HEBDEN :

CHEMISTRY 11

A WORKBOOK FOR STUDENTS

by

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Hebden :
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A Workbook for Students

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PREFACE

This book covers the entire British Columbia Chemistry 11 curriculum as revised in 1996. Some material from my book Chemistry: Theory and Problems, Book One (McGraw–Hill Ryerson, 1980) has also been included, with permission.

This book has been designed for student use and field tested for several years. My students have found the notes and exercises included in this book to be of great use. The final form of this book was greatly influenced by the suggestions and comments of many students. It is hoped that this book will be of use to other teachers and students alike. When students miss a class, having to learn the material on their own is often a frustrating experience for both students and teacher. It is my hope that the notes, explanations and exercises in this book will make it easier to "catch up" (or to work ahead).

ACKNOWLEDGEMENTS

The support, advice and encouragement of my wife, Frances, has played a major role in the creation of this book. Her continued suggestions and help with the entire project, from initial concept though to final editing, have made possible an otherwise almost impossible task.

The experienced advice, mentoring and friendship of Gordon Gore is greatly appreciated and acknowledged. His encouragement and humour were of enormous help in all stages of the preparation of this book. He truly defines the word "educator".

The cartoons in this book were drawn by the highly talented Ehren Stillman ("ZimBoBwe"), of Mission, B.C. The cartoons have enriched this book immeasurably and helped to illustrate the human side of chemistry.

The legions of students and teachers who gave me feedback and suggestions on earlier versions of this book are gratefully acknowledged.

COVER PHOTOGRAPH

Some of the main themes of Chemistry 11 are highlighted in the cover photograph. The molecular model shows the bonding in a diamond crystal. The combustion reaction is giving off heat and light and the broad leaves of the plant are engaged in photosynthesis while producing organic compounds. Photosynthesis is the reverse of a combustion reaction.

The design of the cover by Loren Hebden is gratefully acknowledged.

PERMISSIONS

Data from the tables

Periodic Table of the Elements
Standard Atomic Weights
Physical Constants of Organic Compounds
Physical Constants of Inorganic Compounds
Conversion Factors
Density of Various Solids
Elements in Sea Water
Enthalpy of Combustion of Selected Organic Compounds
Strengths of Chemical Bonds
Table of the Isotopes, and
Ionization Potentials of Atoms and Atomic Ions

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The table "Electronegativities of the Elements" on page 194 is reproduced from L. Pauling, *The Nature of the Chemical Bond and the Structure of Molecules and Crystals*, Cornell University Press, Ithica, New York, 1960. Permission to use this table is gratefully acknowledged.

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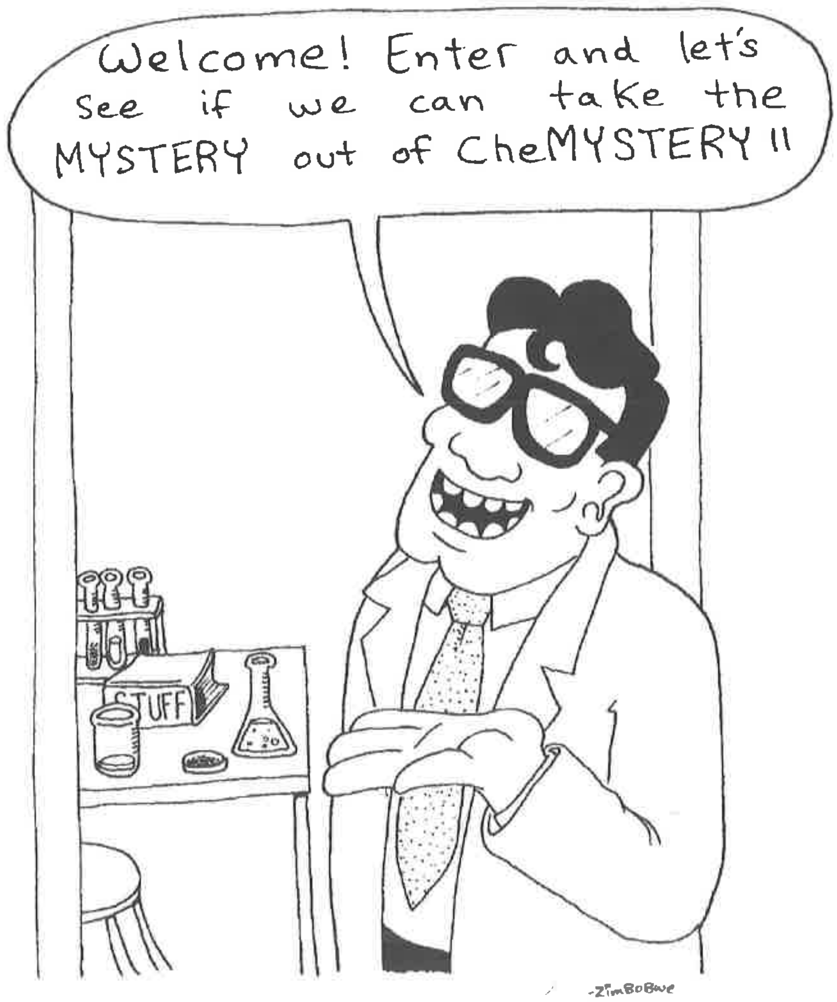
Chemistry 11 is intended to give students an understanding of many principles of modern chemistry. Since not all students taking Chem 11 go on to take Chem 12, Chem 11 is a "survey course" which covers the "basics". Some of the topics are covered in previous science courses but the present course extends those topics and introduces many new ones. After completing Chem 11 you should be able to understand the chemistry underlying many of the natural processes around you, as well as much of what you read in magazines or newspapers.

Chemistry 11 also acts as a stepping stone to Chemistry 12 and beyond that to university or college science courses. One of the most enjoyable parts of Chem 11 is the laboratory work. You will be taught the proper procedures for working safely in the laboratory, for handling chemicals and for using specialized chemical equipment. Much of the pleasure of chemistry is found in the variety of colour changes you will see and the manner in which crystals grow, gases and solids suddenly form, and energy is given off or absorbed. In true experimental work, you should experience a sense of wonder, awe and surprise. Chemistry is fun!



UNIT I : SAFETY IN THE CHEMICAL LABORATORY

In Chemistry 11 you will be taught how to deal with chemicals and chemical equipment in a safe manner: compliance with all safety procedures will be strictly enforced. You will be shown how to use safety equipment and the emergency procedures to follow in the unlikely event of an accident. Then it will be up to you to work safely. Unsafe behavior and "fooling around" in a lab are not funny – they are STUPID and endanger your classmates and yourself.

Before each experiment you will be told the **special** precautions to be observed with the chemicals and equipment to be used. At all times, **general** safety procedures must be followed and you must know how to use the available protective and safety equipment.

In the event of **ANY** emergency, your teacher must be notified. However, it is not necessary to get your teacher's attention and permission **BEFORE** using emergency equipment; a quick yell as you go for the equipment will notify everyone that an emergency has arisen.

I.1. EMERGENCY EQUIPMENT

The following five pieces of equipment are only intended for use in the event of an emergency. STUDENTS DO NOT HAVE TO ASK PERMISSION TO USE ANY EMERGENCY EQUIPMENT, but must be prepared to justify why they used the equipment. Not all schools will necessarily have all the equipment listed.

You should sketch a good map which includes your chemistry room, the surrounding hallways, the designated fire exit for the room, and the location where you are to assemble once the building is evacuated during a fire drill.

A. FIRE EXTINGUISHERS

Location: Indicate on your map the location of
a) **all** the fire extinguishers in your laboratory.
b) the closest fire extinguisher **outside** the laboratory.

When to use: Generally, your teacher will use the fire extinguisher, but if your teacher is absent temporarily or can't get to an extinguisher in time you may need to use it. If, in your opinion, the fire is uncontrolled and small enough to be put out with a fire extinguisher, use one. (Other students must immediately leave the room and someone must pull the fire alarm.) IF YOU DOUBT THAT THE FIRE IS SMALL ENOUGH THAT A FIRE EXTINGUISHER WILL DO THE JOB, quickly leave the room (someone should turn off the main gas supply for the room if that can be done safely), close the door behind you, pull the fire alarm and evacuate the building. Don't try to be a hero; fighting a substantial fire **MUST** be left to trained fire–fighters.

How to use: Take the extinguisher to the location of the fire, grab the extinguisher by the handle and yank the safety pin out of the side of the handle with a sharp pull.

a) Small extinguishers
- If the extinguisher has a fixed "horn", aim the horn at the **BASE** of the flames and pull the trigger, sweeping the spray back and forth over the area in flames.
- If the extinguisher has a hose and nozzle secured to the extinguisher body by clips, remove the hose from the clips so that you can hold the extinguisher in one hand and direct the spray with the other. Aim the nozzle at the **BASE**

of the flame and pull the trigger, sweeping the spray back and forth over the area in flames.

b) Large extinguishers can be set on their base and their hose freed from the retaining clips, leaving both hands free to pull the trigger and direct the spray.

***Time and distance of effectiveness*:** Large extinguishers will spray for about 20 s; smaller ones for about 10 s. You have to get within 4–5 m (12–15 ft) for the spray to be effective, so don't waste the extinguisher contents with "practice sprays" or spray from too far away.

***Comments*:** Do not spray the contents of an extinguisher on a person! The spray from some types of extinguishers can instantly freeze flesh or drive powder into the eyes or lungs. (A fire blanket is used when a person's clothing or hair has caught fire.) The extinguishers available in a high school chemistry laboratory usually are "general purpose" ones which spray out powdered baking soda and can be used with the chemical, electrical and general fires which might occur.

B. FIRE BLANKET

***Location*:** Indicate on your map the location of the fire blanket in the lab.

***When to use*:** There are two possible situations where a fire blanket should be used.

a) A fire blanket should be used when a student's clothing or hair catches fire. The blanket must be used **very quickly** in order to minimize injury.

b) A fire blanket can be used to smother burning material on the floor or a bench, provided the fire can be approached with sufficient safety to allow the blanket to be placed over the entire area involved. Since both a blanket and extinguisher can be used on a small fire, use whichever one you can get to most quickly.

***How to use*:** Pull the cord at the bottom of the fire blanket "tube" to get the blanket out.

- A student on fire must "stop, drop and roll". Throw the fire blanket over the student as soon as possible. Once the fire is out, remove any burned clothing (unless it has melted onto the skin). It is very important to get burned skin cooled as soon as possible.
- If there is a small fire on a lab bench or the floor, the fire blanket may be thrown carefully over the area involved in flames. Be careful not to knock over any beakers or flasks containing flammable materials if they are close to the flames.

***Comments*:** A standing person burns like a candle ... but MUCH QUICKER. If you catch on fire, immediately drop to the floor and roll around. If you panic and run around, someone must get the fire blanket and wrap it around you, get you on the floor and roll you in the blanket until the fire is out. Sometimes it may be necessary to get you onto the floor and roll you around while the fire blanket is coming.

C. EYEWASH FOUNTAIN OR EYEWASH STATION

Location: Indicate on your map the location of the eyewash fountain or eyewash station.

When to use: The eyewash MUST be used any time a chemical or solution gets into eyes.

How to use: As soon as something gets in your eye, yell for help and **HURRY** to the eyewash; if you cannot see, yell for someone to help you.

a) **Use of an Eyewash Fountain:** Push the vertical paddle back with your hand and put your face down into the stream of water so that water strikes your eyes DIRECTLY. You **MUST** keep your eyes **OPEN** in the stream of water, blinking rapidly to help wash underneath the lids. Keep washing the eyes for at least 5 minutes, unless it is just a harmless substance (how do you know?); toward the end of this time the eyes may be sore just from the cold water, but pain is better than blindness. If you are panicking someone may have to hold your head in the fountain and help by giving you encouragement and reminders of what to do.

b) **Use of an Eyewash Bottle:** If possible, get to a sink so that water can be splashed up into your eyes. Then lie on your back so that the people administering first aid can slowly drip liquid from an eyewash bottle into the affected eye(s). This washing should continue for at least 10 minutes.

In all cases someone must call for help. If your teacher and/or medical aide deems it necessary, you will be prepared for transport to a medical facility.

Comments:

a) Contact lenses must be removed for proper cleaning. You may have to wash the "contacts" out if they can't be quickly "popped out". In general, do a quick preliminary wash, then pop out the contacts and wash thoroughly.

b) You have less than one second to get DILUTE acid or base out of your eyes before damage starts to occur — don't waste time getting to the eyewash. When you yell, **everyone** must get out of the way and move obstacles out of your way.

c) Medical treatment will usually be required to assess the extent of harm to your eyes.

d) If there is reason to suspect you have glass in your eyes as well as chemicals, the situation becomes complicated and more serious. If, while your eyes are being washed in the eyewash fountain or with the eyewash bottle, you feel a foreign object in your eyes, call for help. The first aid attendant may attempt to remove the object. If the object cannot be removed, both your eyes will be bandaged to prevent you from blinking and you will be transported on your back as soon as possible to a medical facility.

D. EMERGENCY SHOWER

Location: Indicate on your map the location of the emergency shower.

When to use: The shower is used when hazardous chemicals spray over large areas of the body.

How to use: There are two common types of emergency showers in use. One involves an overhead shower having a pull–ring, while the other is a hand–held shower with an on–off handle. Either stand under the shower head and pull down on the ring or turn the on–off handle and hold the shower head so as to spray water liberally over the affected area. If a hazardous liquid chemical or solution soaks into your clothing, the affected clothing **must** be removed after the washing process. Your teacher will discuss with you the procedures which must be followed to protect the modesty of students. Do not worry if water gets on the floor (but be careful not to slip). Any clothing removed should be thoroughly washed before being put on again. Someone may be able to assist with a change of clothing such as "gym" clothes.

E. ACID–BASE NEUTRALIZING SOLUTION

(This is not actually "equipment" but is a specialized solution to help control the harmful effects of certain accidents. There is no need to alert everyone in the room that you are using the solution.)

Location: Indicate on your map the location of the neutralizing solution.

When to use: Use this solution whenever an acidic or basic ("caustic") solution has come in contact with your skin. DO NOT USE IT IN YOUR EYES – use the eyewash if something gets in your eyes. If unsure as to whether a solution is acidic or basic, ask your teacher for instructions or use the neutralizing solution just to be safe. The solution usually consists of a dilute solution of sodium acetate and acetic acid ("vinegar"), and is more or less harmless to the skin for short periods of exposure.

How to use: First wash the affected area with large amounts of water and then pour some of the neutralizing solution on the affected area and gently wash the skin with the solution. If the skin felt slippery before applying the neutralizing solution, the chemical which contacted the skin was a base and prolonged washing of the skin must take place to help get all of the base removed: basic solutions tend to "eat their way" under the top layers of the skin. After washing with the neutralizing solution, continue to wash with warm soapy water.

SOME FINAL NOTES ON EMERGENCIES

a) Priorities

If more than one piece of equipment is needed or more than one problem has arisen, tend to the most serious problem first.

- a person on fire (immediately life–threatening), then
- a person with chemicals or glass in their eyes (threatens permanent blindness), then
- a person soaked with chemicals (harm to skin; generally a slower reaction due to natural protective oils on skin)

b) Sources of First–Aid Assistance

List the people trained to provide first aid in your school.

I.2. PROTECTIVE EQUIPMENT

The purpose of protective equipment is to protect you from the effects of hazardous chemicals or material BEFORE any problems arise.

A. SAFETY GOGGLES

Location: Indicate on your map where the safety goggles are stored.

When to use: Safety goggles **MUST** be used whenever chemicals are being used or glass–working is being performed. Goggles must be put on **PRIOR** to handling any chemicals or glass and must not be removed until **AFTER** you have disposed of or put away all chemicals or glass. **Putting on safety goggles at the start of a lab period MUST become an automatic reflex!**

How to use: Goggles must fit snugly; if a lens is loose or a strap is broken, get a replacement pair and give the defective pair to your teacher. Goggles must not simply be held on the face with a hand, moved up onto your forehead, or left to dangle around your neck.

B. FUME HOODS

Location: Indicate on your map the location of the fume hoods.

When to use: The fume hoods must be used whenever poisonous or offensive odours are being produced. At no time should such odours be allowed to enter the classroom. Do not carry a reaction mixture out of the fume hood to show your teacher or partner; instead ask them to come to the fume hood. Normally, the instructions for a experiment will specify the use of a fume hood, if required, as part of the procedure.

How to use: Learn where the On/Off buttons for the fume hood are located (they generally are below or beside the fume hoods). It is **NOT** necessary to pull down the sliding glass partition during usage; the hoods are supposed to have sufficient draft to keep fumes out of the room even when the sliding glass is fully up. The only time it should be necessary to pull down the glass is when a reaction may spatter out of its container or a strong draft is required to accelerate an evaporation process. If you pull down the glass and try to work with your hands extended under the glass your cramped movements may endanger you if something unexpected suddenly causes you to try to jerk your hands out.

I.3. IN CASE OF FIRE

a) The FIRST and most important thing to do is to back out of harm's way and evaluate the situation.

b) Next, warn the teacher and other students with a shout.

c) ***CONTROLLED FIRES***: If a fire is controlled, in the sense that it is contained in a beaker, flask or test tube, the fire can often be put out by placing a watch glass or inverted beaker over the top of the container and smothering the fire. **Be VERY careful not to spill the contents.** If unsure that you want to attempt such a maneuver, simply call for help. DO NOT PANIC – even if NOTHING is done the fire normally will burn itself out. If the fire involves a small amount of burning liquid on a benchtop and might take a minute or so to burn itself out, the fire blanket can be used carefully to smother the fire.

d) ***UNCONTROLLED FIRES***: If the fire is not minor and will possibly continue to spread, everyone must immediately evacuate the room except those who may be using a fire extinguisher. If possible, someone should turn off the main gas supply in the room. Also, someone must pull the fire alarm to start the evacuation of the building. The door must be left closed after the last person is out.

Important: Students **must** quickly go to the designated assembly point so that a roll call can be made to check that everyone has safely made it out of the building. You must NOT go elsewhere to "visit with friends" since your teacher is required to be know whether you are safely out of the building. Your whereabouts MUST be known so as to guarantee your life is not in danger!

I.4. SOME LABORATORY HAZARDS

	HAZARD	NATURE OF HAZARD	HOW TO DEAL WITH HAZARD
1.	Spilled chemicals	Chemical burns	Notify teacher for cleanup instructions, but keep away in the meantime.
2.	Broken glass	Cuts; chemicals in cuts	Notify teacher for cleanup instructions if chemicals are mixed with the glass. Otherwise, use broom / dustpan provided and put in special recepticle for broken glass.
3.	Burning chemicals in container	Burns	Step back and notify class, then deal with the fire as outlined above ("IN CASE OF FIRE; CONTROLLED FIRES").
4.	Chemicals on hands	Chemical burns; skin irritation or allergic reaction	Wash off immediately under fast–running water. Then use NEUTRALIZING SOLUTION if the chemicals are an acid, a base or have properties unknown to you.
5.	Being asked to smell chemical vapours	Strong odours may injure nasal passages	Holding the container in front of you, dilute the smell by gently "wafting" the odour to your nose with a wave of the hand over the container and toward the nose.
6.	Bunsen burners	Burns; fires	Tie long hair back or use elastics; don't keep burner gas on for more than a few seconds if burner won't start (seek help rather than filling the room with flammable gas).
7.	Loose hair or "floppy" clothing / accessories	Burns or chemical spillage; equip–ment knocked onto floor	Tie long hair back or use elastics; remove ties or tuck into shirt front; secure or remove loose clothing accessories such as scarves or dangling jewelry. Wide, loose sleeves should be secured at the wrist by an elastic.

I.5. DISPOSAL OF CHEMICALS

1. Disposal of Unused Chemicals

Never put unused chemicals back into their original containers! If you have taken too much of a chemical, ask if anyone else in your class can use it or ask your teacher for disposal instructions. Putting chemicals back into their original container poses two problems:

- the chemical may be put in the wrong container, spoiling the chemicals or starting a reaction.
- the chemicals may be contaminated by using glassware that was not perfectly clean and dry.

2. Disposal of Used Chemicals

You will be given instructions as to how to dispose of the chemicals used in each experiment. The sink or waste paper basket is NOT to be used unless so instructed. Many of the chemicals you will use are harmless to the environment, but some are extremely destructive. Let's take care, eh!

YOU DON'T NEED PERMISSION TO USE SAFETY EQUIPMENT

I.6. GENERAL RULES OF SAFE LABORATORY CONDUCT

1. There **MUST** be no horseplay in the lab.
2. There **MUST** be no running in the lab. Always look where you are going and don't turn around quickly; you may spill chemicals being carried by someone else. Notify people if you are passing behind them with a container of chemicals, so that they know not to make any sudden moves.
3. You must not carry out unauthorized experiments. For example, you must not mix chemicals in other than the way in which you were told.

To summarize: you must always have a "conscious safety attitude".

A **"CONSCIOUS SAFETY ATTITUDE"** means you should always think about the possible safety–related consequences of any action you are planning.

QUESTIONS:

1. What is the first thing you must do in each of the following situations?
   (a) A burning liquid spills on your clothing.
   (b) An unknown solution sprays into your eyes.
   (c) The chemical in a beaker on your lab bench catches fire.
   (d) A beaker full of liquid chemical falls onto the floor, breaking the beaker and spreading the chemical.
   (e) You get a few millilitres of acid or base on your hands.
2. You are having trouble getting a bunsen burner to light and after 15 seconds the gas still won't ignite. What should you do at this point?
3. When using a fire extinguisher, where should you aim the spray?
4. How should you handle a controlled fire in the laboratory?
5. Why is it not necessary to ask permission before using emergency equipment?
6. How long will a small fire extinguisher continue to spray?
7. If a chemical sprays into a student's eyes and simultaneously a flaming liquid sprays on his clothes, what should be done first?
8. A beaker full of acid drops on the floor. In addition to the danger from getting acid on your hands, why shouldn't you attempt to pick up the pieces of glass with your hands?

Safety Notes:

UNIT II : INTRODUCTION TO CHEMISTRY

Before you learn how to make stinks and bangs in your chemistry lab, there are a few tiny little details to attend to ... such as how to read the scales on the equipment you will be using, how to handle the units used in Chemistry 11 and how to decide how good your data is. This unit gives you the background needed for the remainder of Chemistry 11.

II.1. UNIT CONVERSIONS

This section shows how to use a mathematical method called Unit Conversions which will be used extensively in Chemistry 11 and 12. Initially, you will be solving relatively easy problems. **Avoid the temptation to solve the problems by your own method; you should learn the Unit Conversion method.** OK, let's get on with the game.

If eggs are $\frac{\$1.44}{1 \text{ doz}}$, another way to say this is that eggs are $\frac{1 \text{ doz}}{\$1.44}$.

> The statement "$1.44 per dozen" allows us to RELATE or CONNECT one amount ($1.44) to another amount (1 dozen). Both
>
> $$\frac{\$1.44}{1 \text{ doz}} \text{ and } \frac{1 \text{ doz}}{\$1.44}$$
>
> make the same connection implied by the statement:
>
> $$\$1.44 = 1 \text{ doz}$$
>
> where the "=" sign here is interpreted as "**IS EQUIVALENT TO**".

Definition: A **CONVERSION FACTOR** is a fractional expression relating or connecting two different units.

Examples:

STATEMENT FORM	CONVERSION FACTORS
1 min = 60 s	$\frac{1 \text{ min}}{60 \text{ s}}$ and $\frac{60 \text{ s}}{1 \text{ min}}$
$1 = 100 ¢	$\frac{\$1}{100 ¢}$ and $\frac{100 ¢}{\$1}$

Look at one of the conversion factors that relate "minutes" to "seconds".

$$\frac{1 \text{ min}}{60 \text{ s}}$$

the TOP part
EQUALS
the BOTTOM part

Dividing "1 min" by something **EQUAL TO** 1 min produces **a fraction with a value equal to "1".** Multiplying any expression by this conversion factor is equivalent to multiplying by "1" and therefore WILL NOT CHANGE THE VALUE of the expression. The next example shows how a conversion factor is used.

EXAMPLE: **How many minutes are there in 3480 seconds?**

$$\text{\# of minutes} = 3480\text{ s} \times \frac{1\text{ min}}{60\text{ s}} = \mathbf{58\ min}$$

Both "60 s" and "1 min" are the same length of time (multiplying by the conversion factor didn't change the VALUE of the time). However, the units are different after using the conversion factor: the question starts with a ***large number*** of **SMALL** time units and ends up with a ***small number*** of **LARGE** time units.

The method of unit conversions uses conversion factors to change the units associated with an expression to a different set of units.
Every unit conversion problem has three major pieces of information which must be identified: i) the unknown amount and its UNITS, ii) the initial amount and its UNITS, and iii) a conversion factor which relates or connects the initial UNITS to the UNITS of the unknown.

INCREDIBLY, VITALLY IMPORTANT NOTE!

In all the calculations which follow you must **ALWAYS** include the units, for they are the "major players" in the calculation. If you are tempted to omit or "forget about" the units, DON'T! The course you fail could be Chem 11!

EXAMPLES: **a) What is the cost of 2 doz eggs if eggs are \$1.44/doz?**

The first thing to do is tear this problem apart and analyze the information it contains.

$$\underbrace{\text{What is the cost}}_{\text{UNKNOWN AMOUNT}} \quad \underbrace{\text{of 2 doz eggs}}_{\text{INITIAL AMOUNT}} \quad \underbrace{\text{if eggs are \$1.44/doz}}_{\text{CONVERSION STATEMENT}}$$

The **UNKNOWN AMOUNT and its UNIT:**

- The unknown is identified in ***the phrase which asks the question.*** In this problem the UNKNOWN is the "cost".
- Since the only unit of "cost" mentioned in the problem is **dollars** (<u>\$</u>1.44/doz), use this unit with the unknown amount.

UNKNOWN AMOUNT = # of dollars

The **INITIAL AMOUNT and its UNIT:**

- It is in a PHRASE CONNECTED OR DIRECTLY RELATED TO THE PHRASE CONTAINING THE UNKNOWN — "What is the cost ***of*** 2 doz eggs . . .". Notice that in this case the word "of" connects the unknown amount to the initial amount.
- It is a number with a **SINGLE UNIT** ("doz"). The only other number mentioned, \$1.44/doz, involves two units.

INITIAL AMOUNT = 2 doz

The **CONVERSION STATEMENT:**

- involves TWO DIFFERENT UNITS ("\$" and "doz"), AND
- is a separate statement which does not involve a question.

The conversion statement gives the information needed to make the conversion factor. The possible conversion factors are

$$\frac{\$1.44}{1\text{ doz}} \text{ and } \frac{1\text{ doz}}{\$1.44}.$$

PUTTING EVERYTHING TOGETHER completes the unit conversion. (If you follow what happens here, fine. Otherwise, don't worry; you will be shown how to "put everything together" next.)

$$\text{\# of dollars} = 2\ \cancel{\text{doz}} \times \frac{\$1.44}{1\ \cancel{\text{doz}}} = \mathbf{\$2.88}$$

Notice that the unit "doz" cancels

b) If a car can go 80 km in 1 h, how far can the car go in 8.5 h?

Again, dissect the sentence.

$$\underbrace{\text{If a car can go 80 km in 1 h,}}_{\text{CONVERSION STATEMENT}}\quad \underbrace{\text{how far can the car go}}_{\text{UNKNOWN AMOUNT}}\quad \underbrace{\text{in 8.5 h?}}_{\text{INITIAL AMOUNT}}$$

The **UNKNOWN AMOUNT and its UNIT:** The part of the sentence which asks the question ("how far can the car go") implies that the unknown is a distance. Since the only unit of distance mentioned is "km" ("80 **km** in one hour"), use this as the distance unit.

UNKNOWN AMOUNT = # of kilometres

The **INITIAL AMOUNT and its UNIT:** The initial amount, 8.5 h, is connected directly to the unknown — "how far can the car go ***in*** 8.5 h" — and has a single unit ("h").

INITIAL AMOUNT = 8.5 h

The **CONVERSION STATEMENT:** This statement is recognized because it

- makes a statement involving a number with no question asked or implied, AND
- mentions two different units (km and h).

The possible conversion factors are

$$\frac{80\ \text{km}}{1\ \text{h}} \text{ and } \frac{1\ \text{h}}{80\ \text{km}}.$$

PUTTING EVERYTHING TOGETHER in a complete unit conversion:

$$\text{\# of kilometers} = 8.5\ \cancel{\text{h}} \times \frac{80\ \text{km}}{1\ \cancel{\text{h}}} = \mathbf{680\ km}\ .$$

Again, note that the unit "h" cancels.

EXERCISE:

1. For each of the following problem statements identify
   - the unknown amount and its unit,
   - the initial amount and its unit, and
   - the conversion factors and their units.

   (You aren't required to put everything together and solve the problem yet . . . that comes next.)

   a) If a chemical costs $50 per gram, what is the cost of 100 g of the chemical?
   b) Computer disks cost $6.00 for 10 disks. How many disks can you buy for $36.00?
   c) Cork has a density of 0.35 g/mL. What is the volume of 20 g of cork?
   d) If 3 kiwi fruit sell for $1, how many kiwi fruit can you buy for $5?
   e) If 4 bims are worth 5 tuds, how many bims can you buy for 30 tuds?
   f) A farmer trades 2 cows for 7 goats. At this rate, how many goats can he get for 10 cows?
   g) One mole of oxygen has a mass of 32 g. What is the mass of 5.5 moles of oxygen?
   h) One molecule of sulphur contains 8 sulphur atoms. How many sulphur molecules can be made from 104 sulphur atoms?
   i) How long must an electrical current of 35 coulombs/s flow in order to deliver 200 coulombs?
   j) What temperature increase is caused by 100 kJ of heat if 4.18 kJ of heat causes a 1 ^{o}C increase in temperature?

HOW TO PUT EVERYTHING TOGETHER

The method of unit conversions may seem a little awkward at first, but later it will allow you to solve some complicated problems **in one line. Also, it is "SELF–CHECKING", allowing you to check the "correctness" of your results**!

The general form of a unit conversion calculation is shown below.

(UNKNOWN AMOUNT) = (INITIAL AMOUNT) X (CONVERSION FACTOR)

EXAMPLES: **a) If 0.200 mL of gold has a mass of 3.86 g, what is the mass of 5.00 mL of gold?**

The **UNKNOWN AMOUNT and its UNIT**: The question asks "What is the mass", which suggests finding "# of grams".

The **INITIAL AMOUNT and its UNIT** is "5.00 mL", which is tied to the unknown amount ("What is the mass") by the connector "of".

The **CONVERSION STATEMENT** is "If 0.200 mL of gold has a mass of 3.86 g". The amounts being connected are 0.200 mL and 3.86 g.

Now to solve the problem. Put the **unknown amount** on the **left** side of an "=" sign to identify what you are trying to find.

$$\# \text{ of grams} =$$

Then put the **initial amount and unit** on the **right** side of the "=" sign.

$$\# \text{ of grams} = 5.00 \text{ mL}$$

Next multiply the initial value by a conversion factor. Construct the conversion factor from the conversion statement as follows.

The conversion statement connects "0.200 mL" and "3.86 g"; possible conversion factors are

$$\frac{0.200 \text{ mL}}{3.86 \text{ g}} \text{ and } \frac{3.86 \text{ g}}{0.200 \text{ mL}} .$$

Use the conversion factor which has "0.200 mL" on the bottom. **THE PURPOSE OF PLACING "0.200 mL" ON THE BOTTOM OF THE FRACTION IS TO ALLOW THE UNIT "mL" TO *CANCEL*.**

$$\# \text{ of grams} = 5.00 \cancel{\text{mL}} \times \frac{3.86 \text{ g}}{0.200 \cancel{\text{mL}}}$$

Finally, carry out the multiplication and finish the problem.

$$\# \text{ of grams} = 5.00 \cancel{\text{mL}} \times \frac{3.86 \text{ g}}{0.200 \cancel{\text{mL}}} = \mathbf{96.5 \text{ g}}$$

This problem started with the unit "mL" and eventually **converted** to the unit "g"; hence the term "***Unit Conversion***". To show that everything has been done properly, notice that the procedure started with "# of grams" on the left, and found 96.5 g as an answer.

The conversion statement allows you to make two possible conversion factors:

$$\frac{0.200 \text{ mL}}{3.86 \text{ g}} \text{ and } \frac{3.86 \text{ g}}{0.200 \text{ mL}} .$$

The required conversion factor was BUILT by arranging the fraction in such a way as to cancel the initial unit "mL". If the other conversion factor had been used (that is, the fraction was built upside–down), the calculation would have given:

$$\text{\# of grams} = 5.00 \text{ ml} \times \frac{0.200 \text{ mL}}{3.86 \text{ g}} = 0.259 \frac{(\text{mL})^2}{\text{g}} = \text{a mess!!!}$$

Therefore, whenever you multiply the initial value by a conversion factor you have to ask yourself:

"WHICH WAY DO I HAVE TO WRITE THE CONVERSION FACTOR IN ORDER TO ALLOW THE INITIAL UNITS TO CANCEL PROPERLY?"

b) If 0.200 mL of gold has a mass of 3.86 g, what is the volume occupied by 100.0 g of gold?

The **UNKNOWN AMOUNT and its UNIT:** The question asks "what is the volume", which suggests finding "# of millilitres".

The **INITIAL AMOUNT and its UNIT** are "100.0 g", which is tied to the unknown amount ("what is the volume") by the connector "occupied by".

As in the previous example, the **CONVERSION STATEMENT** is "If 0.200 mL of gold has a mass of 3.86 g". The amounts being connected are 0.200 mL and 3.86 g.

Now to solve the problem. Start with the **unknown amount** on the **left** side of an "=" sign.

$$\text{\# of millilitres} =$$

Then put the **initial amount and unit** on the **right** side of the "=" sign.

$$\text{\# of millilitres} = 100.0 \text{ g}$$

Construct a conversion factor from the conversion statement such that the starting unit "g" is cancelled by having "3.86 g" on the bottom.

$$\text{\# of millilitres} = 100.0 \cancel{\text{g}} \times \frac{0.200 \text{ mL}}{3.86 \cancel{\text{g}}}$$

Finally, carry out the multiplication and finish the problem:

$$\text{\# of millilitres} = 100.0 \cancel{\text{g}} \times \frac{0.200 \text{ mL}}{3.86 \cancel{\text{g}}} = \mathbf{5.18 \text{ mL}}$$

Again, notice that the problem tried to find "# of millilitres" and found 5.18 mL as an answer. Also, note that the conversion factor used in this problem, 0.200 mL/3.86 g, was the inverse of the conversion factor used in the problem above, 3.86 g/0.200 mL. **The way the conversion factor is used depends on which unit is to be cancelled.**

SUMMARY OF THE PROCEDURE TO BE USED WITH UNIT CONVERSIONS

1. Identify the unknown amount and its unit. Write these down on the left–hand side of an "=" sign.
2. Identify the initial amount and its unit. Write these down on the right–hand side of the "=" sign.
3. Identify the conversion factor. Multiply the initial amount by the conversion factor in such a way that one of the units in the conversion factor cancels the unit of the initial amount.
4. Complete the problem by multiplying and/or dividing the amounts on the right–hand side.

EXERCISE:

2. Solve the following using the method of unit conversions.
   a) If there are 6.02×10^{23} atoms in 1 mol of atoms, how many atoms are there in 5.5 mol of atoms?
   b) If one mole of a gas has a volume of 22.4 L, how many moles are there in 25.0 L of gas?
   c) If one mole of nitrogen has a mass of 28 g, how many moles of nitrogen gas are in 7.0 g of nitrogen gas?
   d) How many seconds must an electrical current of 35 coulombs/s flow in order to deliver 200.0 coulombs?
   e) A quiet sound exerts a pressure of 4×10^{-8} kPa ("kPa" = kilopascals, an SI pressure unit). What is this pressure in atmospheres if 1 atmosphere is 101.3 kPa?
   f) A large nugget of naturally occurring silver metal has a mass of 3.20×10^{4} troy ounces. What is the mass in kilograms if 1 troy ounce is equivalent to 0.0311 kg?
   g) A reaction is essentially complete in 5.0×10^{-4} s. If one millisecond (1 ms) equals 10^{-3} s, how many milliseconds does the reaction take?
   h) If 1 mol of octane produces 5450 kJ of heat when burned, how many moles of octane must be burned to produce 15 100 kJ of heat?
   i) Our fingers can detect a movement of 0.05 micron. If 1 micron is 10^{-3} mm, what is this movement expressed in millimetres (mm)?
   j) If concentrated hydrochloric acid has a concentration of 11.7 mol/L, what volume of hydrochloric acid is required in order to have 0.0358 mol of hydrochloric acid?

MULTIPLE UNIT CONVERSIONS

So far, hopefully, so good. All of the problems above involve a single conversion factor, which leads to the question "What happens when there is ***more than one*** conversion factor involved in a problem?" In fact, you have already run into such problems in everyday life if you have ever tried to solve a problem such as "How many seconds are there in 1 day?" Consider the following examples.

EXAMPLES: **(a) If eggs are $1.44/doz, and if there are 12 eggs/doz, how many individual eggs can be bought for $4.32?**

Analyzing this problem —

The UNKNOWN AMOUNT is "how many individual eggs can be bought".

The INITIAL AMOUNT is $4.32.

There are two conversion statements: "eggs are $1.44/doz", and "there are 12 eggs/doz".

The overall connection which is required is **($) ⟶ (eggs).**

The first conversion statement, $1.44 = 1 doz, makes the connection

($) ⟶ (doz).

The second conversion statement, 12 eggs = 1 doz, makes the connection

(doz) ⟶ (eggs).

Combining the conversion statements gives the overall connection

($) ⟶ (doz) ⟶ (eggs)

which is the connection required (in bold, above).

To start, set up the problem as usual.

$$\# \text{ of eggs} = \$4.32$$

Now, apply the first conversion factor, which cancels the unit "\$".

$$\# \text{ of eggs} = \cancel{\$}4.32 \times \frac{1 \text{ doz}}{\cancel{\$}1.44}$$

So far, cancelling the unit "\$" on the right side leaves the unit "doz". The unit change (\$) ⟶ (doz) is accomplished. Now apply the second conversion factor, which cancels the unit "doz" and accomplishes the unit change (doz) ⟶ (eggs).

$$\# \text{ of eggs} = \cancel{\$}4.32 \times \frac{1 \cancel{\text{doz}}}{\cancel{\$}1.44} \times \frac{12 \text{ eggs}}{1 \cancel{\text{doz}}} = \mathbf{36 \text{ eggs}}$$

Notice that both the units "\$" and "doz" are cancelled.

(b) The automobile gas tank of a Canadian tourist holds 39.5 L of gas. If 1 L of gas is equal to 0.264 gal in the United States ("gal" is the symbol for "gallon", a measure of volume used in the U.S.), and gas is \$1.26/gal in Dallas, Texas, how much will it cost the tourist to fill his gas tank in Dallas?

UNKNOWN AMOUNT = # of dollars

INITIAL AMOUNT = 39.5 L

Required connection: (L) ⟶ (\$)

Conversion statements available: 1 L = 0.264 gal and 1 gal = \$1.26

Connections available through the conversion statements:

(L) ⟶ (gal) and (gal) ⟶ (\$)

Using the conversion statements together gives the required overall connection.

(L) ⟶ (gal) ⟶ (\$)

Using both conversion statements solves the problem. One statement, 1 L = 0.264 gal, allows the cancelling of the initial unit, "L". The other statement, 1 gal = \$1.26, allows the cancelling of the unit "gal" which was introduced by the first conversion factor.

$$\# \text{ of dollars} = 39.5 \cancel{\text{L}} \times \frac{0.264 \cancel{\text{gal}}}{1 \cancel{\text{L}}} \times \frac{\$1.26}{1 \cancel{\text{gal}}} = \mathbf{\$13.1}$$

At the end, the units "L" and "gal" have been cancelled, leaving the required unit, "\$".

EXERCISES:

3. An old barometer hanging on the wall of a mountain hut has a reading of 27.0 inches of mercury. If 1 inch of mercury equals 0.0334 atm ("atmospheres") and 1 atm = 101.3 kPa ("kilopascals"), what is the pressure reading of the barometer, in kilopascals?
4. It requires 334 kJ of heat to melt 1 kg of ice.
   (a) The largest known iceberg had a volume of about 3.1×10^{13} m^3. How much heat was required to melt the iceberg if 1 m^3 of ice has a mass of 917 kg?
   (b) The explosive "TNT" releases 1.51×10^4 kJ of energy for every kilogram of TNT which explodes. Provided that all the energy of an explosion went into melting the ice, how many kilograms of TNT would be needed to melt the iceberg in part (a) of this question?
5. Sugar costs \$0.980/kg. 1 t = 1000 kg. How many tonnes ("t") of sugar can you buy for \$350?
6. The Cullinan diamond, the largest diamond ever found, had an uncut volume of 177 mL. If 1 mL of diamond has a mass of 3.51 g and 1 carat = 0.200 g, how many carats was the Cullinan diamond?

7. How many kilometres ("km") will a car travelling at 120 km/h go in: (a) 0.25 h? (b) 12 min?
8. Solve the following, using the fact that beakers cost $8.40 per dozen.
   (a) Harry drops 3 dozen beakers. How much will the Chemistry teacher charge Harry?
   (b) Harry drops another 5 dozen beakers (clumsy!). If Burger Bob's hamburgers cost $1.50 each, how many hamburgers could clumsy Harry have bought for the same amount of money as he has to pay for the second batch of beakers?
   (c) Harry does not learn very quickly, and breaks a third batch of beakers. If he has to pay $13.30, what is the number of beakers he breaks the third time? (Express your answer in actual numbers of beakers, rather than in "dozens of beakers".)
9. An ancient Celtic chicken farmer wished to purchase a gift for his wife. The gift was worth 2 horses. At the local market, 3 horses were worth 5 cows, 1 cow was worth 4 hogs, 3 hogs were worth 4 goats, and 1 goat cost 9 chickens. How much was the gift going to cost the farmer, who had to pay in chickens?
10. If 1 yard = 3 feet, 1 foot = 12 inches and 1 centimetre = 0.3937 inch, how many centimetres are there in 5 yards?

In addition to the above, there is a specialized type of unit conversion which you must be able to perform: METRIC CONVERSIONS. Before starting on these conversions, let's review metric usage.

II.2. SI UNITS

The International System (SI) of metric units has numerous "base units", although only a few are used in Chemistry 11. A "base unit" is a basic unit of measurement; all other units are multiples of the base units, or combinations of base units.

A. SOME SELECTED BASE UNITS IN THE INTERNATIONAL SYSTEM (SI)

Quantity	Written Unit	Unit Symbol
length	metre	m
mass	gram *	g *
time	second	s
amount of substance	mole	mol

temperature
waves
volume
energy
power

* The actual base unit for mass in the SI system is the kilogram (kg), which is an inconsistent base unit, but for the purposes of Chemistry 11 the gram (g) is considered to be the base unit.

B. SOME ADDITIONAL UNITS USED

Quantity	Written Unit	Unit Symbol
volume	litre	L
mass	tonne	t

C. MULTIPLES OF BASE UNITS

Written Prefix	Prefix symbol	Equivalent exponential
mega	M	10^{6}
* kilo	k	10^{3}
deci	d	10^{-1}
* centi	c	10^{-2}
* milli	m	10^{-3}
micro	μ	10^{-6}

(The prefixes preceded by a "*" are those used most frequently in Chemistry 11.)

D. SOME IMPORTANT EQUIVALENCES

1 mL = 1 cm^3
1 m^3 = 10^3 L
1 t = 10^3 kg

EXAMPLES: **(a) Re–write the expression "5 kilograms" using**
- **PREFIX and UNIT SYMBOLS, and**
- **an EXPONENTIAL EQUIVALENT.**

The ***prefix symbol*** which stands for "kilo" is "k" and the ***unit symbol*** which stands for "grams" is "g".

Therefore: 5 kilograms = 5 kg.

The ***exponential equivalent*** of "kilo" and "k" is "10^3"

Therefore: 5 kilograms = 5 x 10^3 g.

(b) Re–write the expression "2 ms" using
- **a WRITTEN PREFIX and UNIT, and**
- **an EXPONENTIAL EQUIVALENT.**

The ***written prefix*** which is equivalent to "m" is "milli" and the ***written unit*** which is equivalent to "s" is "seconds.

Therefore: 2 ms = 2 milliseconds.

The ***exponential equivalent*** of "milli" and "m" is "10^{-3}"

Therefore: 2 ms = 2 x 10^{-3} s.

(c) Re–write the expression "2.7 x 10^{-2} m" using
- **a WRITTEN PREFIX and UNIT, and**
- **a PREFIX SYMBOL .**

The ***written prefix*** equivalent to "10^{-2}" is "centi", and the ***written unit*** which is equivalent to "m" is "metres".

Therefore: 2.7 x 10^{-2} m = 2.7 centimetres.

The ***prefix symbol*** which stands for "10^{-2}" is "c"

Therefore: 2.7 x 10^{-2} m = 2.7 cm.

The following multiples are used very infrequently and do not have to be memorized. They are only included for the purpose of completeness. (Like, you never can tell when they might come in handy.)

Written Prefix	Prefix symbol	Equivalent exponential
yotta	Y	10^{24}
zetta	Z	10^{21}
exa	E	10^{18}
peta	P	10^{15}
tera	T	10^{12}
giga	G	10^{9}
hecto	h	10^{2}
deka	da	10^{1}
nano	n	10^{-9}
pico	p	10^{-12}
femto	f	10^{-15}
atto	a	10^{-18}
zepto	z	10^{-21}
yocto	y	10^{-24}

EXERCISES:

11. Re–write the following using PREFIX and UNIT SYMBOLS, and EXPONENTIAL EQUIVALENTS.
    (a) 2.5 centimetres (b) 1.3 kilograms (c) 25.2 millimoles (d) 5.1 decigrams (e) 0.25 megalitres (f) 6.38 micrograms
12. Re–write the following using WRITTEN PREFIXES and UNITS, and EXPONENTIAL EQUIVALENTS.
    (a) 2.5 mm (b) 6.5 dL (c) 1.9 kmol (d) 4 Mt (e) 9.94 cg (f) 1.25 μs
13. Re–write the following using PREFIX SYMBOLS, and WRITTEN PREFIXES and UNITS.
    (a) 4.5×10^{-3} mol (b) 1.6×10^{3} m (c) 0.50×10^{-6} L (d) 2.68×10^{-1} g (e) 8.85×10^{6} t (f) 7.25×10^{-2} m
14. Express (a) 50 cm^3 in millilitres (b) 22.5 t in kilograms (c) 0.125 m^3 in litres

II.3. METRIC CONVERSIONS

Metric conversions involve using unit conversions between prefix symbols and exponential equivalents.

EXAMPLES: (a) Write a conversion statement between **cm** and **m.**

Since "c" stands for "10^{-2}" then 1 cm = 10^{-2} m.

(b) Write a conversion statement between **ms** and **s.**

Since "m" stands for "10^{-3}" then 1 ms = 10^{-3} s.

EXERCISE:

15. Write conversion statements between each of the following.
    (a) kg and g (b) Mm and m (c) μL and L (d) dm and m (e) cs and s (f) mmol and mol (g) kL and L (h) μs and s (i) Mg and g (j) cL and L (k) dmol and mol (l) mg and g

EXAMPLE: **How many micrometres are there in 5 cm?**

Unknown amount = # of μm

Initial amount = 5 cm

You can write your own conversion statements between μm and m, and cm and m because the prefixes ***micro*** (μ) and ***centi*** (c) are mentioned in the problem statement.

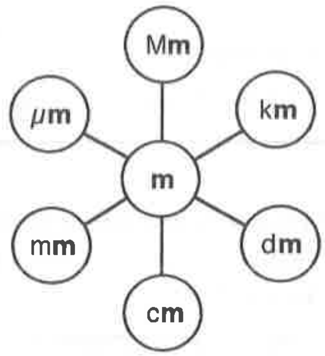
1 μm = 10^{-6} m These statements can be combined to make the connections below.

1 cm = 10^{-2} m (μm) ⟶ (m) ⟶ (cm)

The conversion is now straightforward.

$$\#\text{ of }\mu\text{m} = 5\text{ cm} \times \frac{10^{-2}\text{ m}}{1\text{ cm}} \times \frac{1\ \mu\text{m}}{10^{-6}\text{ m}} = \mathbf{5 \times 10^{4}\ \mu m}$$

The diagram below shows the manner in which a given base unit (for example, meters) is related to the important prefix symbols.



As can be seen, all the prefix symbols are directly related to the "central" base unit. (The central unit "m" also could be any other base unit such as g, s or mol.) In order to connect any two metric prefixes, first connect the initial prefix symbol to the base unit and then connect the base unit to the prefix symbol of the unknown.

EXAMPLE: **Express 8 kg in milligrams.**

Unknown amount = # of mg

Initial amount = 8 kg

Since the prefix symbols "k" (kilo) and "m" (milli) are mentioned in the problem statement, write down the conversion statements.

$1\text{ kg} = 10^3\text{ g}$ These statements can be combined to make the connections below.
$1\text{ mg} = 10^{-3}\text{ g}$ (kg) ⟶ (g) ⟶ (mg)

Carry out the conversion.

$$\#\text{ of mg} = 8\text{ kg} \times \frac{10^3\text{ g}}{1\text{ kg}} \times \frac{1\text{ mg}}{10^{-3}\text{ g}} = \mathbf{8 \times 10^6\ mg}$$

EXAMPLE: **Express 5 Mg/mL in kilograms/litre.**

Unknown amount = # of $\frac{\text{kg}}{\text{L}}$

Initial amount = $\frac{5\text{ Mg}}{\text{mL}}$

This problem requires the conversion of both the numerator and the denominator. Again, write conversion statements for each metric prefix mentioned.

$1\text{ kg} = 10^3\text{ g}$
$1\text{ Mg} = 10^6\text{ g}$
$1\text{ mL} = 10^{-3}\text{ L}$

Treat the top and bottom of the initial fraction separately.

- to convert the top: (Mg) ⟶ (g) ⟶ (kg)
- to convert the bottom: (mL) ⟶ (L)

Depending on how comfortable you are with conversion factors, you can carry out the overall conversion in two ways.

a) ***The 3–step method*:**

1st: Convert the top (ignoring the bottom).

$$\# \text{ of kg} = 5 \text{ Mg} \times \frac{10^6 \text{ g}}{1 \text{ Mg}} \times \frac{1 \text{ kg}}{10^3 \text{ g}} = 5 \times 10^3 \text{ kg}$$

2nd: Convert the bottom (ignoring the top).

$$\# \text{ of L} = 1 \text{ mL} \times \frac{10^{-3} \text{ L}}{1 \text{ mL}} = 1 \times 10^{-3} \text{ L}$$

3rd: Divide the converted top amount by the converted bottom amount.

$$\# \text{ of } \frac{\text{kg}}{\text{L}} = \frac{5 \times 10^3 \text{ kg}}{1 \times 10^{-3} \text{ L}} = 5 \times 10^6 \frac{\text{kg}}{\text{L}}$$

b) ***The 1–step method*:** Simply convert the top and then the bottom (or vice versa), applying all the conversion factors one after another.

- arbitrarily, first convert **Mg** to **g** : $\# \text{ of } \frac{\text{kg}}{\text{L}} = \frac{5 \text{ Mg}}{1 \text{ mL}} \times \frac{10^6 \text{ g}}{1 \text{ Mg}}$
- then immediately convert **g** to **kg** : $\# \text{ of } \frac{\text{kg}}{\text{L}} = \frac{5 \text{ Mg}}{1 \text{ mL}} \times \frac{10^6 \text{ g}}{1 \text{ Mg}} \times \frac{1 \text{ kg}}{10^3 \text{ g}}$
- then convert the BOTTOM by cancelling the **mL** on the bottom of the initial amount with **mL** in the top of a final conversion factor

$$\# \text{ of } \frac{\text{kg}}{\text{L}} = \frac{5 \text{ Mg}}{1 \text{ mL}} \times \frac{10^6 \text{ g}}{1 \text{ Mg}} \times \frac{1 \text{ kg}}{10^3 \text{ g}} \times \frac{1 \text{ mL}}{10^{-3} \text{ L}} = \mathbf{5 \times 10^6 \frac{kg}{L}}$$

Sneaky Short–Cut

Situations which simply require changing from prefix symbol form to base unit form can use a direct substitution of the exponential equivalent for the prefix symbol. This procedure helps to eliminate writing one conversion factor in longer problems.

EXAMPLE: If aluminum is worth \$0.00116/g, what is the cost of 725 kg of aluminum?

$$\# \text{ of dollars} = 725 \times 10^3 \text{ g} \times \frac{\$0.00116}{\text{g}} = \mathbf{\$870}$$

"725 kg" is simply written as "725 x 10^3 g", eliminating the conversion factor $\frac{10^3 \text{ g}}{1 \text{ kg}}$.

EXERCISES:

16. (a) If 1 mg = 10^{-3} g and 1 Mg = 10^6 g, how many milligrams are there in 0.25 Mg?
(b) If 1 μs = 10^{-6} s and 1 cs = 10^{-2} s, how many centiseconds are there in 10 μs?
(c) If 1 mm = 10^{-3} m and 1 cm = 10^{-2} m, how many millimetres are there in 15.8 cm?
(d) If 1 kg = 10^3 g and 1 mg = 10^{-3} g, how many kilograms are there in 250 mg?
(e) If 1 dL = 10^{-1} L and 1 kL = 10^3 L, how many decilitres are there in 0.5 kL?

17. Convert the following

(a) 3 s into milliseconds
(b) 50.0 mL into litres
(c) 2 L into microlitres
(d) 25 kg into grams
(e) 3 Mm into metres
(f) 2 L into decilitres
(g) 7 μs into milliseconds
(h) 51 kg into milligrams
(i) 3125 μL into kilolitres
(j) 1.7 μg into centigrams
(k) 1 year into seconds
(l) 1 mg/dL into grams per litre
(m) 1 cm/μs into kilometres/second
(n) 1 cg/mL into decigrams/litre
(o) 5 cg/ds into milligrams/second

18. Light travels at a rate of 3.00×10^8 m/s.
    (a) It takes light 8.3 min to travel from the surface of the sun to the earth. What is the distance of the earth from the sun?
    (b) The moon is 3.8×10^5 km from the earth. What time will pass between the instant an astronaut on the moon speaks and the instant his voice is heard on earth? (His voice travels by modulated laser beam at the speed of light.)
    (c) A robot vehicle is travelling on the surface of Mars while Mars and Earth are at their closest approach (7.83×10^7 km). Suddenly, a video camera on the robot shows a yawning crevasse dead ahead! How many minutes will it take for an electronic signal travelling at the speed of light to go from Earth to Mars in order to tell the robot to stop immediately?

19. (Care: Nasty!) A measurement is given as 9.0 lb/in^3. If 1 kg = 2.2 lb and 1 m = 39 in, convert the measurement into kg/m^3.

OPTIONAL EXERCISES:

20. If sugar is $9.80 for 10 kg, what is the cost of: (a) 90.0 kg of sugar? (b) 6.00 tonnes of sugar?

21. If 1 inch = 2.54 cm, what is the length, in centimetres, of a 20.0 inch rod? What is the length, in metres, of a 36 inch ruler?

22. Express 90 μg in centigrams.

23. A car travels at a constant speed of 105 km/h.
    (a) How many hours does it take to go 450 km?
    (b) How many seconds does it take to go 2.0×10^2 m?
    (c) How many kilometres are travelled in 10.0 min?
    (d) How many centimetres are travelled in 1.00 ms?

24. If 1 L of granite has a mass of 5.50 kg,
    (a) what is the mass of 7.00 L of granite?
    (b) what is the volume occupied by 22 kg of granite?
    (c) what is the mass, in grams, of 5.00 mL of granite?

25. The SI unit of energy is the joule (unit symbol = J). If 0.334 kJ of energy is required to melt 1.00 g of ice and 1 kJ = 1000 J then:
    (a) what mass of ice can be melted by 10.0 kJ of heat?
    (b) how many kilojoules of heat are required to melt 50.0 g of ice?
    (c) how many joules of heat are required to melt 2.00 kg of ice?

26. Express 80.0 Mg in micrograms.

27. Express 2 cL/ms in kilolitres/second.

28. Express 50.0 mL/min in microlitres/second.

DERIVED QUANTITIES

Definitions: A **DERIVED QUANTITY** is a number made by combining two or more other values.

A **DERIVED UNIT** is a unit which is made by combining two or more other units.

EXAMPLE: The heat change occurring when the temperature of a water sample increases is given by

$$\Delta H = c \cdot m \cdot \Delta T$$

where: ΔH = the change in heat ; "Δ" is the Greek letter "delta" and is used to indicate "the change in" (ΔH is measured in joules, **J**).

m = the mass of water being heated (measured in grams, **g**),

ΔT = the temperature change of the water (measured in degrees Celsius, **°C**)

and c = a derived quantity called the specific heat capacity, which can be calculated by rearranging the above equation.

$$c = \frac{\Delta H}{m \cdot \Delta T}$$

The units of **c** are derived by substituting the units of each symbol into the equation. For example, using the values: $\Delta H = 4.02 \times 10^4$ J, $m = 175$ g and $\Delta T = 55.0^\circ$C gives

$$c = \frac{4.02 \times 10^4 \text{J}}{175\,\text{g} \times 55.0^\circ\text{C}} = 4.18\,\frac{\text{J}}{\text{g} \cdot {}^\circ\text{C}}.$$

Therefore, **c** is a **derived quantity**, having **derived units**, found by combining three other quantities (ΔH, m and ΔT) and their units.

EXERCISE:

29. Find the derived value and units for

(a) the molar concentration, **c**, using the equation $\boldsymbol{c = \frac{n}{V}}$,

where: $\boldsymbol{n} = 0.250$ mol and $\boldsymbol{V} = 0.500$ L.

(b) the Universal Gas Constant, **R**, using the equation $\boldsymbol{R = \frac{P \cdot V}{n \cdot T}}$,

i) where $\boldsymbol{P} = 1$ atm, $\boldsymbol{V} = 22.4$ L, $\boldsymbol{n} = 1$ mol and $\boldsymbol{T} = 273$ K (K is the temperature on the Kelvin scale.

ii) where $\boldsymbol{P} = 202.6$ kPa, $\boldsymbol{V} = 24.45$ L, $\boldsymbol{n} = 2$ mol and $\boldsymbol{T} = 298$ K.

(c) the entropy change for the boiling of water , **ΔS**, using the equation $\boldsymbol{\Delta H = T \cdot \Delta S}$, where: $\boldsymbol{\Delta H} = 44.0$ kJ and $\boldsymbol{T} = 373$ K. (Hint: you will have to rearrange the equation first.)

(d) the kinetic energy of hydrogen gas at 0°C, **KE**, using the equation $\boldsymbol{KE = \frac{1}{2}\, m \cdot v^2}$,

where: $\boldsymbol{m} = 3.35 \times 10^{-27}$ kg and $\boldsymbol{v} = 1692\,\frac{\text{m}}{\text{s}}$.

II.4. DENSITY

Definitions: **Mass** = the quantity of matter in an object

Density = the mass contained in a given volume of a substance

In other words, density is mass divided by volume.

$$\boxed{d = \frac{m}{V}}$$

where: d = the density
m = the mass
V = the volume

EXERCISE:

30. If "m" is measured in grams, and "V" is measured in litres, what are the units of "d"?

Density calculations involve direct substitution of information into the density equation, after rearranging the equation to solve for the unknown.

EXAMPLES: **(a) An iron bar has a mass of 19 600 g and a volume of 2.50 L. What is the iron's density?**

Substitute the given values into the density equation.

$$d = \frac{m}{V} = \frac{19\,600\ \text{g}}{2.50\ \text{L}} = \mathbf{7.84 \times 10^3\ g/L}$$

(b) If mercury has a density of 13 600 g/L, what volume (in millilitres) is occupied by 425 g of mercury?

First, rearrange the density equation to solve for V

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

and then substitute the given information.

$$V = \frac{m}{d} = \frac{425\ \text{g}}{13\,600\ \text{g/L}} = 0.0313\ \text{L}$$

Finally, perform a unit conversion to express the answer in millilitres.

$$V = 0.0313\ \text{L} \times \frac{1\ \text{mL}}{10^{-3}\ \text{L}} = \mathbf{31.3\ mL}$$

Note: You could also start the problem by using a unit conversion – a density is actually a conversion factor: 1 L = 13 600 g.

$$\#\text{ of litres} = 425\ \text{g} \times \frac{1\ \text{L}}{13\,600\ \text{g}} = 0.0313\ \text{L}$$

IMPORTANT FACT:

For water at 4°C –
d = 1000.0 g/L
or **d = 1.0000 g/mL**

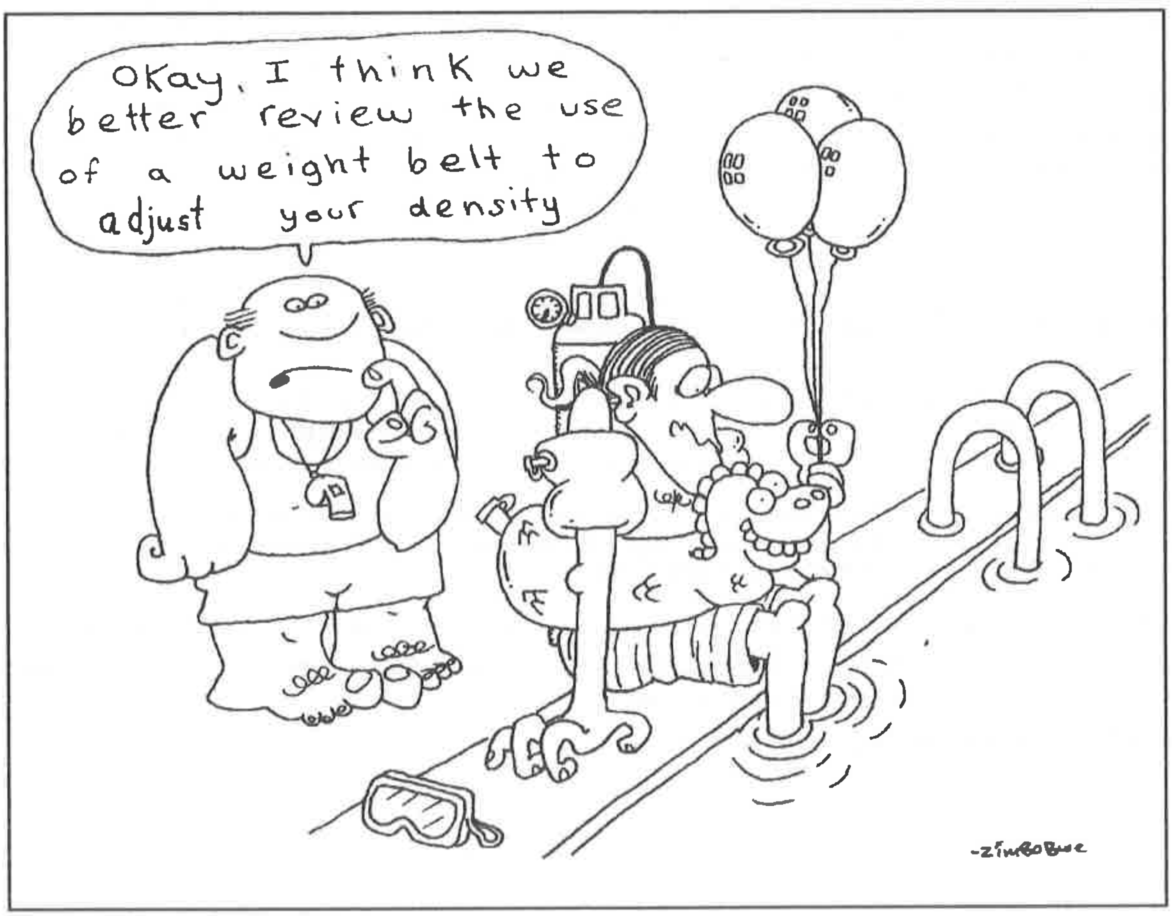
Note that measuring the volume of a sample of water allows you to immediately know its mass, and vice versa. Density can be translated into a conversion statement.

1 g = 1 mL (this is **ONLY** TRUE for water at 4°C, but is quite close for other temperatures)

For example, 50 mL of water will have a mass of 50 g.

Another important fact is: **LESS DENSE LIQUIDS AND OBJECTS FLOAT ON LIQUIDS HAVING A GREATER DENSITY** (in other words, corks float in water and rocks don't). This fact is summarized below.

Objects will sink in a liquid if $d_{OBJECT} > d_{LIQUID}$.
Objects will float in a liquid if $d_{OBJECT} < d_{LIQUID}$.



EXERCISES:

31. A 3.50 mL chunk of boron has a mass of 8.19 g. What is the density of the boron?
32. An iron bar has a mass of 125 g. If iron's density is 7.86×10^3 g/L, what volume does the bar occupy ?
33. A block of beeswax has a volume of 200.0 mL and a density of 961 g/L. What is the mass of the block?
34. Alcohol has a density of 789 g/L. What volume of alcohol is required in order to have 46 g of alcohol?
35. A gas called neon is contained in a glass bulb having a volume of 22.4 L. If the density of the neon is 0.900 g/L, what is the mass of the neon in the bulb?
36. A 70.0 g sphere of manganese (density = 7.20×10^3 g/L) is dropped into a graduated cylinder containing 54.0 mL of water. What will be the water level indicated after the sphere is inserted?
37. A 25.0 mL portion of each of W, X, Y and Z is poured into a 100 mL graduated cylinder. Each of the 4 compounds is a liquid and will not dissolve in the others. If 55.0 mL of W have a mass of 107.3 g, 12.0 mL of X have a mass of 51.8 g, 42.5 mL of Y have a mass of 46.8 g and 115.0 mL of Z have a mass of 74.8 g, list the layers in the cylinder from top to bottom.
38. Explain why boats made of iron are able to float. The density of iron is 7.86×10^3 g/L.
39. If the density of copper is 8.92×10^3 g/L and the density of magnesium is 1.74×10^3 g/L, what mass of magnesium occupies the same volume as 100.0 g of copper?
40. The sun has a volume of 1.41×10^{30} L, an average density of 1.407 g/mL, and can be thought of as more or less pure hydrogen. If the sun consumes 4.0×10^6 t of hydrogen per second, how many years will it take at this rate to burn all of the hydrogen? Hint: use the results of exercise 17(k). The sun will actually cease burning its hydrogen in far less time than indicated by this simple calculation.
41. (OPTIONAL: A Stinker!) A hollow cylinder, closed at both ends, has a volume of 250.0 mL and contains 4.60 g of argon gas. A 90.0 g cube of sodium (density = 970.0 g/L) is inserted into the tube in such a way that no gas escapes. What is the density of the gas afterwards?

II.5. SIGNIFICANT FIGURES AND EXPERIMENTAL UNCERTAINTY

When **COUNTING** a small number of objects it is not difficult to find the **EXACT** number of objects. On the other hand, when a property such as mass, time, volume or length is **MEASURED** you can ***never*** find the exact value. It is possible to find a mass, say, very precisely but it is ***impossible*** to find an object's exact mass.

All measurements have a certain amount of "uncertainty" associated with them. The purpose of this section is to show you how to correctly report and use the results of the experimental measurements you will be making in Chemistry 11.

You will need to learn
(a) how to find and report the uncertainty associated with each measurement, and
(b) the number of digits which can be claimed when reporting results and carrying out calculations with the results.

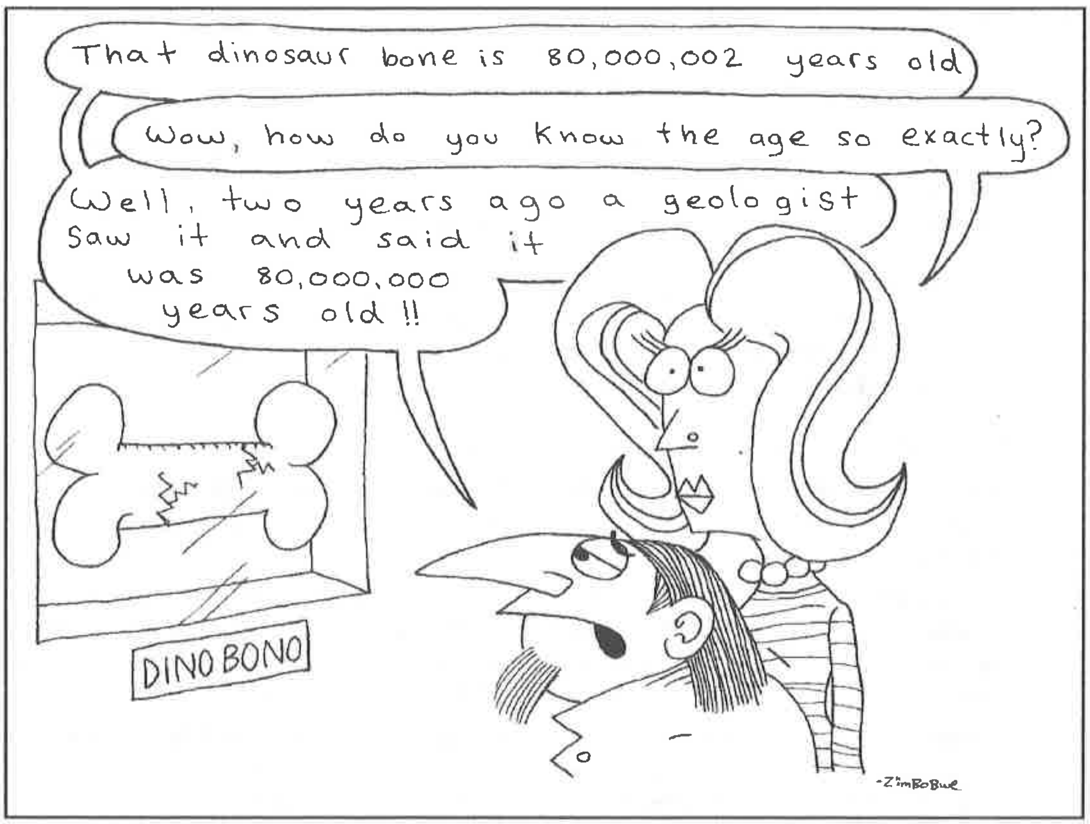
Let's look at the rules of the game.

SIGNIFICANT FIGURES

A. **A significant figure is a measured or meaningful digit.**

EXAMPLE: If a stopwatch is used to time an event and the elapsed time is 35.2 s, then the measurement has 3 significant figures (3, 5 and 2). If the stopwatch can only be read to 0.1 s then it is silly to claim that the time according to the stopwatch is 35.2**168497** s. Since the stopwatch cannot measure the time to 7 decimal places, the last digits (**168497**) have **no significance** – in other words these last digits are "imagined" or a joke.

EXAMPLE: A balance gives a reading of 97.53 g when a beaker is placed on it. This first reading has 4 significant figures since the measurement contains 4 digits. The beaker is then put on a different balance, giving a reading of 97.5295 g. In this second case there are more significant figures to the measurement (6 significant figures).



SPECIAL NOTE: When a ***measurement*** is reported it is usual to assume that in numbers such as

10, 1100, 120, 1000, 12 500

any zeroes at the end are NOT SIGNIFICANT WHEN NO DECIMAL POINT IS SHOWN. That is, we assume the last digits are zeroes because they are rounded off to the nearest 10, 100, 1000, etc. The number of significant digits in the above examples are shown in parentheses, below.

10 (1), 1100 (2), 120 (2), 1000 (1), 12 500 (3)

To complicate matters, SI usage dictates that a decimal point cannot be used without a following digit. For example, 10.0 and 100.0 are legal examples of SI usage with 3 and 4 significant digits respectively, but 10. and 100. are "illegal" ways of showing numbers. If you need to show that a number has been measured to 3 significant figures and has a value of 100 or that the number 1000 has actually been measured to 4 significant figures, the solution is to use exponential notation.

$1.00 \times 10^2 = 100$ (to 3 significant figures)
$1.000 \times 10^3 = 1000$ (to 4 significant figures)

EXERCISE:

42. How many significant figures do each of the following measurements have?

(a) 1.25 kg	(c) 11 s	(e) 1.283 cm	(g) 2 000 000 years
b) 1255 kg	(d) 150 m	(f) 365.249 days	(h) 17.25 L

B. An **ACCURATE** measurement is a measurement that is close to the CORRECT or ACCEPTED value. (The closer to the correct/accepted value, the more accurate the measurement.)

A **PRECISE** measurement is a reproducible measurement. In general, the more precise a measurement, the more SIGNIFICANT DIGITS it has.

Notes: 1. The accuracy of a measuring instrument depends on whether the instrument is properly "calibrated". For example, if a reference mass of 50.000 g is put on an electronic balance and the balance gives a reading of 48.134 g, the balance is not accurate. A special adjustment on the electronic balance is then used to make the balance give a reading of 50.000 g, in agreement with the reference mass. This adjustment process is called a "calibration".

2. For the purposes of Chemistry 11, "high precision" shall be used to mean a "high number of significant figures". In general, it is reasonable to assume that if an instrument in good operating condition can give a reading to 8 significant figures a first time, it will give the same (and therefore reproducible) measurement a second time — provided what is being measured does not change. (There are cases where an instrument can record many digits but not give reproducible results as the result of "machine malfunction" or random errors, but these are ignored for our purposes.)

EXAMPLES: Assume the **CORRECT** width of a room is 5.32000 m.

- A measurement of 5.3 m is **ACCURATE but not VERY PRECISE.** (The value "5.3" is very close to correct as far as its significant figures go, but there are not many significant figures so the value is not very precise)
- If several measurements with some device consistently give the width as 5.45217 m, the measurements are **PRECISE but not ACCURATE.** (Apart from the initial "5" none of the significant digits are correct, so the measurements lack accuracy. The measurements DO have several significant figures, however, so they are precise.)
- If a measurement is consistently given as 5.32001 m, it is **ACCURATE and PRECISE.** (The measured value has many significant figures, so it is precise, and the measured digits agree very well with the correct value, so the measurement is accurate.)
- If a measurement is 7.1 m, it is **not ACCURATE and not PRECISE.** (There are very few significant figures, so the measurement is not very precise, and all the digits are in error so the measurement is not accurate.)

EXERCISES:

43. Assume you have a balance which gives very precise measurements. What might be true about the balance in order that its readings would be precise but not accurate?

44. A "calibration weight" has a mass of exactly 1.000 000 g. A student uses 4 different balances to check the mass of the weight. The results of the weighings are shown below.

    mass using balance A = 0.999 999 g
    mass using balance B = 1.00 g
    mass using balance C = 3.0 g
    mass using balance D = 0.811 592 g

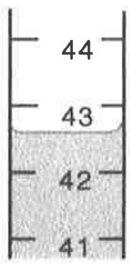
    (a) Which of the balances give accurate weighings?
    (b) Which of the balances give precise weighings?
    (c) Which balance is both accurate and precise?

45. An atomic clock is used to measure a time interval of 121.315 591 s. Assume you have to measure the same time interval. Give an example of a time interval you might actually measure if your measurement is:

    (a) not accurate, but is precise.
    (b) not precise, but is accurate.
    (c) both inaccurate and imprecise.
    (d) both accurate and precise.

C. The **number** of significant figures is equal to all the **certain** digits PLUS the **first uncertain** digit.

EXAMPLE: In the figure at the right, the liquid level is somewhere between 42 and 43 mL. You know that it is at least 42 mL, so you are "certain" about the first two digits. As a guess, the volume is about 42.6 mL; it could be 42.5 or 42.7 but 42.6 seems reasonable. There is some "significance" to this last, guessed digit. It is somewhat uncertain, ***but*** not completely so. For example, the reading is NOT 42.1 or 42.9. As a result, there are two CERTAIN digits (4 and 2) and one uncertain–but–still–significant digit (6) for a total of THREE significant figures.



NOTE: If you are given a measurement without being told something about the device used to obtain the measurement, assume that the LAST DIGIT GIVEN IS SOMEWHAT UNCERTAIN.

EXERCISE:

46. How many "certain" digits are contained in each of the following measurements?
    (a) 45.3 s (b) 125.70 g (c) 1.85 L (d) 2.121 38 g

D. "**Defined**" numbers and "**counting**" numbers are assumed to be **PERFECT** so that they are "exempt" from the rules applying to significant figures.

Defined or counted values involve things which cannot realistically be subdivided and must be taken on an "all–or–nothing" basis.

EXAMPLES: When "1 book" or "4 students" is written, it means exactly "1 book" and "4 students", not 1.06 books and 4.22 students.

The conversion factor 1 kg = 1000 g is used to define an exact relationship between grams and kilograms, so that the numbers involved are assumed to be perfect.

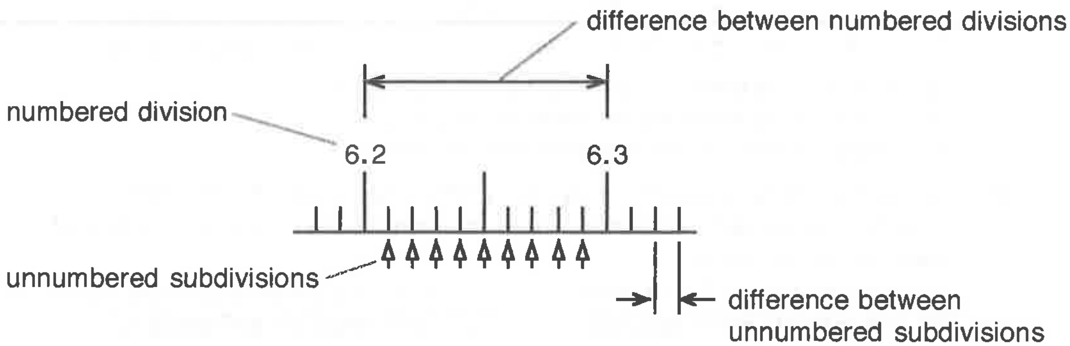
EXERCISE:

47. In the space following each value below put "M" if the value was likely obtained by a **M**easurement, or "C" if the value was probably determined by **C**ounting.
    (a) 4 comets (b) 45 seconds (c) 6.5 litres (d) 12 TV sets (e) 12 grams

HOW TO READ A SCALE

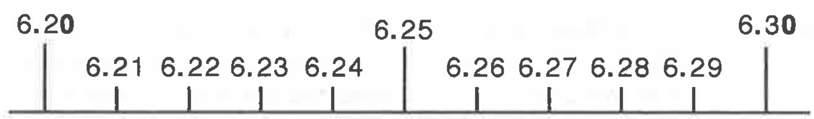
Before learning more about uncertainty, you must first be able to read a scale properly.

IMPORTANT: The following terms are used –



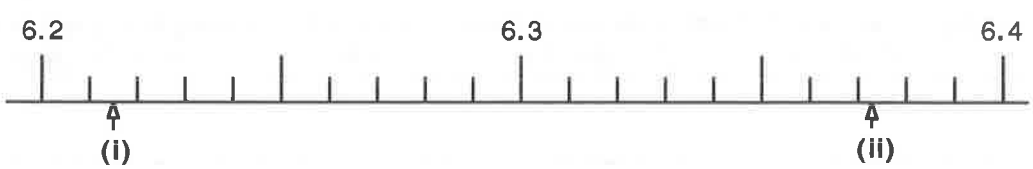
Both the numbered divisions and unnumbered subdivisions are **CALIBRATED** DIVISIONS because the overall scale has been "marked off" or "calibrated" at regular intervals.

If the unnumbered subdivisions ***were numbered***, they would be labelled as shown below.



The "numbered divisions" would then read "6.20" and "6.30", rather than "6.2" and "6.3". The unnumbered subdivisions allow two more decimal places to be read. For example, the numbered divisions above differ in the first decimal place and the unnumbered subdivisions allow a reading to the second decimal place. The estimated distance between unnumbered subdivisions allows a reading to the third decimal place.

EXAMPLES: **(a) What is the value of (i) and (ii) on the following centimetre scale?**



The first two digits of (i) are 6.2 and the first two digits of (ii) are 6.3. The problem is to read the next two digits for each point.

FIRST: Find the difference between each NUMBERED DIVISION.
In the above example: 6.3 – 6.2 = 0.1 cm.

SECOND: Find the number of unnumbered subdivisions between numbered divisions and calculate the value of each unnumbered subdivision. Each numbered division above has 10 subdivisions and each unnumbered sub–division is

$$\frac{0.1\ \text{cm}}{10} = 0.01\ \text{cm}.$$

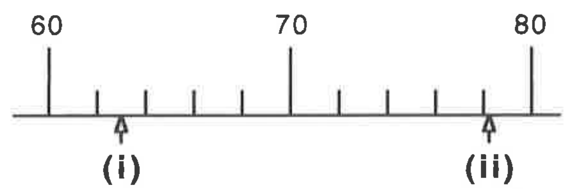
Since the unnumbered subdivisions have a value of 0.01 cm, the value at (i) is a little more than 6.21 cm and the value at (ii) is a little more than 6.37 cm.

THIRD: Estimate how far along their respective unnumbered subdivisions (i) and (ii) are; this gives a reading to the next decimal place, which is the uncertain digit.

Reading at (i): The 3 certain digits are "6.21". The pointer is half–way from 6.21 to 6.22, so the uncertain 4th digit is probably a "5". Therefore, the reading is **6.215 cm.**

Reading at (ii): The 3 certain digits are "6.37". The pointer is $^3/_{10}$ of the way from 6.37 to 6.38, so the uncertain 4th digit is probably a "3". Therefore, the reading is **6.373 cm.**

(b) What is the value of (i) and (ii) on the following centimetre scale?



The value of (i) lies between 60 and 70 cm; the value of (ii) lies between 70 and 80 cm.

FIRST: The difference between numbered divisions is 10 cm.

SECOND: There are 5 subdivisions between each numbered division, so each unnumbered subdivision is equal to

$$\frac{10\,\text{cm}}{5} = 2\,\text{cm}\,.$$

THIRD: Pointer (i) lies between 62 and 64, and pointer (ii) is between 78 and 80.

Reading at (i): The pointer is about half–way ($^5/_{10}$) between 62 and 64. Therefore the reading is more than 62 cm by $^5/_{10}$ of 2 cm (the subdivision value).

reading = 62 cm + 0.5 x 2 cm = **63.0 cm**

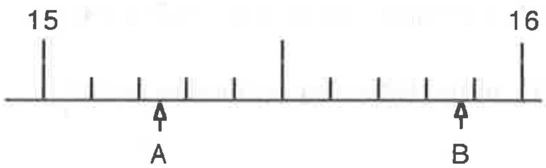
(The numbered divisions differ by "tens", the unnumbered subdivisions are read to the "ones" and an estimate between unnumbered subdivisions are read to "tenths".)

Reading at (ii): The pointer is about $^1/_{10}$ of the way from 78 to 80. Therefore the reading is more than 78 by $^1/_{10}$ of 2 cm.

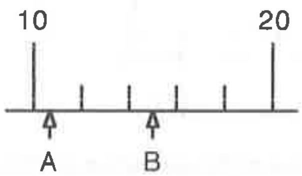
reading = 78 cm + 0.1 x 2 cm = **78.2 cm**

EXERCISE:

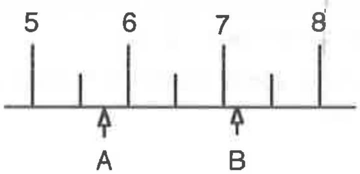
48. In each of the following, determine the reading as follows. ***Note:*** all measurements are in "cm".
    i) Find the difference between each numbered division.
    ii) Find how many unnumbered subdivisions lie between each numbered division and calculate the value of the intervals between each unnumbered subdivision.
    iii) Estimate the value at the pointer (you will have to estimate how far the pointer is from one unnumbered subdivision to the next).

(a)


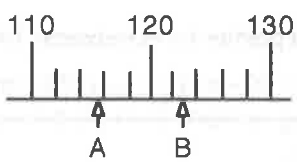
A = ________
B = ________

(b)


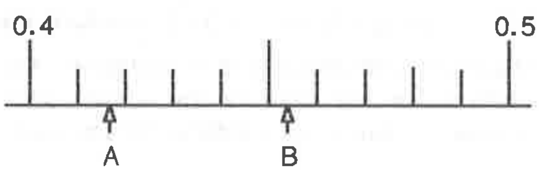
A = ________
B = ________

(c)


A = ________
B = ________

(d)


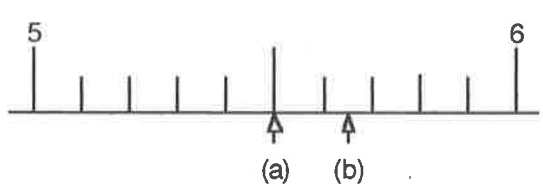
A = ________
B = ________

(e)


A = ________
B = ________

There is one last problem associated with reading scales that must be examined: **what to do when the "pointer" is exactly on one of the markings.**

EXAMPLE: Look at the centimetre ruler and indicated values below.



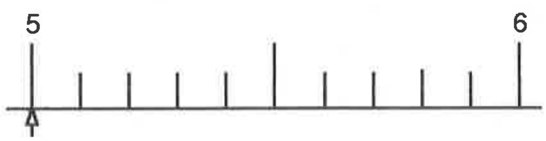
The pointer at (a) seems to indicate a value of 5.5 cm but that is ***not*** the correct value. Look at the pointer at (b). Since the value at (b) is about 5.65, both the value at (a) and (b) can be guessed to the nearest **0.01 cm.** The value for (a) must be given as **5.50 cm.**

BE VERY CAREFUL WHEN A VALUE APPEARS TO COINCIDE EXACTLY WITH A MARKING ON A MEASURING DEVICE. The following procedure should help when such a situation occurs.

THE PROCEDURE FOR CORRECTLY READING MEASURING SCALES WHEN A POINTER IS EXACTLY ON A NUMBERED DIVISION

- Determine the value that the measurement seems to have.
- Pretend you have a value in between two of the unnumbered subdivisions on your measuring device.
- Determine how many decimal places you could read off the measuring device at the "in–between value".
- Add a sufficient number of zeroes to the actual reading to give you the correct number of decimal places for your reading.

In the example above, the intervals between unnumbered subdivisions can be read to 0.01 cm; that is, to 2 decimal places. The reading appears to be 5.5, which is only 1 decimal place, so an extra zero is added to get the value: **5.50 cm.** Similarly, consider the value of the measurement below.

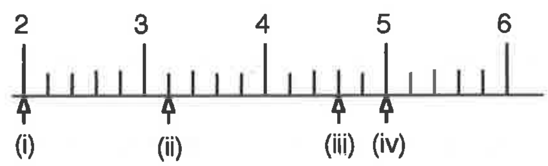


The value seems to be 5 cm, but the previous example shows that an "in–between value" can be read to 0.01 cm (2 decimal places) and so 2 extra zeroes are added to arrive at the final reading: **5.00 cm.**

EXERCISES:

49. Determine the readings on the following centimetre rulers.

(a)



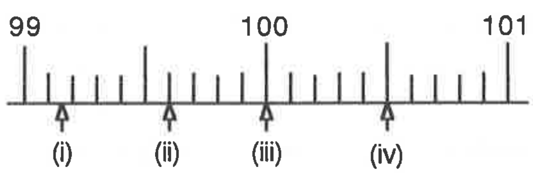
Reading at (i) = ____________

Reading at (ii) = ____________

Reading at (iii) = ____________

Reading at (iv) = ____________

(b)



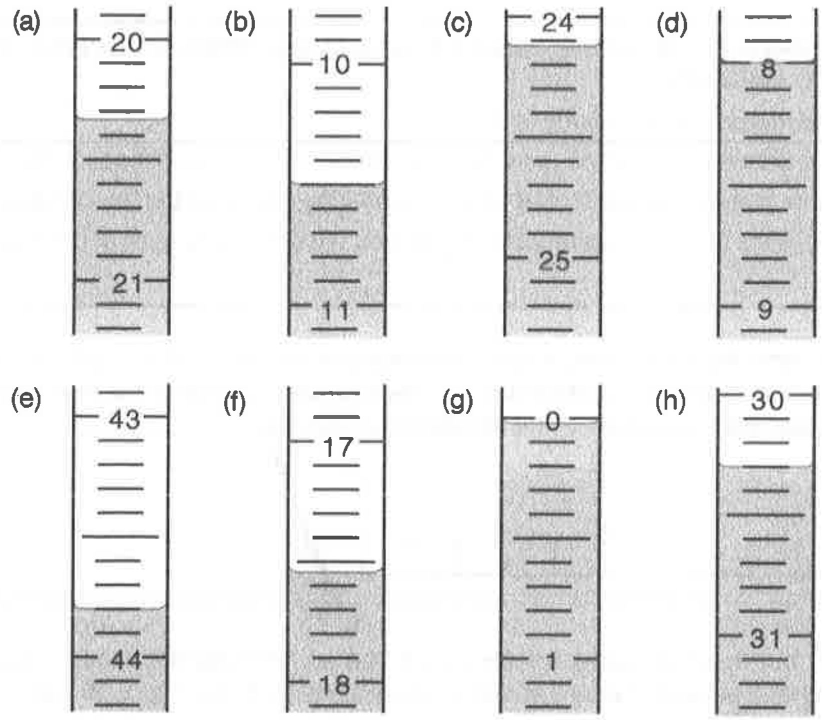
Reading at (i) = ____________

Reading at (ii) = ____________

Reading at (iii) = ____________

Reading at (iv) = ____________

50. Determine the volume readings of the following burettes. **Care!** The numbers ***increase*** going ***down*** the scale.



EXPERIMENTAL UNCERTAINTY

Having seen how to deal with significant figures and make proper readings, the next step is to learn about experimental uncertainty.

Definition: The experimental uncertainty is the estimated amount by which a measurement might be in error.

E. When adding an uncertainty to a measurement, the uncertainty goes after the measured value but before the unit.

EXAMPLE: Assume that a measured temperature is 39.6°C and the uncertainty in the measurement is ± 0.1°C (Part F shows how to estimate the uncertainty). The measurement and uncertainty are shown below.

$$\underbrace{39.}_{\text{certain digits}}\underbrace{6}_{\text{uncertain digit}} \pm \underbrace{0.1}_{\text{uncertainty which is possessed by the uncertain digit}}\underbrace{{}^{\circ}C}_{\text{put the unit at the end}}$$

NOTE: If the uncertain digit is in the first decimal place, the uncertainty will be in the first decimal place also.

INTERPRETATION OF UNCERTAINTIES

When a measurement is said to be 39.6 ± 0.1°C, this implies that the actual value most likely lies in the range from (39.6 – 0.1)°C to (39.6 + 0.1)°C; that is, from 39.5°C to 39.7°C.

Similarly the measurement 15.55 ± 0.02 mL implies a volume in the range from 15.55 – 0.02 = 15.53 mL to 15.55 + 0.02 = 15.57 mL.

If only the range of probable values is known, for example 88.0 g to 89.0 g, the uncertainty is simply one–half of the stated range.

$$\text{range} = 89.0 - 88.0 = 1.0 \text{ g}$$
$$\text{uncertainty} = \tfrac{1}{2}(1.0) = 0.5 \text{ g}$$

The measurement reported is the MIDPOINT of the range, plus/minus the uncertainty. The midpoint of the range is simply the AVERAGE of the endpoints of the range.

$$\text{midpoint} = \tfrac{1}{2}(88.0 + 89.0) = 88.5 \text{ g}$$

Therefore, the reported measurement is **88.5 ± 0.5 g.**

Similarly, if the range is 15.0 g to 15.5 g then the midpoint of the range is 15.3 g (to one decimal point) and the uncertainty is $\tfrac{1}{2}$ (0.5) = 0.3 g (to one decimal point). [Yes, 15.3 ± 0.3 g predicts the range as 15.0 – 15.6 g, but the range, midpoint and uncertainty are all recognized as simply "good guesses".]

IMPORTANT: **The place values (tens, units, first decimal, etc.) of the experimental uncertainty and the first uncertain digit of a measurement must agree with each other.**

EXAMPLES: 15.5°C ± 0.01°C is ***wrong*** because the measurement is only read to the nearest 0.1°C, which means the first decimal place is uncertain. An uncertainty of 0.01°C implies the measurement can be read with at least partial certainty to the second decimal place.

5.52 ± 0.01 mL is ***correct*** because the last (uncertain) digit in the measurement and the uncertainty quoted are both to the second decimal place.

EXERCISES:

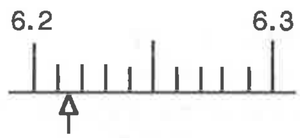
51. Write the following measurements and uncertainties in the correct form.
    (a) A balance gives a mass reading of 51.32 g. The balance has an uncertainty of 0.01 g.
    (b) A student records a volume of 55 mL with an uncertainty of 1 mL.
    (c) A measurement is in the range from 452 g to 458 g.
    (d) A measurement is in the range from 0.5128 g to 0.5132 g.
    (e) Several people time the same event. Their times range from 98.2 s to 99.5 s.
    (f) A series of white mice have masses ranging from 48.9 g to 50.6 g.

52. What is the range of values possible for the following?
    (a) 15.25 ± 0.01 mL (b) 110.0 ± 0.2 mL (c) $1.528 \times 10^{-6} \pm 0.005 \times 10^{-6}$ s

F. **NORMALLY USE UNCERTAINTIES TO THE NEAREST 0.1 OF THE SMALLEST UNNUMBERED SUBDIVISION.** ***If*** you can only estimate a value to the nearest ±0.2 or even ±0.5 of the smallest unnumbered subdivision, feel free to do so, but be prepared to justify your decision. (Sometimes values are hard to read.)

Now that you know HOW to read a scale, estimating the uncertainty is relatively easy.

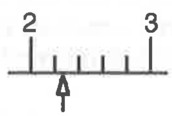
EXAMPLE: Look at the centimetre scale below.



The pointer indicates a value of 6.214, and the last digit ("4") is somewhat uncertain. **The place value (third decimal place) of the experimental uncertainty and the first uncertain digit of a measurement must agree with each other.** Therefore the value and uncertainty are

6.214 ± 0.001 cm .

EXAMPLE: The value on the scale below is 2.26 cm and $^{1}/_{10}$ of an unnumbered subdivision is 0.02 cm.



Therefore, the value and uncertainty are **2.26 ± 0.02 cm .**

EXERCISES:

53. Record values for the experimental uncertainty encountered in using the following apparatus.

Apparatus	Difference between numbered divisions	# of unnumbered subdivisions between numbered divisions	Smallest unnumbered subdivision	Uncertainty of measurement
thermometer				
10 mL graduated cylinder				
25 mL graduated cylinder				
100 mL graduated cylinder				
250 mL graduated cylinder				
50 mL burette				
clock				
balance				
Other:				
Other:				
Other:				

54. Determine the uncertainty for each of the measurements in exercises 48–50, and record the measurement, uncertainty and units in the correct fashion.

G. **Leading** zeroes are **not** significant.

EXAMPLE: The mass "25 g" has 2 significant figures. Using a unit conversion to express 25 g in kilograms gives

$$\# \text{ of kg} = 25 \text{ g} \times \frac{1 \text{ kg}}{10^3 \text{ g}} = 0.025 \text{ kg} .$$

A more precise measurement was not performed so the measurement must still have two significant figures, the 2 and 5. The leading zeroes (in bold) in **0.0**25 kg are NOT SIGNIFICANT. Notice that re–expressing 25 g in megagrams increases the number of leading zeroes — 0.000 025 Mg — but the leading zeroes are not significant.

The number of leading zeroes depends on the size of the unit used to express the measured value, and is ***not related*** to the precision, accuracy or number of significant figures.

H. **Trailing** zeroes are **all** assumed to be significant and must be justified by the precision of the measuring equipment.

EXAMPLE: The zeroes at the end of the following 2 numbers are called "trailing zeroes".

25.00 g	represents the precision of a common lab balance (4 significant figures in this case)
25.000 000 g	represents a highly precise microbalance (8 significant figures)

A balance precise to at least 0.000 001 g is required in order to ensure that the trailing zeroes in 25.000 000 g are zeroes and not some other digits.

(If a balance capable of making a measurement to ± 0.01 g is used and the result is written as "35.6 g", it is not correct to say "*Oh, I forgot to record the second decimal place, it must have been a zero.*" In fact, any reading from "35.60 g" to "35.69 g" was equally possible. Write "35.60 g" only when the balance actually shows "0" in the second decimal place.)

EXCEPTION: Recall that a number which is written without a decimal and has been rounded off to the nearest 100 (say) does not claim that the last zeroes are significant.

Example: In 38 500 g the trailing zeroes are NOT SIGNIFICANT.

To summarize all the comments on leading and trailing zeroes —

There are two ways to count the number of significant figures.

EXPRESS THE NUMBER IN SCIENTIFIC NOTATION AND THEN COUNT ALL THE DIGITS.

Or even simpler:

starting from the left side of the number, ignore all "leading zeroes" and only start counting at the first NON–ZERO digit. Once you start counting, continue until you run out of digits.

EXAMPLE: 0.000 035 000 = 3.5000×10^{-5}, and therefore the number has 5 significant figures.

EXERCISE:

55. State the number of significant figures in each of the following.

(a) 3570	(c) 41.400	(e) 0.000 572	(g) 41.50×10^{-4}	(i) $1.234\ 00 \times 10^{8}$
(b) 17.505	(d) 0.51	(f) 0.009 00	(h) $0.007\ 160 \times 10^{5}$	(j) $0.000\ 410\ 0 \times 10^{7}$

I. After **MULTIPLYING** or **DIVIDING** numbers, round off the answer to the **LEAST NUMBER OF SIGNIFICANT FIGURES** contained in the calculation.

EXAMPLE: 2.00 x 3.000 00 = 6.00

2.00	3.000 00	6.00
↑	↑	↑
3 significant figures	6 significant figures	3 significant figures

Since the calculation involves a lower precision number (3 significant figures) and a higher precision number (6 significant figures), the precision of the result is limited by the LEAST precise number involved. The answer has only 3 significant figures.

When multiplying 2 numbers like 5.0×20.0 you must be careful how you write the answer.

$5.0 \times 20.0 = 100$ **IS WRONG!**

In this case a 2 significant figure number (5.0) is being multiplied by a 3 significant figure number (20.0), so the answer is only allowed to have 2 significant figures. Since "100" implies 1 significant figure, you MUST change to EXPONENTIAL FORM to properly show that the answer has 2 significant figures.

$5.0 \times 20.0 = 1.0 \times 10^2$

EXAMPLE: $\dfrac{15.55 \times 0.012}{24.6} = 0.0076$

This example involves numbers with 4, 3 and 2 significant figures. Since the least precise number, 0.012, has only 2 significant figures, the answer is rounded off to 2 significant figures.

EXAMPLE: $\dfrac{2.4000}{8.000} = 0.3000$

If you perform this calculation on your calculator, the result shown will be "0.3". ***BUT,*** the answer must have 4 significant figures, so three ZEROES are added to indicate that the answer is "0.3000" to 4 significant figures.

EXAMPLE: $\dfrac{2.56 \times 10^5}{8.1 \times 10^8} = 3.2 \times 10^{-4}$

The exponential parts of the numbers do not contribute to the number of significant figures. This calculation has a 3 significant figure number, 2.56×10^5, divided by a 2 significant figure number, 8.1×10^8. Putting these numbers into a calculator, and rounding the final answer to 2 significant figures, gives 3.2×10^{-4}.

IMPORTANT: YOU MUST **ALWAYS** PERFORM CALCULATIONS TO THE MAXIMUM NUMBER OF SIGNIFICANT FIGURES ALLOWED BY YOUR CALCULATOR AND **ONLY** YOUR FINAL ANSWER SHOULD BE ROUNDED OFF TO THE CORRECT NUMBER OF SIGNIFICANT FIGURES. ROUNDING OFF **INTERMEDIATE** ANSWERS OFTEN PRODUCES INCORRECT RESULTS.

If you cannot keep all your calculated values in your calculator (or its memory), then always round off intermediate results so as to keep at least ONE "SIGNIFICANT FIGURE" more than you will eventually use in your final result.

EXERCISE:

56. Perform the indicated operations and give the answer to the correct number of significant figures.

(a) 12.5 x 0.50	(e) $(6.40 \times 10^8) \times (5 \times 10^5)$	(i) 4.75 x 5
(b) 0.15 x 0.0016	(f) $4.37 \times 10^3 / 0.008\,560\,0$	(j) 0.000 01 / 0.1000
(c) 40.0 / 30.0000	(g) 51.3 x 3.940	(k) 7.4 / 3
d) 2.5 x 7.500 / 0.150	(h) $0.51 \times 10^{-4} / 6 \times 10^{-7}$	(l) 0.000 43 x 0.005 001

J. After **ADDING** or **SUBTRACTING** numbers, round off the answer to the **LEAST NUMBER OF DECIMAL PLACES** contained in the calculation.

The idea behind this rule is simple. The number with the least number of decimal places is least precise and limits the precision of the final result.

EXAMPLE:

$$\begin{array}{r} 12.56 \text{ cm} \\ +\ 125.8 \text{ cm} \\ \hline 138.36 \text{ cm} \end{array}$$

The second value, 125.8 cm, is only precise to the1st decimal place so the final answer is rounded off to one decimal place: **138.4 cm.**

EXAMPLE:

$$\begin{array}{r} 41.037\,6 \text{ g} \\ -\ 41.037\,584 \text{ g} \\ \hline 0.000\,016 \text{ g} \end{array}$$

Since the least precise number has 4 decimal places, the answer must be rounded to 4 decimal places: **0.0000 g.** This answer is interpreted to mean there is no significant difference between the numbers being subtracted.

EXAMPLE: $1.234 \times 10^6 + 4.568 \times 10^7 = ?$

Since the exponents are different, one of the exponents must be changed to the size of the other. Arbitrarily, change the smaller exponent so that it equals the larger one.

1.234×10^6 becomes **0.1234×10^7**

(Since the exponent becomes one power of 10 larger (10^6 becomes 10^7), the number in front is made one power of 10 smaller to compensate.)

Now the numbers are added.

$$\begin{array}{r} 0.1234 \times 10^7 \\ +\ 4.568 \times 10^7 \\ \hline 4.6914 \times 10^7 \end{array}$$

Since the second number is only known to the third decimal place (in its present exponential form), the answer is rounded to the third decimal place.

Answer = **4.691×10^7**

Note: When an uncertain number is multiplied or divided by an exact (counting) number, the result obeys the rules for adding or subtracting the uncertain number. In other words, the answer is rounded to the same number of decimal places as the uncertain number.

EXAMPLES: The weights of 3 boys are: 51.0 kg, 52.4 kg and 49.8 kg. The average of their weights is

$$\frac{(51.0 + 52.4 + 49.8)}{3} = \mathbf{51.1\ kg.}$$

DO NOT try to restrict the answer to 1 significant figure, because "3 boys" is an exact (counting) number rather than a "1 significant figure" number.

A Canadian nickel has a mass of 4.53 g. The mass of three such nickels is:
3 x 4.53 = 13.59 g (the "3" is exact; both "4.53" and "13.59" have 2 decimal places)

IN SUMMARY

When ***multiplying*** or ***dividing*** two numbers, the result is rounded to the least number of significant figures used in the calculation.

When ***adding*** or ***subtracting*** two numbers, the result is rounded to the least number of decimal places used in the calculation.

EXERCISES:

57. Perform the indicated operations and give the answer to the correct number of significant figures.
    - (a) $15.1 + 75.32$
    - (b) $178.904\ 56 - 125.8055$
    - (c) $4.55 \times 10^{-5} + 3.1 \times 10^{-5}$
    - (d) $0.000\ 159 + 4.0074$
    - (e) $1.805 \times 10^{4} + 5.89 \times 10^{2}$
    - (f) $0.000\ 048\ 1 - 0.000\ 817$
    - (g) $7.819 \times 10^{5} - 8.166 \times 10^{4}$
    - (h) $45.128 + 8.501\ 87 - 89.18$
    - (i) $0.0589 \times 10^{-6} + 7.785 \times 10^{-8}$
    - (j) $89.75 \times 10^{-12} + 6.1157 \times 10^{-9}$
58. Perform the indicated operations and give the answer to the correct number of significant figures.
    - (a) $7.95 + 0.583$
    - (b) $1.99 / 3.1$
    - (c) $4.15 + 1.582 + 0.0588 - 35.5$
    - (d) $1200.0 / 3.0$
    - (e) $5.31 \times 10^{-4} / 3.187 \times 10^{-8}$
    - (f) $45.9 - 15.0025$
    - (g) 375.59×1.5
    - (h) $5.1076 \times 10^{-3} - 1.584 \times 10^{-2} + 2.008 \times 10^{-3}$
    - (i) $1252.7 - 9.4 \times 10^{2}$
    - (j) $0.024\ 00 / 6.000$

Mixed calculations involving the addition, subtraction, multiplication and/or division of uncertain values are treated in a step–by–step manner, as shown in the next example.

EXAMPLE: Perform the indicated operations and give the answer to the correct number of significant figures.

$$50.35 \times 0.106 - 25.37 \times 0.176 = ?$$

First: Evaluate the multiplications (and divisions, if present). All digits given by the calculation are shown, with significant figures shown in bold.

$$50.35 \times 0.106 = \mathbf{5.337}1$$
$$25.37 \times 0.176 = \mathbf{4.465}\ 12$$

Second: Perform all additions and subtractions last. The results of the two multiplications are subtracted from each other. Since both of the multiplications are good to the second decimal place, the final result is rounded to the second decimal place.

$$\mathbf{5.337}1 - \mathbf{4.465}\ 12 = \mathbf{0.87}1\ 92$$

Therefore the answer is rounded to **0.87.**

EXERCISE:

59. In the following mixed calculations perform multiplications and divisions before doing the additions and subtractions. Keep track of the number of significant figures at each stage of a calculation.
    - (a) $25.00 \times 0.1000 - 15.87 \times 0.1036$
    - (b) $35.0 \times 1.525 + 50.0 \times 0.975$
    - (c) $(0.865 - 0.800) \times (1.593 + 9.04)$
    - (d) $\dfrac{(0.3812 + 0.4176)}{(0.0159 - 0.0146)}$
    - (e) $\dfrac{3.65}{0.3354} - \dfrac{6.14}{0.1766}$
    - (f) $\dfrac{5.3 \times 0.1056}{0.1036 - 0.0978}$
    - (g) $(0.341 \times 18.64 - 6.00) \times 3.176$
    - (h) $9.34 \times 0.071\ 46 - 6.88 \times 0.081\ 15$

UNIT III : THE PHYSICAL PROPERTIES AND PHYSICAL CHANGES OF SUBSTANCES

III.1. SOME BASIC DEFINITIONS IN SCIENCE

In this course you will be asked to describe substances in many different ways. To make sure that we are using a common vocabulary, some important terms first must be defined and agreed upon.

Definitions: **QUALITATIVE** information is ***NON–NUMERICAL*** information.

QUANTITATIVE information is ***NUMERICAL*** information.

Example	Qualitative Description	Quantitative Description
Your height	tall, short	5' 10" , 180 cm
Your weight	normal, heavy	110 lb, 123 kg
Chemistry 11 mark	awesome, fail	100%, 55/200

(Qualitative and quantitative information serve different purposes. "The boy is five feet tall" is a quantitative statement which makes no judgement about the boy's height. If the boy was only six years old, the qualitative statement "the boy was extraordinarily tall" would let us know that an unusual situation existed.)

An **OBSERVATION** is ***qualitative*** information collected through the direct use of our senses.

An **INTERPRETATION** (or "inference") is an attempt to put meaning into an observation.

A **DESCRIPTION** is a ***list*** of the properties of something.

DATA is ***quantitative*** information which is experimentally–determined or obtained from references.

An **EXPERIMENT** is a test or a procedure that is carried out in order to discover a result.

A **HYPOTHESIS** is a SINGLE, UNPROVEN assumption or idea which attempts to explain why nature behaves in a specific manner. When initially put forward, hypotheses are tentative but, if they survive testing, eventually gain general acceptance.

A **THEORY** is a set of hypotheses that ties together a large number of observations of the real world into a logically consistent and understandable pattern. In other words, a theory is a TESTED, REFINED and EXPANDED explanation of why nature behaves in a given way.

A **LAW** is a broad generalization or summary statement which describes a large amount of experimental evidence stating how nature behaves when a particular situation occurs.

Some Additional Comments about Hypotheses, Theories and Laws

The following are general characteristics of HYPOTHESES.

- Hypotheses are normally single assumptions.
- Hypotheses are narrow in their scope of explanation.
- When originally proposed, hypotheses are tentative (being based on suggestive but VERY incomplete evidence) but may become generally accepted after more complete testing.

The following are general characteristics of THEORIES.

- Theories are composed of one or more underlying hypotheses.
- Theories are broad in scope and may have subtle implications which are not foreseen when they are proposed because they provide explanations for entire "fields" of related behaviour.

- Theories are sometimes called **models** because they often provide a concrete way to examine, predict and test the workings of nature.
- A theory can't be "proven" but at some point it may have such a tremendous record of explanation and prediction that we place a high probability on its correctness as a model capable of describing reality.
- Theories must be "falsifiable"; that is, they must make ***testable predictions*** about the behaviour of the system under NEW conditions. (If a theory makes no predictions then it is not "wrong", but it **is** discarded as useless.)

The following are general characteristics of LAWS.

- Laws summarize the results of many experiments or observations and state what will happen when a specific situation occurs.
- Laws do NOT try to explain WHY something occurs.
- Laws are NOT "proven theories", as sometimes is erroneously stated. Laws are often stated before any theory exists as to WHY the law is true. In the past, new experiments have occasionally shown a particular law to be invalid, causing an upheaval in scientific thought while new theories were proposed to explain the new observations.

Examples of hypotheses

1. All gases are made up of tiny, fast–moving particles.
2. The tiny particles in a gas transfer some or all of their energy of movement when they collide with one another or the walls of their container.
3. The tiny particles in a gas act like miniature billiard balls and the entire system undergoes no net change in energy when particles collide.

Notice that all the above hypotheses are assumptions about the way gases exist and behave.

Examples of theories

The **KINETIC THEORY OF GASES** states that gases act as they do because they are made up of point–like particles which are constantly moving, colliding and exchanging energy.

The KINETIC THEORY OF GASES is the result of combining the three hypotheses above. Pressure is explained as a result of the collision of gas particles with the walls of the container. Temperature is defined in terms of the average velocity of the moving particles.

The **ARRHENIUS THEORY OF ACIDS AND BASES** is the result of combining the following hypotheses.

- Some substances can break up in water to form positive and negative charged particles called ions.
- Acids are substances which break up in water to form an ion of H^+ and a negative ion.
- Bases are substances which break up in water to form an ion of OH^- and a positive ion.
- The reaction between an acid and a base is the result of the H^+ from the acid combining with the OH^- from the base to produce water.
- When an acid and a base react, the left–over positive and negative ions combine to produce a substance called a salt.

Examples of laws

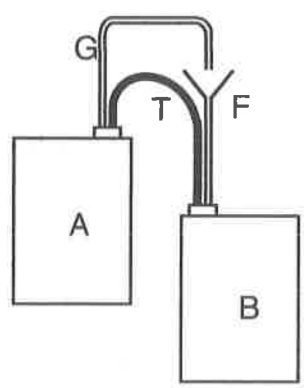
BOYLE'S LAW states that if the temperature is unchanged, then the greater the pressure applied to a sample of gas, the smaller its volume.

CHARLES' LAW states that if the applied pressure is unchanged, then the greater the temperature of a gas sample, the greater its volume.

EXERCISES:

1. Give a quantitative description of (a) a length of time. (b) a temperature.
2. Give a qualitative description of (a) a length of time. (b) a temperature.
3. Which parts of the description in the following passage are quantitative and which are qualitative?

   Copper is a reddish–coloured element with a metallic lustre. It is an excellent conductor of heat and electricity, melts at 1085°C and boils at 2563°C. Archeological evidence shows that it has been mined for the past 5000 years and presently is considered to be one of the most important metals available. Copper is insoluble in water and virtually all other solvents, reacts easily with nitric acid but only slightly with sulphuric and hydrochloric acids. It has a density of 8.92 g/mL, which makes it more dense than iron.
4. "I observed that the long tube had a bright white glow". Give at least two different interpretations that could explain why the tube had a bright white glow. Propose a simple experiment that would help you to decide which of your interpretations is most likely to be correct.
5. What is the difference between
   (a) "observations" and "data"? (c) "observation" and "interpretation"?
   (b) "observation" and "description"?
6. Which of the following can be incorrect?
   (a) data (b) observations (c) interpretations (d) hypotheses (e) theories (f) laws
7. Ideally, experiments should be done to find out WHAT HAPPENS when a procedure is performed, NOT to confirm an expected result. What might happen if an experiment is carried out with the intention of proving that what you think will happen does, in fact, happen?
8. The tiny "protons" inside an atomic nucleus behave in a weird way if a beam of high energy particles is fired at them. One way to explain such behaviour of protons is to assume protons are made up of extremely small particles called "quarks". Is this assumption a hypothesis, theory or law?
9. A person proposes that the world was created 15 minutes ago, in such a way that it looks as if it has been here for billions of years, including a created set of memories which makes us believe we have existed individually for many years. Is this "proposition" an acceptable theory?
10. "Whatever you throw out in the trash today, you will need desperately within four days." Assuming that this whimsical statement is true, is it stated as a hypothesis, a theory, or a law?
11. Explain the important differences between (a) a hypothesis and a theory. (b) a theory and a law.
12. You have two 4 L cans, **A** and **B**, (shown at right). They are sealed with rubber stoppers. A piece of black rubber tubing **T** goes through the stopper in can **A**, over to can **B** and through its stopper. A piece of bent glass tubing **G** goes through the stopper in can **A**. Finally, a glass funnel **F** goes through the stopper in can **B**. The exposed end of **G** is about 3 cm above the funnel top. When about 300 mL of water is poured down the funnel into can **B**, water is seen to rise up glass tube **G**, pour out the tip and into the funnel. At this point no more water is poured into the funnel, but water continues to pour out of tube **G**. After a litre of water has come out, the outpouring still has not slowed down.



   (a) Based on these observations, propose a model for the interior of the cans. Hint: will the water keep flowing forever?
   (b) Apart from actually taking out the stoppers and looking in the cans, what further experiments could you propose to test your model?

III.2. THE PHYSICAL PROPERTIES OF MATTER

What is "Chemistry" anyway? Chemistry is such a broad field that sometimes it is said that "Chemistry is what chemists do". Before giving a better definition for Chemistry, a special term must be defined.

Definition: **MATTER** = anything that has mass and occupies space. (Matter is what makes up the universe, other than energy.)

A better definition of Chemistry can now be given.

CHEMISTRY is the science concerned with the properties, composition and behaviour of matter.

This section extensively examines the physical properties of matter; later sections examine the composition and behaviour of matter.

Definitions: A **SUBSTANCE** is something with a unique and identifiable set of properties.

Therefore, two objects with different properties must be made of different substances.

A **PHYSICAL PROPERTY** of a substance is a property that can be found without creating a new substance.

Example: Density, colour, hardness and melting temperature are physical properties.

A **CHEMICAL PROPERTY** is the ability of a substance to undergo chemical reactions and change into new substances, either by itself or with other substances.

Example: One chemical property of hydrogen is its ability to burn in air and produce water; another is its ability to react with chlorine gas and produce hydrogen chloride.

A physical property can be either **INTENSIVE** or **EXTENSIVE.**

Definitions: An **EXTENSIVE** property of a substance is a physical property which depends on the amount of the substance present.

Example: Mass and volume are extensive properties. The more substance you have (the more its EXTENT), the greater the mass and volume it occupies.

Extensive properties are NOT used to identify substances. For example, a 5 g chunk of material could be dirt or gold.

An **INTENSIVE** property of a substance is a physical property which depends solely on the nature of the substance, and NOT on how much of the substance is present.

Example: Density and melting temperature are intensive properties. Pure gold always has the same density and melting temperature, regardless of amount.

Intensive properties are used to identity a substance.

EXERCISES:

13. Which of the following statements describe physical properties and which describe chemical properties?
    (a) Glass is transparent.
    (b) Salt melts at 801°C.
    (c) Adding lye to fat makes soap.
    (d) Copper conducts electricity.
    (e) Fumes from ammonia and hydrochloric acid mix to produce a white smoke.
14. Give an example of something which is observable but which does not contain matter.

15. Which of the following are intensive properties and which are extensive?
    (a) shape (c) length (e) electrical conductivity (g) hardness
    (b) smell (d) colour (f) time required to dissolve a solid

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Matter exists in THREE COMMON states or "phases": **solid, liquid** and **gas.** The three common phases of matter each have a unique set of properties which allow a given substance to be classified as one of a solid, liquid or gas.

(a) **Solids** are rigid, do not readily change their shape, and experience very small changes in volume when heated or subjected to pressure.

(b) **Liquids** conform to the shapes of their containers and experience only slight changes in volume when heated or subjected to pressure.

(c) **Gases** conform to the shapes of their containers and experience drastic changes in volume when heated or subjected to pressure.

Liquids can be seen as an "intermediate" phase between a solid and a gas when you consider that

- both ***solids*** and ***liquids*** undergo only slight changes in volume when heated or subjected to pressure
- both ***liquids*** and ***gases*** conform to the shapes of their containers.

Therefore, liquids share some properties with both solids and gases.

The space which exists between molecules is quite different in solids, liquids and gases.

In a solid, all of the particles are packed into a given volume in a highly organized and rigid manner which requires particles to be in direct contact with each other. So, solids are not compressible.

Example: The volume occupied by 28.0 g of solid nitrogen is 27.3 mL.

In a liquid, the particles remain in close contact with each other but have sufficient room to slide past one another easily and prevent an organized packing. Because of the close contact between particles, liquids are not compressible.

Example: The volume occupied by 28.0 g of liquid nitrogen is 34.6 mL.

In a gas, the particles are widely separated and only contact each other during collisions. Because the wide separation of particles can be decreased, gases are compressible.

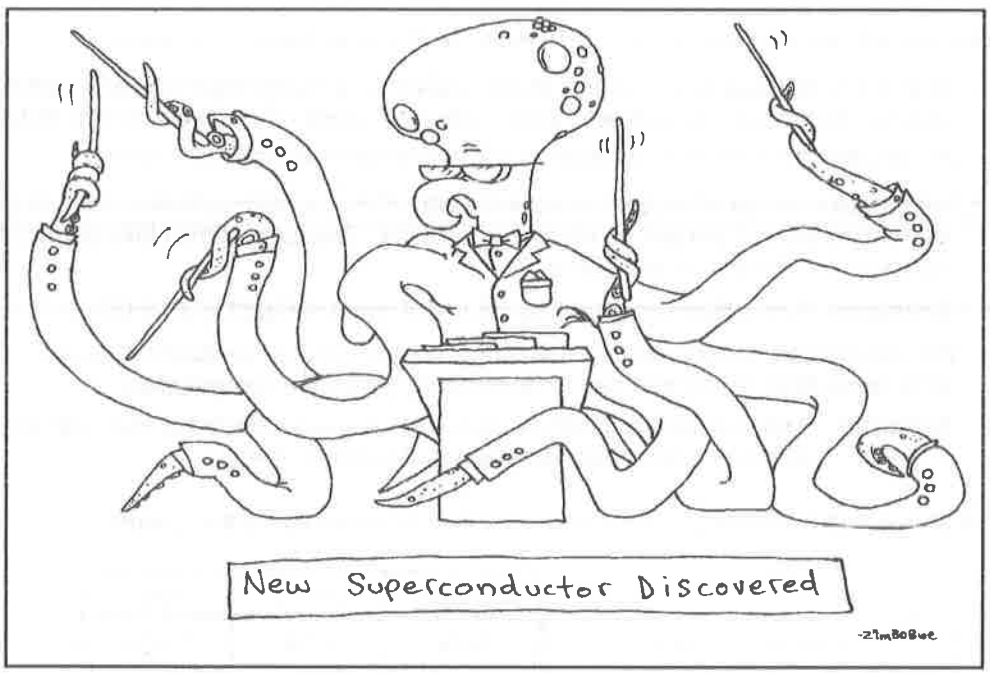
Example: The volume occupied by 28.0 g of gaseous nitrogen at room temperature (20°C) and one atmosphere pressure is 24 000 mL (24 L).

The table below shows that liquids are an intermediate phase between solids and gases.

	solid	liquid	gas
volume occupied	small	small	large
movement allowed in substance	small	large	very large

There are also several EXOTIC states of matter. You do not have to know them but you may find them of some interest.

- **PLASMA** – A plasma is a gas made of charged particles such as electrons and naked atomic nuclei at extremely high temperatures (10^7 °C). Examples are: lightning, the interior of stars, and neon lights.
- **SUPERCONDUCTIVE STATE** – Superconductive material allows electricity to flow without ANY resistance. An electrical pulse flowing in a superconducting loop can theoretically flow forever. Previously, the superconductive material had to be at very low temperatures (–250 °C), but exciting new discoveries have found superconductors which operate close to room temperature. These may allow new types of electronic devices and supercomputers to be created.
- **SUPERFLUID STATE** – A special type of helium (He–3) at –272°C has no frictional forces. If set in motion, the helium keeps moving at a constant rate. A container of He–3 also spontaneously self–siphons out of its container if it can eventually reach a lower level.
- **SUPERCONDENSED STATES** – These states are found in collapsed stars. In "white dwarf" stars the atoms are squeezed into a solid having a density of about 50 000 g/mL; in "neutron stars" the electrons, protons and neutrons are crushed together until the nuclear particles touch and effectively coalesce into a giant "neutron" with a density of 2×10^{14} g/mL; in a "black hole" matter is pushed together so tightly that it is crushed out of existence in our universe.



In order to continue examining physical properties, a few additional definitions are required.

Definitions: **HARDNESS** is the ability of a solid to resist abrasion or scratching.

MALLEABILITY is the ability to be rolled or hammered into thin sheets.

DUCTILITY is the ability to be stretched or drawn into wires.

LUSTRE is the manner in which a solid surface reflects light. Lustres can vary from metallic to adamantine (diamond–like), glassy, oily, pearly, silky or dull.

VISCOSITY is the ***resistance*** of a fluid to flow.

DIFFUSION is the intermingling of fluids as a result of motion within the fluid (this applies to both gases and liquids).

VAPOUR is the gaseous material formed by the evaporation of a substance which boils above room temperature.

Example: Acetone is a liquid which boils at 56°C. The acetone that evaporates at room temperature (25°C) is called a "vapour".

Oxygen boils far below room temperature and therefore is called a "gas" at room temperature.

VAPOUR PRESSURE is the pressure created by the vapour evaporating from a liquid. (Vapour pressure is abbreviated as VP.)

Note: The **BOILING TEMPERATURE** of a liquid (which you understand to a certain extent, and which will be properly defined later) is often called the "boiling point" and is abbreviated as **BP**.

Similarly, the **MELTING TEMPERATURE** of a solid is often called the "melting point" and is abbreviated as **MP**.

EXERCISE:

16. Colour is a physical property that MAY distinguish between two solids, two liquids, or two gases. Suggest as many other physical properties as you can which might also distinguish between two
(a) solids. (b) liquids. (c) gases.

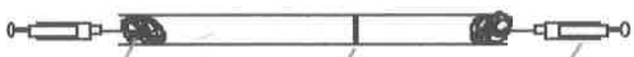
AN INVESTIGATION OF: VAPOUR PRESSURE, EVAPORATION RATE, DIFFUSION RATE, VISCOSITY, AND GAS COMPRESSIBILITY

For each experiment below, read the description of results and answer the questions which follow.

EXERCISES:

17. A sample of liquid butanol at room temperature is put into a pressure–measuring device and the vapour pressure created by the evaporation of vapour from the butanol is found to be 0.9 kPa. When a few millilitres of butanol are placed in an open evaporating dish on an electronic balance, the balance shows that the mass of butanol remains almost unchanged for several minutes. A thermometer placed in boiling butanol registers a temperature of 117°C. When the series of measurements is repeated with liquid acetone, the vapour pressure is found to be 31 kPa and the boiling temperature is found to be 56°C. When acetone is placed in an open evaporating dish on an electronic balance, the mass of acetone is found to decrease quickly and steadily.
    (a) Which substance, butanol or acetone, has the higher vapour pressure?
    (b) What relationship exists between a liquid's boiling temperature and its vapour pressure?
    (c) What relationship exists between a liquid's vapour pressure and its evaporation rate?
    (d) What would you expect to be true about the vapour pressure of iron metal?
    (e) Diethyl ether ("hospital ether") boils at 35°C. Compare the expected evaporation rate and vapour pressure of diethyl ether to acetone.

18. A few crystals of a water–soluble dye are dropped simultaneously into a basin containing hot water and a basin containing cold water. The dye spreads out quickly in the basin of hot water and much more slowly in the cold water. In a second experiment, a few dark purple crystals of iodine are sealed into an upright 30 cm glass tube at 50°C. After a few minutes a faint purple cloud of gaseous iodine is seen just above the iodine crystals, but the purple colour remains close to the bottom. Next, iodine crystals are sealed in a similar tube at 100°C. The purple gas formed spreads relatively quickly throughout the entire tube. In a third experiment, a 30 cm glass tube sealed with cotton batting at both ends has gaseous hydrogen chloride injected into one end while simultaneously gaseous ammonia is injected into the other end. After about 30 s, a white ring is seen inside the tube, as shown below. The white ring is solid ammonium chloride, produced when gaseous ammonia meets and reacts with gaseous hydrogen chloride.



(a) Propose a general relationship between temperature and diffusion rate.
(b) Which gas travels faster through air: ammonia or hydrogen chloride?
(c) The ammonia molecule is lighter than the hydrogen chloride molecule. What relationship appears to exist between a molecule's mass and the speed with which it travels? (In other words, what relationship appears to exist between a molecule's mass and its diffusion rate?)

19. Small steel pellets are simultaneously dropped down three glass tubes, each of which is one metre long and each of which contains a different liquid: hexane, carbon tetrachloride and glycerol. A pellet drops very quickly to the bottom of the tube containing hexane. A pellet drops quickly to the bottom of the tube containing carbon tetrachloride, but less quickly than in the tube of hexane. A pellet passing through the tube containing glycerol seems to "take forever" to reach the bottom. When 25 mL samples of each of carbon tetrachloride, glycerol and hexane are poured into the same tube, three distinct layers are seen. The carbon tetrachloride is found on the bottom, the glycerol lies in the middle and the hexane floats on top.
(a) Rank the liquids – hexane, glycerol and carbon tetrachloride – from lowest to highest viscosity.
(b) Rank the liquids – hexane, glycerol and carbon tetrachloride – from lowest to highest density.
(c) What relationship appears to exist between the viscosity and density of a liquid?

20. (a) If you use a tire pump to pump up a bicycle tire, what happens to the pressure exerted on the gas in the pump when you push down on the pump handle?
(b) What happens to the gas volume in the pump when you push down on the pump handle?
(c) Complete the following statement.

When the pressure on a gas increases, the volume of the gas ____________________.

(d) If you scoop up some freshly–fallen snow and compress it between your hands, what happens to the volume of the snow when you
(i) apply pressure, and (ii) release the pressure?
(e) If you squeeze a sponge in your hands, what happens to the volume of the sponge when you
(i) apply pressure, and (ii) release the pressure?
(f) What can you conclude about the ability of a gas sample to "recover" when an externally applied pressure is released?

ADDITIONAL EXERCISES:

21. Does corn syrup have a high or low viscosity? Does gasoline have a high or low viscosity? If you heat up a glass of syrup, what happens to the viscosity of the syrup?

22. Which of the following are extensive properties and which are intensive properties?
(a) viscosity
(b) the time required to melt a sample of ice
(c) the pressure exerted by helium in a balloon

23. Ice evaporates at a very slow rate. What can you conclude about the vapour pressure of ice?

24. Which of solids, liquids and gases can possess each of the properties below? There may be more than one phase possible in each case.
(a) the ability to flow rapidly
(b) transparency
(c) the ability to be easily compressed
(d) the ability to melt
(e) the ability to create a vapour pressure

25. Which phase of matter can be described as occupying a relatively small volume and allowing a large movement of the particles within the phase?

26. The volumes occupied by 32.0 g of oxygen are (in no particular order): 27.9 mL, 22.4 L and 22.4 mL for its different phases. What is the volume occupied by (a) solid oxygen? (b) gaseous oxygen?

27. A balloon is filled with air and sealed. **At this point the balloon is neither expanding nor shrinking.**
    (a) Is the pressure inside the balloon greater than, equal to or less than the atmospheric pressure pushing on the outside of the balloon? [Hint: What happens if the atmospheric pressure rises?]
    (b) What would you expect to observe if you put a filled balloon in a sealed glass tank and then created a vacuum in the tank?
28. Water vapour is more viscous than chlorine gas. What can you predict regarding the density of water vapour versus chlorine gas?
29. (a) Harry is sitting three metres away from Harriet in a cold room. John is sitting three metres away from Juanita in a warm room. Will Juanita smell John's aftershave lotion before or after Harriet smells Harry's identical aftershave lotion? Why?
    (b) Assume Harry and John wore different aftershave lotions, but were in rooms having the same temperature. What physical property would determine whether the scent of John's aftershave would reach Juanita before the scent of Harry's aftershave reached Harriet?
30. Viscosity is measured in "poise". The viscosities of argon gas and chlorine gas are 222 micropoise and 133 micropoise respectively at 25°C. Will a small plastic sphere fall faster in argon or chlorine gas?
31. If the boiling temperature of chloroform is 62°C, should gaseous chloroform at room temperature be called a "gas" or a "vapour".
32. Glycerol has a lower vapour pressure than does ethanol at room temperature. Which of glycerol and ethanol has the lower boiling point?

III.3. THE CLASSIFICATION OF MATTER

The physical properties possessed by substances are used to classify the substances into various categories.

Definitions: A **SYSTEM** is the part of the universe being studied in a given situation.

A **PHASE** is any part of a system which is uniform in both its composition and properties.

Phases in a system are distinct REGIONS separated from each other by VISIBLE BOUNDARIES (although sometimes a microscope is needed to see the boundaries).

EXAMPLE: When a copper rod is placed in water, all the points within the copper have the same properties, and all the points within the water have the same properties. In this TWO PHASE system there is a distinct boundary where the copper stops and the water exists.

Before proceeding, a few more definitions are needed. You should recall these definitions from previous science courses, but they are included here for completeness.

Definitions: An **ELEMENT** is a substance which cannot be separated into simpler substances as a result of any chemical process.

Examples: silver metal, copper metal, hydrogen gas

An **ATOM** is the smallest possible unit of an element which retains the fundamental properties of the element.

Examples: silver (Ag), copper (Cu), hydrogen (H)

A **MOLECULE** is a cluster of two or more atoms held together strongly by electrical forces.

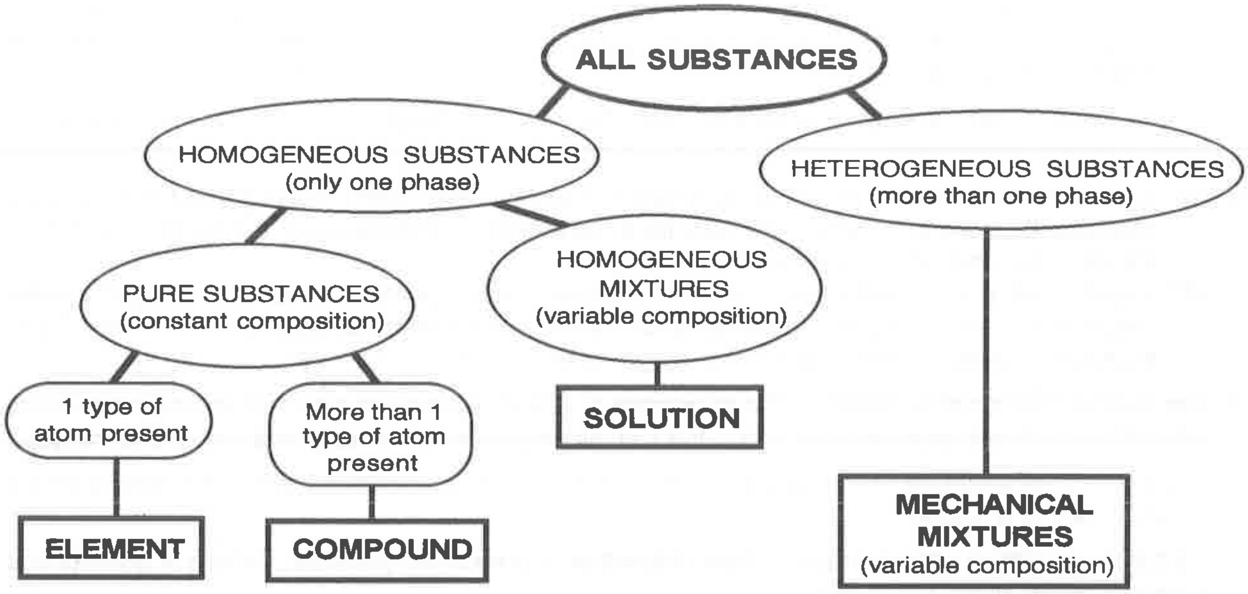
Examples: water (H_2O), ethanol (CH_3CH_2OH), table sugar ($C_{12}H_{22}O_{11}$)

An **ION** is an atom or molecule which possesses an electrical charge.

Examples: sodium ion (Na^+), chloride ion (Cl^-), nitrate ion (NO_3^-)

A **PARTICLE** is a general term used to describe a small bit of matter such as an atom, molecule or ion.

The following scheme is generally accepted for classifying substances.



The following terms are used to classify matter. (Refer to the above classification diagram as the terms below are discussed.)

1. A **HOMOGENEOUS** substance is a substance consisting of only one phase.

   Examples: air, water, salt water, a piece of iron

2. A **HETEROGENEOUS** substance is a substance consisting of more than one phase.

   Examples: a human being, a pencil, gravel

3. A **PURE SUBSTANCE** is a substance that is homogeneous and has an unchangeable composition.

   Examples: sugar, water, copper, iron

4. A **MIXTURE** is a system made up of two or more substances, such that the relative amounts of each substance can be VARIED.

   Examples: salt dissolved in water, alcohol dissolved in water

   "Mixture" is a general term which includes both heterogeneous mixtures (better known as "mechanical mixtures") and homogeneous mixtures (better known as "solutions").

5. A **MECHANICAL MIXTURE** is a heterogeneous mixture of two or more substances.

   Examples: gravel, sand and iron filings, a pencil

   Note: ALL HETEROGENEOUS substances are MECHANICAL MIXTURES, and vice versa.

6. A **SOLUTION** is a homogeneous mixture of two or more substances.

   There are several different types of solutions.

Type of Solution	Example
gas–in–gas solution	air (oxygen, nitrogen, etc.)
gas–in–liquid solution	soda pop
liquid–in–liquid solution	water and alcohol
solid–in–liquid solution	salt water
solid–in–solid solution	alloys (metals melted together)