

3. Based on your observations of the solutions, which solution probably contains more solute per unit volume? Explain your answer.

4. What is meant by the term *concentration*?

5. What factors should be considered when determining the concentration of a solution?

6. When a solution is diluted, additional solvent is added to the more concentrated solution. How does the amount of *solute* in the initial solution compare to the amount of solute in the final solution?

The most common measure of concentration used by chemists is molarity (M).

Molarity is defined as the number of moles of solute (mol) divided by the total volume (V) of the solution in liters (L).

concentration = moles of solute per liter of solution (mol/L or M)

$$C = n / V$$

Molarity also is called molar concentration. When the symbol *M* is accompanied by a numerical value, it is read as "molar." For example, a solution labeled 3.0 *M* $NaC\ell$ is read as "three molar sodium chloride solution".

Use [ ] to represent concentration. For example, $[NaC\ell]$ = 0.11 mol/L

Exercises

1. In a problem a student is given the amount of solute in grams & the volume of the solution in milliliters. What must be done with the information before the molarity can be calculated?

2. Calculate the molarity of a solution in which 0.50 moles of $MgC\ell_2$ are dissolved to produce 1.5 liters of solution.

3. Intravenous (IV) saline solutions are often administered to patients in the hospital. Normal saline solution contains 0.90g $NaC\ell$ in exactly 100 ml of solution. Calculate the molarity of this solution.

4. Calculate the molarity of the following solutions.
   a. 1.0 moles of $NaNO_3$ in 500 ml of solution.

   b. 85.0 g of $NaNO_3$ in 250 ml of solution.

   c. Which of the solutions, 4a or 4b, is more concentrated? Explain your answer.

Making solutions of a fixed concentration is just as important as being able to calculate it. Carefully read the following two methods & answer the questions below.

Method A:	Method B:
When preparing one litre of a 1.0 molar solution one should pour some solvent (water) into a 1.0 litre volumetric flask. The measured amount of solute (1.0 mol) is added to the volumetric flask. The flask is stirred to dissolve the solute, & then additional solvent is added to bring the volume to the 1.0 litre mark.	Fill a 1.0 litre volumetric flask with water up to the 1.0 litre mark & then add measured amount of solute (1.0 mol).

1. In Method A, why is the solute added to some of the solvent & dissolved before more solvent is added to bring the volume to the 1.0 litre mark on the volumetric flask?

2. Which of the solutions contains one litre of solvent? Explain.

3. Which of the solutions contains a 1.0 *M* solution? Explain your answer.

Problems

1. What volume of 0.25 *M* solution can be prepared using 0.50 mol of $KC\ell$?

2. What volume of 0.10 *M* solution can be prepared using 11.6 g of $NaC\ell$?

Extension Activity: Mini-lab

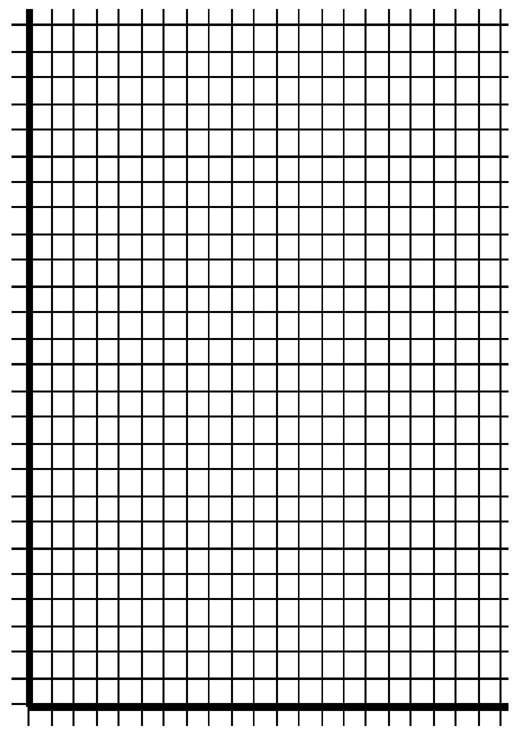
1. Determine the mass of an empty 100 mL volumetric flask.

2. Find the mass of the following sucrose solutions:

   0.0625 *M* 0.125 *M* 0.250 *M* 0.500 *M*

3. Calculate the density of each solution.

4. Prepare a graph of density versus concentration for sucrose. Include the density of pure water on the graph, using the value 1.00 g/mL.



5. Questions:
   a. What is the relationship (direct, inverse, etc.) between concentration & density? Explain your answer.

   b. Based on your graph, what is approxima

   c. What is the minimum density of any aq

Revised by: Lorraine Giloni, Kenneth Levy & Kelly Levy
Edited by Linda Padwa & David Hanson, Stony Brook University

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Solubility curves

Let's review the solution terminology we've learned so far:

1. Which term applies to a sponge that is dry?
____

2. Which term applies to a sponge that is soaked?
____

3. Can you add more water to a sponge that is soaked? ______________

4. Explain how a sponge is analogous to an aqueous solution.

5. List all the units used to report concentration.

There is another unit used by chemists. **A chemical's solubility is given in units of g/100mL**.

6. (a.) Explain the numerator and denominator of the solubility unit.

b.) How else could the amount of water be expressed?

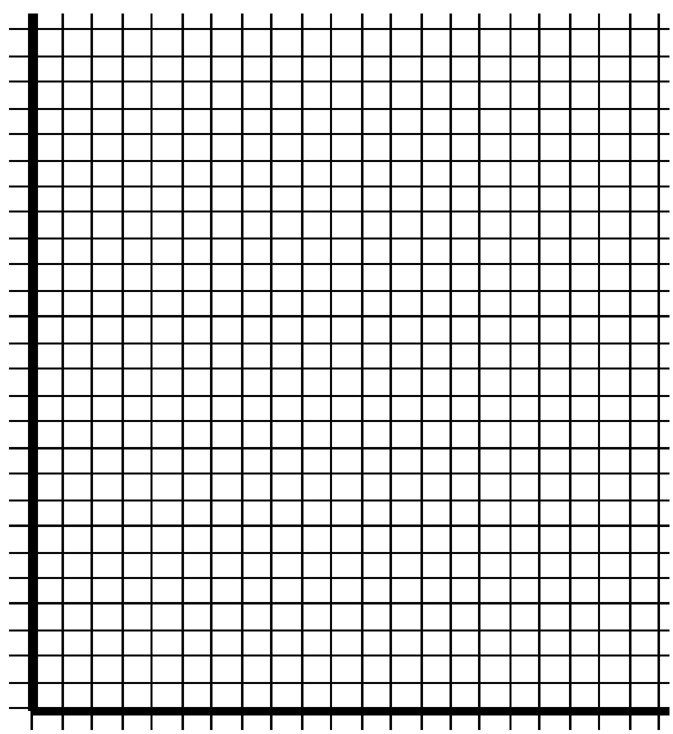
7. If you read "The solubility of $NaNO_2$ is 88g/100mL at 20°C" what kind of solution (unsaturated, saturated, supersaturated) is being discussed?

8. Why is it important to include the temperature in when discussing solubility? If the temperature is not stated, what do we assume the temperature to be?

9. Using the grid, make a graph of the solubility data. Label the axis.

Table 1: Solubility data

Temperature (°C)	Solute (g/100mL)
10	33
30	42
50	52
70	62
90	73



10. What is the relationship between temperature and solubility for this solute?

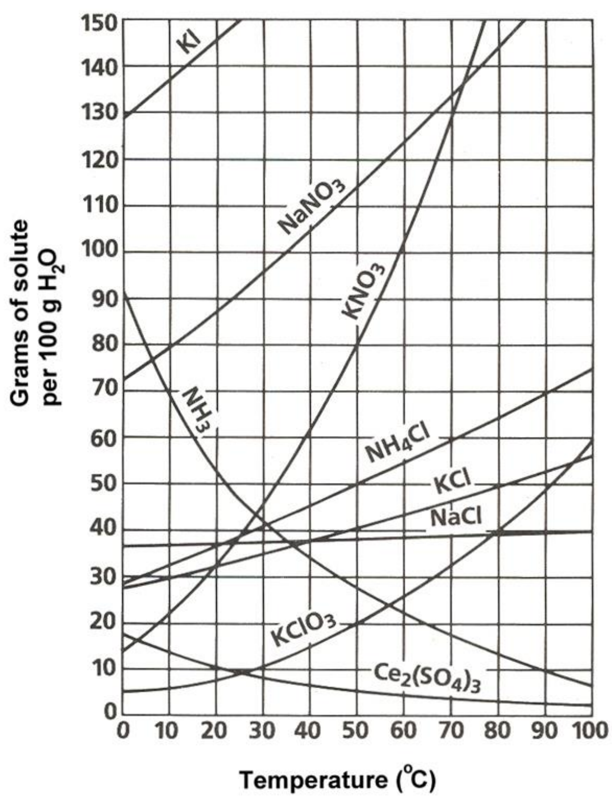
11. What will happen if 12 g of this solute is added to 100g of water at 20°C?

12. What kind of solution is formed?

13. At 20°C, 50 g of this solute is added to 100g of water. What will happen to the extra solute?

14. What kind of solution is formed?

15. What kind of solution is formed when 50g of solute is added to 250mL of water at 30°C?



16. Compare this curve to the one you constructed. What is the solute in your graph?

17. Which substance is most soluble at 60°C?

18. Which substance would you say is most soluble overall? Explain.

19. Consider the state of the chemicals. What is generally true for solubility of solids and gasses as temperature increases?

20. Explain using a diagram why gasses have lower solubility at higher temperature. Can you think of a real world example?

21. Using diagrams, explain why solid solutes have higher solubility at higher temperature. Do your rough copy on the whiteboard, then put your good copy on this sheet.

Problems

22. Omar likes sweet coffee, but he notices that iced coffee is never as sweet as hot coffee. Explain why.

23. A driveway is 550cm but 305 cm. If there is 5 cm of snow on the driveway, what is the maximum amount of road salt that can be dissolved by the water from the snow? (use the graph and state any simplifying assumptions made)

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Solution Stoichiometry

We need to make calculations for amounts of reactants & products in aqueous solution.
Set up the following problem like we did in the last unit (let's focus on Part D of a full solution)

1. If 30.0 mL of 0.100 mol/L silver nitrate is mixed with 15.0 mL of 0.900 mol/L potassium sulphate, what mass of precipitate will form?

   Balanced equation:

2. Can we do % yield & % purity questions in solution stoichiometry? Explain.

3. What's the most common error students make in solution stoichiometry?

4. In a double displacement reaction, potassium iodide & lead (II) nitrate react to form lead (II) iodide & potassium nitrate. What minimum volume of 0.500 mol/L potassium iodide is necessary to produce 10.0 grams of precipitate?

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Introduction to Acids & Bases

What makes a solution acidic, basic or neutral? Also, we'll consider the concept of conjugate acids & bases.

1. Complete the table for the H^+ ion:

Subatomic Particle	Number of Particles
protons (p^+)	
electrons (e^-)	
neutrons (n^0)	

2. Give two correct names for the H^+ ion.

3. The hydrogen ion will immediately combine with water to form the hydronium ion. What is its chemical formula? Draw a diagram of its structure.

Following the H^+ is key to understanding acid base reactions. One chemical species loses a proton, & another gains a proton. This leads us to the definition of acids & bases.

<u>Arrhenius Theory of Acids & Bases:</u>

- Acids dissociate to form H^+ & an anion when dissolved in water
- Bases dissociate to form OH^- & a cation when dissolved in water

<u>Bronsted-Lowry Theory of Acids & Bases:</u>

- Acids are proton donors
- Bases are proton acceptors

4. Label the following as Arrhenius bases (A) or Bronsted-Lowry bases (B-L) or both:
   a. LiOH b. NH_3

5. Is it possible to have a reaction with an acid only & no base?

A **neutral solution** contains hydrogen ions & hydroxide ions in equal concentration. Let's take a closer look at some chemical equations.

i. $NaOH_{(s)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$
ii. $HC\ell_{(aq)} + H_2O_{(\ell)} \rightarrow H_3O^{+}_{(aq)} + C\ell^{-}_{(aq)}$
iii. $NH_{3\,(g)} + H_2O_{(\ell)} \rightarrow NH_4^{+}{}_{(aq)} + OH^{-}_{(aq)}$
iv. $H_2CO_{3\,(g)} + H_2O_{(\ell)} \rightarrow H_3O^{+}_{(aq)} + HCO_3^{-}{}_{(aq)}$
v. $HC\ell_{(aq)} + NH_{3\,(aq)} \rightarrow NH_4^{+}{}_{(aq)} + C\ell^{-}_{(aq)}$

Questions:

6. In equation (i), is NaOH an acid or base? Which kind? (Arrhenius? B-L?) Explain.

7. In equation (ii), is $HC\ell$ an acid or base? Explain.

8. In equation (iii), is NH_3 an acid or base? Explain

9. In equation (iii), is H_2O an acid or base? Explain.

10. In equation (iv), is H_2O an acid or base? Explain.

11. In equation (iv), is $H_2CO_{3\,(g)}$ an acid or base? Explain.

12. Compare the behaviour of NH_3 in equations (iii) & (v). Note any differences or similarities.

13. The prefix "amphi" means "both kinds". Which chemicals in the model are amphiprotic? Explain.

14. In the reaction below, identify which of the reactants is an acid (A) or a base (B)

$$HC_2H_3O_{2\,(aq)} + H_2O_{(\ell)} \rightarrow C_2H_3O_2^{-}{}_{(aq)} + H_3O^{+}_{(aq)}$$

   a. Now consider the products. Which product can now donate a proton? Label this chemical as the conjugate acid (CA).
   b. Which product can now accept a proton? Label this chemical as the conjugate base (CB).
   c. If a chemical is an acid on the reactant side is it a CA or CB on the product side? Explain.

   d. Link together the A with the CB & the B with the CA.

15. For the following reaction,

$$HCO_3^-{}_{(aq)} + OH^-{}_{(aq)} \rightarrow CO_3^{2-}{}_{(aq)} + H_2O_{(\ell)}$$

a. Label the A, B, CB & CA.
b. Identify any amphiprotic chemicals in the reaction.

c. Write the reaction for HCO_3^- when it acts as an acid & the reaction for when it acts as a base.

16. Is water an acid, base, neither or both? Explain.

17. Give an example of a base that is a B-L acid, but is not an Arrhenius acid.

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Acids and Bases Summary: Think of three properties of acids and bases

Acids	**Bases**
•	•
•	•
•	•

1. By what method are acids soluble in water?
2. Show the chemical reaction when $HC\ell$ and H_2SO_4 are added to water:

In general, acids follow the following format $HA \rightarrow H^+_{(aq)} + A^-_{(aq)}$

3. What does A represent?

4. What method do bases dissolve in water?
5. Write the equation for NaOH and $Mg(OH)_2$ in water

6. Do you remember the term dibasic? Which compound in #5 is dibasic, and why.

7. Write the general equation representing how bases dissolve in water. Explain the letters you use.

Acids can be separated into two categories: Strong acids and Weak acids:

Strong Acids:
- Dissociate completely in water (> 99%)
- Conducts electricity easily
- HNO_3, H_2SO_4, $HC\ell O_4$
- All binary acids except HF

Weak Acids:
- Dissociate incompletely in water (< 50%)
- Does not conduct electricity easily
- All organic acids (acids containing carbon)

9. Classify the following as weak acids (WA) or strong acid (SA)

a.) $H_2S_{(aq)}$ b.) $HF_{(aq)}$ c.) $HNO_{2(aq)}$ d.) $CH_3COOH_{(aq)}$

10. Explain the difference between a dilute strong acid and a concentrated weak acid:

Strong Bases:
- Dissociate completely in water (> 99%)
- Conducts electricity easily
- Includes all metal hydroxides

Weak Bases:
- Dissociate incompletely in water (< 50%)
- Does not conduct electricity easily
- Includes all organic bases

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Logarithms & the pH of a Solution

What are Logarithms?

- a logarithm is a pre-calculator system of mathematics that converts very large & very small numbers into values that are easier to work with
- the conversion is done by changing the value into an exponent of base 10

1. Using base 10, the logarithm (or log) of 10,000 is_________________________________

2. Using base 10, the logarithm (or log) of 0.001 ___________________________________

Example: Finding the logarithm of different values:

3. Using your calculator, find the logarithm (base 10) of the following values:

Value	Logarithm
10,000,000	
0.0001	
0.5	
64	

Depending on the calculator,
press LOG then the value,
OR,
press the value, then press the LOG button

Example: Finding the inverse logarithm of different values:

4. Using your calculator, find the inverse logarithm (base 10) of the following values:

Logarithm	Value
3	
-5	
0.5	
2.7	

Depending on the calculator,
press 2nd, then LOG then the value,
OR,
press the value, then press 2nd, then LOG button,
alternatively, use 10^x button

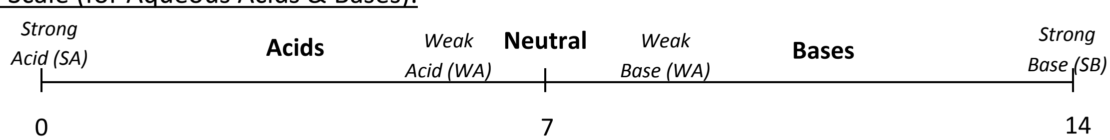
pH of a Solution

5. pH = ____

6. $[H_3O]^+ = [H]^+$, & can be obtained from the dissociation of acids

_______________________________________ ___________________________________

pH Scale (for Aqueous Acids & Bases):



7. If the pH is 3.5, what is $[H^+]$ of the solution?

8. If HNO_3 is 0.21 mol/L, what is the pH?

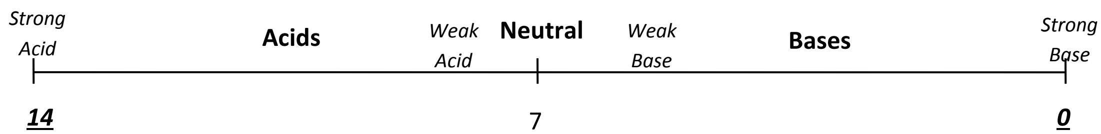
pOH of a Solution

pOH = Power of Hydroxide, that is, the logarithm of the hydroxide (or OH^-) concentration in a solution

9. $[OH^-]$ can be obtained from the dissociation of bases

______________________ ______________________

pOH Scale (for Aqueous Acids & Bases):



Notice:

$pH = -\log [H^+]$ ↔ $[H^+] = 10^{-pH}$

↕ (A) ↕ (B)

$pOH = -\log [OH^-]$ ↔ $[OH^-] = 10^{-pOH}$

10. What is the sum of the pH & pOH of a solution?

11. Can a solid have a pH value?

12. What is $[OH^-]$ of a basic solution whose pH is 2.8? If the base was NaOH, what is its concentration?

13. What is $[OH^-]$ of a basic solution whose pH is 2.8? If the base was $Mg(OH)_2$, what is its concentration?

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Acid/Base Neutralization

These calculations are an example of solution stoichiometry.

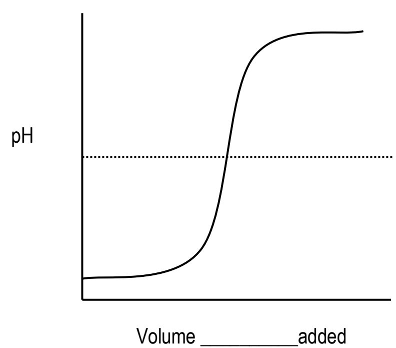
1. If an acid, HA, is reacted with a base, MOH, write the balanced chemical reaction. Label all chemicals.

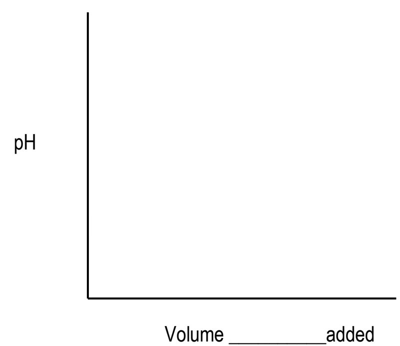
2. Classify the reaction as S, D, SD or DD. Explain.

3. What is true of the moles of acid & moles of base in a neutral solution?

When the moles of acid = moles of base, the titration has reached the <u>equivalence point.</u>

4. This graph is called a titration curve.
   a. Based on the curve, was base added to acid or acid to base? Explain.
   b. Sketch the graph if the order is reversed





5. Work through the following example of a complete neutralization. Use the technique from solution solubility:
   In a titration experiment, 30.0 mL of 0.100 mol/L magnesium hydroxide is titrated against 0.150 mol/L hydrochloric acid. What volume of hydrochloric acid is necessary for the titration?

6. In the previous example, are the moles acid = moles base?

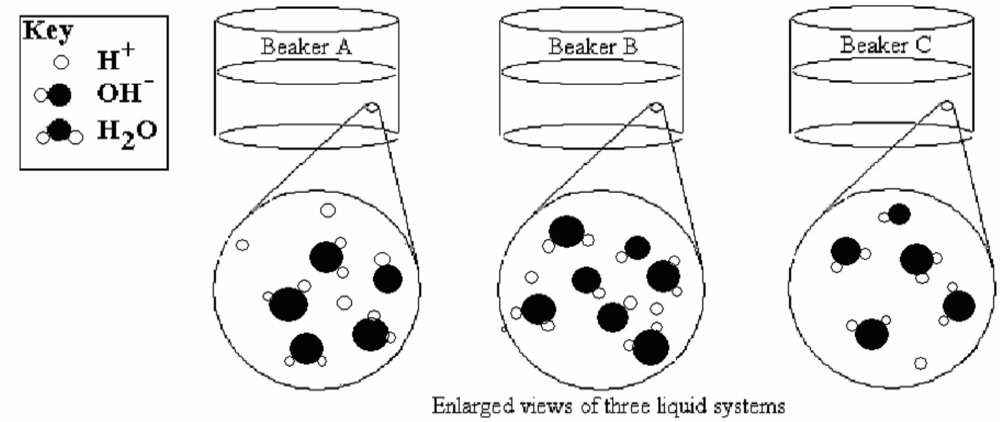
7. Are the moles of $Mg(OH)_2$ = moles of HCℓ? Explain.
8. Let's assume we're not at the equivalence point in a titration. What may be true about the amount of base added to the acid?

9. Define the terms: Incomplete neutralization vs overshooting the titration

10. In the following example, the solution is not at the equivalence point. Was the titration incomplete or overshot? (Hint: calculate the moles of acid & moles of base & see which is in excess)

Suppose 40.0 mL of 0.300 mol/L barium hydroxide was added to 90.0 mL of 0.200 mol/L hydrochloric acid. What is the pH of the resultant solution?

Model:



Note: spectator ions are not shown in this model, but they are present in each solution.

<u>Key Questions</u>

1. How does the concentration of H^+ compare to the concentration of OH^- in solution A?

2. How does the concentration of H^+ compare to the concentration of OH^- in solution B?

3. How does the concentration of H^+ compare to the concentration of OH^- in solution C?

4. Identify the acidic solution in the model.

5. Identify the basic solution in the model.

6. Identify the neutral solution in the model.

<u>Exercises:</u>

1. Based upon the information presented in the key of the Model, draw reactants & products that form when H^+ ion is added to an OH^- ion.

2. What would happen if solution A & solution B were mixed? Explain your answer.

3. Classify the solution that forms in Exercise 2 as acidic, basic, or neutral &justify your classification in terms of the concentration of H^+ ions & OH^- ions.

4. Can a neutral solution contain H^+ &/or OH^- ions? Explain.

<u>Problems</u>

1. How many moles of H^+ ions are present in one litre of 2 *M* HCℓ?

2. How many moles of OH^- ions are needed to completely neutralize one litre of 2 *M* HCℓ?

3. How many moles of OH^- ions are present in one litre of 0.5 *M* NaOH?

4. How many moles of H^+ ions are needed to completely neutralize one litre of 0.5 *M* NaOH?

5. How many moles of OH^- ions are needed to completely neutralize 0.50 litres of 2 *M* of HCℓ?

Authored by: Neil Brosnan & Kenneth Levy; Revised by: Kelly Levy Edited by Linda Padwa & David Hanson, Stony Brook University

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Gas Unit – Pressure Units

On the weather report they tell you the current atmospheric pressure. The pressure we feel is due to the mass of all the air molecules above your head pushing down on you!

1. What is pressure in terms of force & area? Write an equation. What are the SI units of force & area?

2. The unit for pressure is the Kilopascal, of KPa. Standard Pressure is 101.325 kPa at sea level at 0°C. Most experiments are done above sea level at 25°C. Will the pressure be higher or lower that 101.325 above sea level? Explain.

SATP – standard ambient temperature & pressure: ____

STP – standard temperature & pressure: ____

3. Chemists are exceptionally smart people. They came up with a much better unit for pressure called the atmosphere (atm). They defined 1 atm to be the standard pressure at sea level at 0°C.
   a. State one advantage of this unit

   b. What is 101.325kPa in atmospheres?

4. The doctor measures your blood pressure with an old fashion unit. A normal blood pressure is 120/80 (systolic over diastolic). The units are mmHg. They are also called torr.
   Write mmHg in words.

5. With so many units for pressure (and there are more!) we need to convert between them. Note the following:

$$\frac{\quad}{101.325\ kPa} = \frac{\quad}{1\ atm} = \frac{\quad}{760\ mmHg}$$

6. What is normal diastolic blood pressure in atm? In kPa?

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Temperature

As you know, there is more than one scale for temperature. For example, Americans report temperature in Fahrenheit. In chemistry, we need another scale…

1. A student was overheard saying, "Today it's 20°C & yesterday was 10°C, so it's twice as hot today" Do you agree with them? Why or why not.

2. Celsius is referred to as a relative scale. What is it relative to?

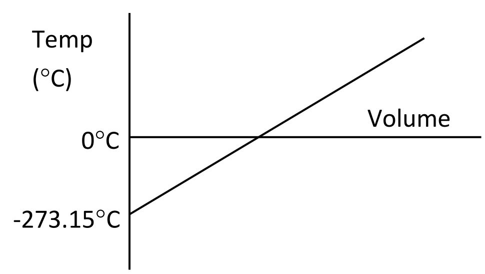
We need a temperature scale that is not relative but absolute.

3. If temperature is defined as the average kinetic energy (energy of motion) of particles in a substance, what is the coldest possible temperature?

   Absolute zero: ____

4. If zero degrees Celsius is 273.15K & 1°C increase in temperature is also a 1K increase in temperature, what's the equation to convert °C into Kelvin.
   (Note: there is no " ° " symbol for Kelvin).

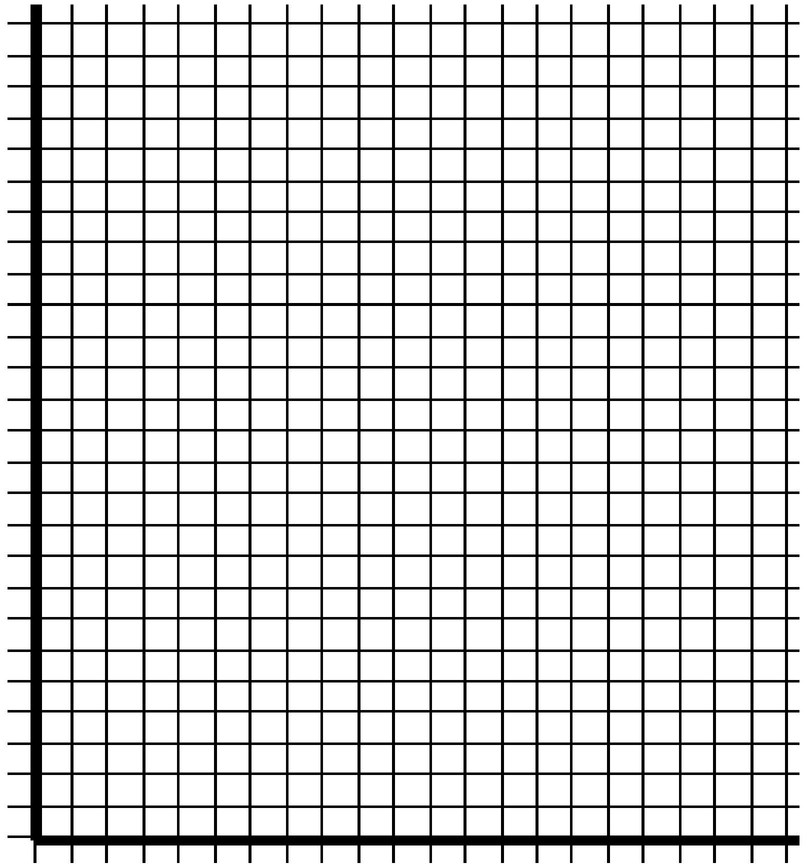
5. The following graph illustrates how absolute zero can be determined in the lab (at normal temperatures). Explain how it is done:



6. Go back to question 1. Using the Kelvin scale, how much hotter is 20°C than 10°C.

7. What temperature is twice as hot as 10°C (in Celsius). Show your calculation.

8. Use the apparatus at the front to determine absolute zero experimentally.



9. Watch the video, “Making stuff colder.”

a.) Can cold flow from one object to another?

b.) What is temperature?

c.) What is the issue with freezing organs for transplant?

d.) What is the advantage of a quantum computer?

e.) How many states of matter are there thought to be?

f.) Will scientists ever be able to achieve absolute zero?

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Gas Laws

1. List the five ideal gas law assumptions:

2. What are four variables that relate to gases? Include their units.

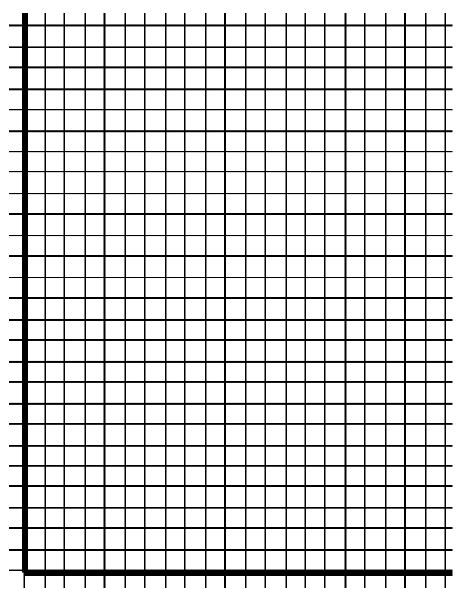
3. When doing an experiment, how many variables can we manipulate at a time? What are the variables that are held constant called?

Boyle's Law

4. With your group, predict how pressure & volume vary (direct? inverse?). Explain your reasoning.

5. What variables must be held constant?

6. Collect & graph the data from the demo. Which variable goes on which axis?



7. State the relationship between pressure & volume. This is called Boyle's law.

8. Can volume be measured in millilitres (mL)? Explain.

Example: Boyle's Law Calculation

9. 2.00 moles of a gas was placed in a balloon. If the original volume of the gas is 150 mL at 100 kPa, what will the volume become when the pressure is increased to 101.3 kPa?

Charles'Law

10. Predict how volume & temperature vary. Give an example to illustrate your prediction.

11. Which variables must be held constant?

    12. Test your prediction using the computer simulation. Any surprises?

13. Sketch a graph of volume & temperature. Which variable goes on the x-axis? Explain.

14. Use your graph to derive the relationship for volume & temperature. This is Charles' Law

15. Can temperature be measured in Celsius? Explain.

16. Must volume be measured in litres? Explain.

Example: Charles' Law Calculation

17. 2.00 moles of a gas was placed in a balloon. If the original volume of the gas is 150 mL at 10.0°C, what will the volume be for the gas at -20.0°C?

Gay-Lussac's Law:

18. Predict how pressure & temperature vary. Give an example to illustrate your prediction.

19. Test your prediction using the computer simulation. Any surprises?

20. Which variables must be held constant?

21. Sketch a graph of pressure & temperature. Which variable goes on the x-axis? Explain.

22. Use your graph to derive the relationship for pressure & temperature. This is Gay-Lussac's Law.

Example: Gay-Lussac's Law Calculation

23. A popcan has a volume of 355 mL. If the can experiences a pressure of 250 kPa at 20.0°C, what is the pressure when the temperature is 35.0°C?

The Combined Gas Law:

So far we have looked at three laws that relate pressure, volume & temperature.

24. Write the variation statements for P, V & T

25. Combine these relationships into one equation.

26. What must be held constant?

27. What are the possible units for:
    a. pressure
    b. volume
    c. temperature

Example: Combined Gas Law Calculation

28. A 0.75 L balloon at STP (standard temperature & pressure) undergoes a change to SATP (standard ambient pressure & temperature). What is the new volume of the balloon?

The Ideal Gas Law: (Version A)

The Combined gas law assumes that moles of gas is held constant. What if we change the moles?

29. Draw a 2.0L container with 5 gas molecules inside. Draw another 2.0 L container with 10 gas molecules. Which has greater pressure? What is this pressure a result of?

30. Write the variation statement for pressure (P) & moles (n)

31. Using the relationship in #29, take the Combined gas law & add the mole variable:

32. What assumption do we make when using the Ideal Gas Law?

33. Rewrite the Ideal Gas Law with mass instead of moles.

<u>The Ideal Gas Law: (Version B)</u>

Up to now, we have considered one gas under two different conditions. Now we want an ideal gas law where we can solve for any variable. If we graph the PV vs nT for any ideal gas, we get a straight line. The slope of the graph is referred to as the Universal Gas Constant & has its own symbol R.

34. Draw the graph of PV vs nT & label the slope R. Derive the Ideal gas law:

35. The Universal Gas Constant (R) depends on the units of pressure:
    - if pressure is in kPa, R = 8.3145 kPa.L/mol.K
    - if pressure is in atm, R = 0.08206 atm.L/mol.K

Let's rearrange the Ideal Gas Law to solve for other variables.

36. If molar mass is mass divided by moles, write the Ideal Gas Law to solve for molar mass of a gas (in g/mol).

37. If density is mass divided by volume calculate the density of a gas (in g/L). Start with the equation from #34.

38. If molar volume is the volume of one mol of a gas, calculate molar volume of a gas (in L/mol).

<u>Example: Ideal Gas Law Calculation</u>

39. What mass of carbon dioxide gas (in grams) is necessary to fill a 350 mL bottle at 0.980 atm & 12.0°C? Use your equation from #34.

40. What is the molar volume of any gas at STP?

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Gas Stoichiometry

Use the Ideal Gas Law (PV=nRT) to solve the following gas stoichiometry problem. Provide a full solution with attention to significant digits & a therefore statement.

For the complete combustion of propane, C_3H_8, all species are gaseous at 40.0°C & 110 kPa. If 2.40 L of propane was burned, what volume of water vapour at STP should be produced?

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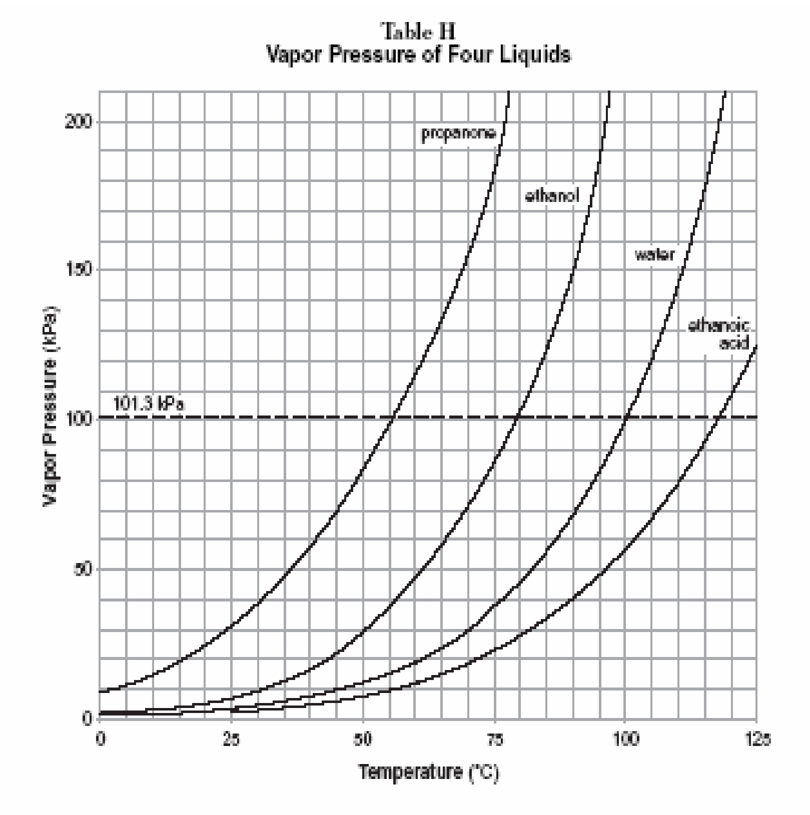
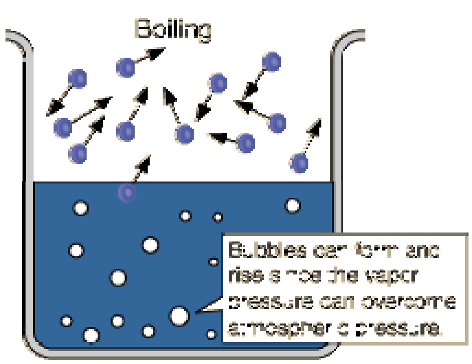
Vapour Pressure

A liquid boils when the vapour pressure equals the atmospheric pressure. Water boils at 100 ^{o}C at sea level, but in Calgary, Alberta, which is a kilometer above sea level and has lower atmospheric pressure, water boils at a different temperature. Certain substances, such as nail polish and paint, dry quickly because they have high vapour pressures.



Water's normal boiling point is 100^{o}C. At this temperature the vapour pressure of water is equal to 100 kPa, standard atmospheric pressure. If we were in a location with a different atmospheric pressure the boiling point would be different. For example, if the atmospheric pressure were 90 kPa, the boiling point of water would be

95^{o}C.



Key Questions

1. The vapour pressure curves of four liquids are shown in the graph in Model 1. What is plotted on the x-axis and what is plotted on the y-axis of the graph?

2. What happens to the vapour pressure of a substance when the temperature increases?

3. According to the information provided in Model 1, what determines the temperature at which a liquid boils?

4. What is the normal boiling point of propanone?

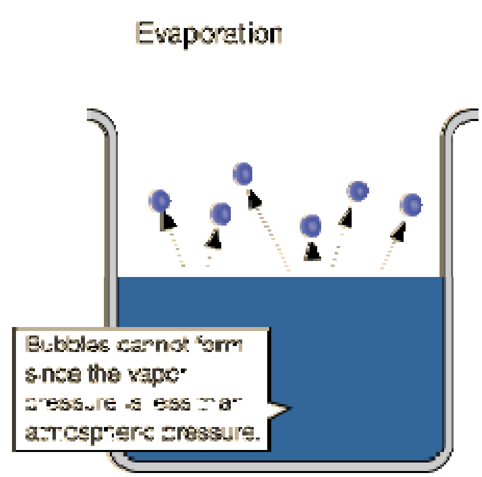
5. At what temperature will propanone boil if the atmospheric pressure is 70 kPa?

Exercise

1. List the four liquids in Model 1 in order of increasing vapour pressure at 60 ^{o}C.

Model 2

Evaporation, unlike vaporization, happens on the surface of liquids at all temperatures. This process is related to the strength of the forces holding the molecules in the liquid phase. The weaker the forces, the faster the molecules will escape from the liquid into the gas phase. A liquid with weak intermolecular forces will have a relatively large amount of vapour (gas phase) present above its surface.



Task

Place an equal amount of ethanol, propanone (acetone), and water on three separate cotton balls. Wipe the cotton balls on the desk at the same time. Observe the relative rate of evaporation for the liquids. Record your observations below.

Key Questions

1. Which liquid evaporated at the fastest rate?

2. Which liquid evaporated at the slowest rate?

3. Based on your observations, which liquid has the highest vapour pressure? Explain your answer.

4. Predict which of the three liquids used in this task would have the highest boiling point. Support your answer with an explanation.

5. Which of the three liquids has the strongest intermolecular forces of attraction? Support your answer with an explanation.

6. How do the intermolecular forces in propanone compare to the intermolecular forces found in water? Support your answer with an explanation.

Applications

1. A thermometer is placed in a beaker of water at room temperature. The beaker, water, and thermometer are covered by a bell jar attached to a vacuum pump. The pump is turned on and the pressure inside the bell jar is reduced. Predict what would be observed inside the bell jar.

2. Based on your predictions, suggest a possible boiling point for the water in the bell jar by using the information on the vapour pressure curve in Model 1. (Specify both temperature and pressure.)

3. Suggest a reason why changes need to be made in the cooking time when eggs are boiled in a location with a high altitude such as Calgary, when compared to the cooking time at a sea level location such as Vancouver. Support your answer with insight you have gained from this activity.

Authored by: Lisa Caputo; Revised by: Erin Graham, Rohini Quackenbush, Lizabeth Tumminello, Kenneth Levy and Kelly Levy
Edited by Linda Padwa and David Hanson, Stony Brook University

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Dalton's Law of Partial Pressures & Water Vapour Pressure

There are two possibilities when gases are combined:

- gases undergo a chemical reaction → solve with stoichiometry
- gases do not react → solve using Dalton's Law of Partial Pressures

1. What gases is air composed of?

2. Draw a diagram illustrating the composition of air in a container. Use the box to represent a container.

3. What is causing the pressure the container feels? Write an expression for the total pressure inside the container.

4. What would happen to the total pressure if inert helium gas was also put into the container?

<u>Dalton's Law of Partial Pressures:</u>The total pressure of a mixture of NON-REACTING gases is equal to the sum of the partial pressures of the individual gases: $P_{total} = P_1 + P_2 + P_3 + \ldots$
<u>Water Vapour Pressure:</u> When collecting gases under water (or via downward displacement of water), some water vapour may be included in the collected gas, thus it is important to subtract the water component from the total pressure before applying any gas law calculation

<u>Example: Finding Total Pressure</u>

If a 5.00 L canister is 7.00% argon & the remainder is oxygen. If the partial pressure of argon is 10.0 kPa, what is the total pressure of the canister?

<u>Example: Water Vapour Pressure</u>

A "Bic" lighter contains butane gas, C_4H_{10}. It is collected over water at 24.0°C & 107.4 kPa. If 1.45 L of butane was collected, what mass of butane was this?

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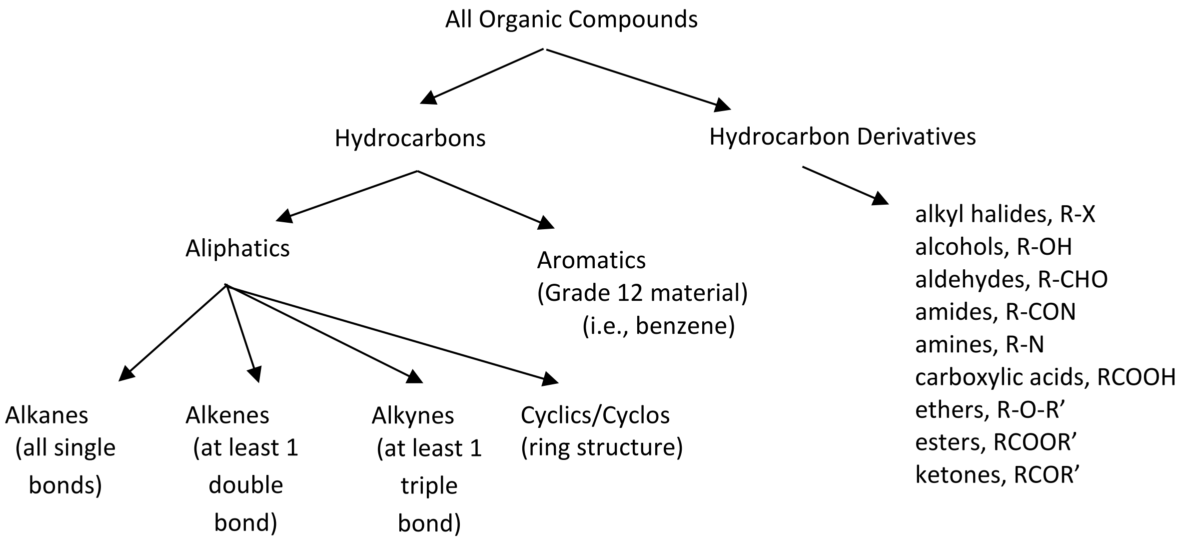
Organic chemistry

So far in grade 11 chemistry we've only studied inorganic chemistry. Organic chemistry concerns carbon based chemicals found in fuels, medicines & biologically important molecules (studied in grade 12 biology). In this course we'll look at just a few organic molecules & consider their physical properties. You will study organic chemistry more fully in grade 12 chemistry.

Let's start with an overview (including molecules you'll study in grade 12):



Alkanes:

1. Use your model kit to build a 6 carbon long chain. Use as many H atoms as you need.
   Draw the molecule Write the chemical formula

2. The molecule you built in question #1 can be drawn multiple ways. Draw at least two.

3. Draw the 3-D shape of a carbon atom. What is its VSEPR notation? What is its shape name?

4. Is C_6H_{14} (hexane) a polar or non-polar molecule? Calculate ΔEN for each bond in the molecule.

5. Does hexane have a high water solubility? Explain.

6. The gasoline in your car contains octane (an 8 carbon alkane). Does gasoline mix with water? Why or why not?

7. Draw the structural diagrams for two molecules of hexane. What intermolecular force(s) (LF, dd or H-bond) hold the two molecules together? Is it a strong or weak force?

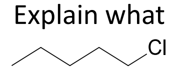
8. Does the length of the carbon chain affect the strength of the London Forces? Explain.

9. Is the boiling point of an alkane high or low? Explain

10. Summarize the general physical properties of alkanes:
    Water solubility: Boiling point: Flammability:

Alkyl Halides:

11. The following molecule is an alkyl halide. Explain what each word means:



12. Calculate the ΔEN for the C-Cℓ bond. Is it pure covalent or polar covalent?

13. Based on your calculation, is the molecule in #11 polar or non-polar?

14. Are alkyl halides soluble in water? Explain.

15. What intermolecular forces are found between alkyl halides?

16. Draw a six carbon alkane & a six carbon alkyl halide (you can use any halogen). Which has a higher boiling point? Explain.

17. Summarize the physical properties of Alkyl Halides:
Water solubility: Boiling point: Flammability:

Alcohols:

18. All alcohols contain an OH group called a "hydroxyl" group. What intermolecular forces hold two alcohols together?

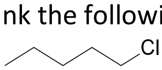
19. Will alcohols be highly water soluble? Explain.

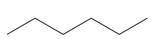
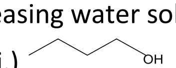
20. Complete the following table:

Molecule	Example	Polarity	Intermolecular Force	Water Solubility	Boiling Point
Alkane					
Alkyl halide					
Alcohol					

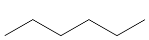
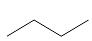
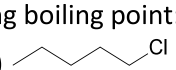
Questions:

21. Rank the following molecules in order of increasing water solubility:

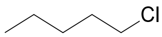
i.) 

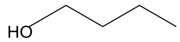
ii.) 
iii.) 

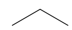
22. Rank the following in order of increasing boiling point:

i.) 
ii.) 
iii.) 

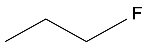
23. Which of the following molecules have H-bonds?

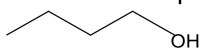
i.) 

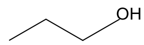
ii.) 

iii.) 

24. Which of the following molecules are polar? Explain.

i.) 

ii.) 

iii.) 

25. Which molecule in #24 is the MOST soluble in water? Explain.

EXAMINATION INFORMATION SHEET

Name: ______________________________

Phone Number: ______________________

Metal Activity Series:

Li K Ba Ca Na Mg Aℓ Zn Cr Fe Cd Co Ni Sn Pb H Cu Hg Ag Pt Au

Halogen Activity Series:

F Cℓ Br I

Water Vapour Pressure:

Temp (°C)	17.0	18.0	19.0	20.0	21.0	22.0	23.0	24.0	25.0	26.0	27.0	28.0	29.0	30.0
Vapour Pressure (kPa)	1.94	2.06	2.20	2.34	2.49	2.64	2.81	2.98	3.17	3.36	3.57	3.78	4.01	4.24

Constants:

R = $8.3145\frac{L.kPa}{mol.K}$ **or R =** $0.08206\frac{L.atm}{mol.K}$	**N_A = 6.02 x 10^{23} particles/mole**
1.00 atm = 101.325 kPa = 760 torr = 760 mmHg	

Equations:

m = nM	$ppm = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \cdot 10^6$	$P_1V_1 = P_2V_2$
$N = nN_A$	$ppb = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \cdot 10^9$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
n = CV	$\%composition = \frac{\text{mass of component}}{\text{total molar mass}} \cdot 100\%$	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
$C_1V_1 = C_2V_2$	$\%yield = \frac{\text{experimental yield}}{\text{theoretical yield}} \cdot 100\%$	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
$[H^+][OH^-] = 1.00 \times 10^{-14}$	$\%purity = \frac{pure yield}{impure sample} \cdot 100\%$	$\frac{P_1V_1}{m_1T_1} = \frac{P_2V_2}{m_2T_2}$
$pH = -\log[H^+]$	$m/m\ \% = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \cdot 100\%$	$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$
$pOH = -\log[OH^-]$	$v/v\ \% = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \cdot 100\%$	PV = nRT
pH + pOH = 14	$m/V\ \% = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \cdot 100\%$	$P_{total} = P_1 + P_2 + ...$
$d = \frac{m}{V}$	$\text{solubility} = \frac{\text{mass of solute (g)}}{\text{100mL of solution}}$	$M = \frac{mRT}{PV}$
Average Atomic Mass = $\sum(\%abundance)(isotopic mass)$		$d = \frac{MP}{RT}$
formal charge = valence electrons – ½(bonding pair electrons) – (all non-bonding electrons)		Mass # = # protons + # neutrons