Shika na Mikono
Version 4.0 TZ

Hands-On Science Resource Manual

Tanzania

June 9, 2014

Questions or Comments?

Thank you for using the *Shika na Mikono* manual! If you have any questions, comments, or would like to request a copy of this manual, please use the contact information given below.

Shika na Mikono Primary Contact (shika.mikono.tz@gmail.com)

You may also contact any of the current members of the Shika na Mikono team directly using the following contact information.

Joe Antonacci (mdfb30@gmail.com)
Cait Baumhart (cbaumhart73@gmail.com)
Riley Burgon (riley.burgon@gmail.com)
Megan Gooley (ooley.gooley@gmail.com)
Rickie Likens (rlikens@gmail.com)
Joel Nightingale (joelnightingale6@gmail.com)
Nikki Yates (nikki.yates.0909@gmail.com)

Belle Archaphorn (belledundi@gmail.com)
Willie Blackmon (willieblackmon@gmail.com)
Steve Bonomo (sbonomo3@gmail.com)
Ryan Early (ryan.early@gmail.com)
Ben Savonen (blsavone@mtu.edu)

To download a digital version of this manual, please visit the digital download page at the following address: http://pctanzania.org/shika

Feedback and requests for revisions to the content of any of the manuals can be made using the primary contact information above, or by visiting the Shika Project web page at the following address: https://code.google.com/p/shika-na-mikono/issues/list

About This Book

Shika na Mikono is a training manual for U.S. Peace Corps Volunteers serving in Tanzania. Many of these Volunteers work in schools without laboratories and all are new to the details of the Tanzanian syllabus. This manual is prepared to help them teach science more effectively.

Just before the rains returned at the end of 2009, Peace Corps Tanzania recruited Leigh Carroll and I to conduct a training on laboratory methods for the new cohort of education Volunteers. We decided to write a hand-out for our presentation, summarizing "everything that Volunteers needed to know about the lab." We quickly realized that this was a much larger undertaking than we could manage on our own and began recruiting other Peace Corps Volunteers. Thus the *Shika na Mikono* project was born.

As the first edition came together, several themes emerged. We believe that hands-on activities are essential for learning science. We believe that the use of local and low cost materials can enable any school to do these activities. Finally, we believe that for hands-on science to be successful, it must be safe. These three ideas – interactive learning, equity, and safety – remain the core of the *Shika* vision.

Given the enthusiastic response to the first edition and the significant work the project had done in the past year, we decided to produce a second edition, this one for the 2010 Peace Corps Pre-Service Training. Once again, many people contributed to this effort, most especially PCVs Michael Rush, Kristen Grauer-Gray, and Peter McDonough – without their excellent ideas and passionate hard work, this book would not exist.

The third edition of *Shika na Mikono* for Tanzania included additional content created by PCVs Jessi Bond, Carolyn Rhodebeck, and Dylan Masters. Peter McDonough also made substantial additions and revision. PCV Dave Berg provided the inspiration, instruction, and much of the labor to import the manual into LaTeX.

Many of the ideas for locally available materials come from or were inspired by the Source Books published by the Mzumbe Book Project, Morogoro, Tanzania. Several other ideas for locally available materials were developed at Bihawana Secondary by Mwl. Mohamed Mwijuma. PCVs Peter Finin and Gregor Passolt wrote a book on physics demonstrations in 2008 that has been incorporated wholesale into the activities portion of this manual. My own knowledge of the laboratory was greatly increased by a brilliant if ancient book found on the shelves of Bihawana Secondary – the cover and title pages with the title have long since been lost, but the preface sites G.P.Rendle, M.D.W. Vokins and P.M.H. Davis as authors, and 1967 as the date of publication.

We are all grateful to our schools for giving us the opportunity to work in such supportive environments, the freedom to explore these ideas, and the time to document them. We have certainly benefited from the wisdom and creativity of many other teachers, both in this country and in America. Many of us working without reliable electricity or internet connections benefited enormously from the hospitality of the numerous families who sheltered us in town. We are all grateful to Peace Corps Tanzania for supporting our work, especially James Ogondiek and the now retired and much beloved Thomas Msuka, both of whom recognized early on the value of this project and advocated for us to undertake the work required to develop and spread the ideas in this book.

Most of all, we are grateful to our students, for it is their curiosity and enthusiasm that has motivated everything.

Aron Walker
Klerruu Teachers' College, Iringa
aronwalk@alum.mit.edu

Preface to the Fourth Edition

Since its original release in 2009, this manual has served as a teaching resource and guide to incorporating interactive learning methods into Tanzanian laboratories and classrooms using locally available materials.

Now in its fourth edition, the manual has broadened its scope in terms of its content and application to the Tanzanian secondary school curricula. The resources within the manual have been divided to better isolate and address the different challenges faced in laboratory and classroom settings. Despite the ability of both of these settings to foster an interactive science education, the two are distinguished only in their adherence to the educational requirements enforced by the Ministry of Education, such as NECTA practicals.

Information regarding laboratory practicals (including a newly added compilation of NECTA past papers), as well as general guidance for starting, developing and maintaining school laboratories, is the focus of this manual. Version 4 now also contains a guide to incorporating and hosting various math- and science-related events for students and teachers, including science competitions, conferences and trainings.

Hands-on activities for the classroom, as well as extensions to outside learning, science clubs, projects and community involvement are now located in the separate subject-specific *Shika Express* companion manuals for Biology, Chemistry, Physics and Mathematics.

Much of the newly added content to the *Shika Express* manuals was inspired by or directly came from the Source Books of the Mzumbe Book Project in Morogoro, Tanzania, as well as *The Science Teachers' Handbook* and *The Maths Teachers' Handbook*, published by the VSO ECOE Programme. Additional activities came from *The Everyday Science Sourcebook* by Lawrence F. Lowery and *150+ Easy Science Experiments*, published by Mark Twain Media/Carson-Dellosa Publishing Company, Inc.

Continued development of the *Shika na Mikono* resources is made possible by a dedicated team of individuals made up of Peace Corps Volunteers and Tanzanian teachers and facilitators. This collaboration among volunteers and Tanzanian nationals is what makes possible the continued success and relevance of all of the *Shika na Mikono* teaching resources.

Steve Bonomo
Wilima Secondary School, Ruvuma
sbonomo3@gmail.com

Belle Archaphorn
Mwatisi Secondary School, Tukuyu
belledundi@gmail.com

Ben Savonen
Ismani Secondary School, Iringa
blsavone@mtu.edu

Ryan Early
Chamwino Secondary School, Dodoma
ryan.early@gmail.com

Willie Blackmon
Rungwe Secondary School, Mbeya
willieblackmon@gmail.com

The Shika na Mikono Teaching Philosophy

The Value of an Interactive Math and Science Education

Math and science education provides a major challenge throughout the world. In fact, formal education as a whole is subject to a large number of constraints which limit a student's educational environment and may hinder her overall capacity for understanding and thinking critically.

As human beings, we do not live merely to assimilate to our surroundings and mimic what is being done by others. Rather, it is our unique advantage to continually challenge ourselves to extend our current limits of understanding. Utilization of this remarkable opportunity as a global community requires individuals with a scientific mindset who are able to ask critical questions and invent solutions to the increasingly difficult challenges facing a sustainable coexistence with nature.

It is the aim of education to immerse developing minds in an environment that fosters these kinds of analytical and creative abilities. Unfortunately, such goals often fail to be realized due to a number of limitations – time, money, teacher-to-student ratio and educational resources, to name a few. Although these challenges can put many communities at a great disadvantage and create divides among learners, they should not prevent educators from inspiring imagination and innovativeness among their students. The materials we need are already available in our villages and in our towns. The key ingredients in science education are not precision glassware, imported reagents, or massive loans. The key ingredients are curiosity, creativity, and the ability of teachers to use what they already have to provide students with experiences that broaden their understanding of the world. When students explore the world with readily available materials instead of expensive lab apparatus and high-end chemicals – when they see parts of their own world appear in the classroom – they gain an understanding that bridges scientific theory and daily life. That understanding sheds light on the world beyond the laboratory, which lets them wield scientific thinking anywhere.

While this goal can be achieved by anyone regardless of location, it does require a suitable learning environment. Training students to follow commands and adhere to a linear method of learning can only produce obedience, not intelligence. In order to continue to adapt and learn in a quickly changing world, students must be able to interact with their surroundings and engage themselves in the content of their studies.

No fields offer such readily available opportunities to practice these skills as do math and science. The ability to observe and appreciate the applications of math and science in our daily lives is an essential component of a person's intellectual development. It is through interactive education that this can be acquired.

A hands-on, interactive approach to learning truly has many benefits inherent in and of itself. It fosters a deeper understanding through experiential discovery as opposed to rote memorization. It encourages learning through enjoyment and excitement, which are much stronger motivators than the fear of punishment. It allows for the development of creativity and curiosity within students, invaluable qualities with benefits far beyond the realm of formal education. It empowers students to seek their own answers and truth in all matters – not just for test questions in school, but for the ones that really matter in life.

Rather than merely encouraging an interactive approach to learning, the aim of *Shika na Mikono* is also to illuminate the relationship between the classroom and all the aspects of a student's life and natural environment. Thus, science becomes the study of reality. An understanding and appreciation of the world around us is what allows us as human beings to interact with and respect all other things in our universe. It is in this way that students can become equipped to take on new challenges and ensure the prosperity of generations to come.

The Teacher's Role in Student Development

To learn science, students must interact with the world around them. They must ask their own questions and seek their own answers. They must see things and they must grasp them in their hands; hence the name of this book: *shika na mikono* is Swahili for "grasping in hands."

It is our fervent belief that every student in the world should perform science practical exercises. For too long we have heard complaints that schools lack the materials necessary for these exercises. This book attempts to make clear that students may perform science practicals at any school, most especially at those without traditional laboratories, starting today. Everything teachers need to create these hands-on learning experiences is available locally and/or at low cost.

Many national syllabi require practical exercises, often on their national examinations. This is good. Critically, however, we urge teachers to expand the scope of students' hands-on work beyond the practicals for the national exams. Every topic, every lesson may be a "practical" – not just a demonstration on the front bench but an opportunity for students to touch and manipulate and discover on their own.

In this vision, the science teacher becomes a guide, someone who can assemble parts of the natural world into a compelling lesson and ask the questions that help students see how things work. In this capacity, the science teacher remains a resource irreplaceable by the march of technology. Photocopy machines produce student editions of notes much more efficiently than teachers copying them to the board for students to copy again. Instructional films shown on low-cost solar-powered projectors offer students articulate explanations and demonstrations. But no technology can replace the essential role of the modern science teacher: s/he is an architect, one who builds a space in which students can learn for themselves, and a shepherd, who tends to their learning through that discovery.

The aim of this book is to inspire and empower this sort of teaching. For years, many educators have bantered about the phrase "student-centered teaching." This sounds rather like patient-centered medicine – anything else is simply absurd. The focus of a lesson must always be the experience of the student. To prepare such a lesson, the teacher should answer the following questions:

- What will the student do in class?
- How will she use her hands to interact with the world?
- What will the student observe with her own senses?
- Given these experiences, what questions will arise from the student's observations?
- Given these questions, how might the teacher respond to provoke further inquiry and critical thinking?
- How might the student's peers respond to build a common understanding?
- How might the student, through further observation and experimentation, arrive at the answer herself?
- Given these goals, what experiences will put her on the journey to that answer?
- What series of activities should be offered to her to facilitate that discovery?

This is student-centered teaching – a lesson plan crafted around the experience of the student – the internal, cognitive, and emotional experience of being in class that day.

The process of answering these questions involves several steps. The teacher must first organize the material that the student is to understand into a well-structured framework: logical, sequential, and hierarchical. Then, using this framework, the teacher should design activities for the student to discover each aspect of the material. These activities should be sequenced to expand understanding, moving from simple phenomena to the more complex; from the specific to the general. Discussion questions should seek first to uncover core phenomena and then to link each new insight with what the students already understand about their world. Targeted questions catalyze introspection, group discussions, and the realizations necessary for the students themselves to start articulating scientific theories. Once the students have discovered phenomena, linked them to pre-existing understanding, and begun articulating general theory, the teacher can help focus and form these articulations into the accepted vocabulary and nomenclature of modern science.

Finally, we teachers must embrace questions; we must encourage students to ask about what they do not understand. Rather than answer these questions directly, whenever possible we should design experiments or ask questions in return that allow students to find answers for themselves. As role models, we also must embrace the limits of our own understanding. Often students ask questions to which we do not know the answer. This is a fundamental aspect of science education. Our job is to help students to understand the world better, to guide them in that discovery. Our job is not to know everything; this is neither necessary, nor is it possible, nor even desirable. When our students observe us confronting the

unknown, when they see how we ask questions and perform experiments ourselves to seek out the truth, then they become more comfortable asking questions and seeking answers themselves. This experience helps them to understand the true power of science, that a person anywhere may always find the answer.

Let us gather the world around us and put it in the hands of our students, so they might understand how it works. Let us let them grasp it in their hands – *walishike na mikono yao.*

Contents

About This Book 3

Preface to the Fourth Edition 4

The Shika na Mikono Teaching Philosophy 5

I Laboratory Development 11

1 Starting School Laboratories 12

2 Specific Technical Needs of a School Laboratory 13

3 Improving an Existing School Laboratory 16

4 Salvaging Old Equipment 18

5 Identifying Unknown Chemicals 21

II Laboratory Safety 29

6 Guidelines for Laboratory Safety 30

7 First Aid 34

8 Dangerous Chemicals 37

9 Dangerous Techniques 43

III Laboratory Management 45

10 Classroom Management in the Laboratory 46

11 Routine Cleanup and Upkeep 48

12 Waste Disposal 50

13 Recycling Materials 52

IV Laboratory Techniques 54

Laboratory Techniques in Chemistry 55

14 Use of the Beam Balance for Measuring Chemicals 56

15 Use of a Plastic Syringe to Measure Volume 58

16 Measures of Concentration 60

17 Calculating the Molarity of Bottled Liquids 62

18 Preparation of Solutions 63

19 Preparation of Solutions Without a Balance 65

20 Relative Standardization 66

Laboratory Techniques in Biology 68

21 Preservation of Specimens 69

22 Dissection 70

23 Preparation of Culture Media 72

24 Using a Microscope 74

25 Low Tech Microscopy 76

V NECTA Practicals 78

26 Biology Practicals 79
26.1 Introduction to Biology Practicals . . . 79
26.2 Food Tests . . . 80
26.3 Classification . . . 85
26.4 Respiration . . . 88
26.5 Transport . . . 89
26.6 Photosynthesis . . . 91

27 Chemistry Practicals 94
27.1 Introduction to Chemistry Practicals . . . 94
27.2 Volumetric Analysis . . . 95
27.3 Qualitative Analysis . . . 106
27.4 Chemical Kinetics and Equilibrium . . . 119

28 Physics Practicals 123
28.1 Introduction to Physics Practicals . . . 123
28.2 Mathematics . . . 125
28.3 Mechanics . . . 127
28.3.1 Hooke's Law . . . 127
28.3.2 Simple Pendulum . . . 129
28.3.3 Principle of Moments . . . 130
28.4 Light . . . 132
28.4.1 Plane Mirror (Reflection) . . . 132
28.4.2 Rectangular Prism (Refraction) . . . 133
28.5 Electricity . . . 135
28.5.1 Potentiometers . . . 136
28.5.2 Metre Bridges . . . 137
28.5.3 Ohm's Law . . . 139

VI Science Activities and Competitions 141

29 Hosting Science Events 142
29.1 Box of Fun . . . 142
29.2 Shika Express Gallery Walk . . . 142
29.3 Science Fair . . . 143
29.4 Science Competitions . . . 143
29.5 Science Day . . . 143
29.6 Math and Science Conferences . . . 144
29.7 Teacher Trainings . . . 144

30 Science Competitions **145**
Egg Drop . . . 146
Jenga Jengo . . . 147
Raft Rally . . . 148
Drop Zone . . . 149
Bridge Challenge . . . 150

31 The Scientific Procedure **152**
Biology . . . 152
Chemistry . . . 154
Physics . . . 156

Appendix **160**

NECTA Practical Past Papers **160**

Biology Past Papers **161**
2013 - BIOLOGY 2A (ACTUAL PRACTICAL A) . . . 162
2012 - BIOLOGY 2A (ACTUAL PRACTICAL A) . . . 163
2011 - BIOLOGY 2A (ACTUAL PRACTICAL A) . . . 164
2010 - BIOLOGY 2A (ALTERNATIVE A PRACTICAL) . . . 165
2009 - BIOLOGY 2A (ALTERNATIVE A PRACTICAL) . . . 166
2008 - BIOLOGY 2A (ALTERNATIVE A PRACTICAL) . . . 167
2007 - BIOLOGY 2A (ALTERNATIVE A PRACTICAL) . . . 168

Chemistry Past Papers **169**
2013 - CHEMISTRY 2A ACTUAL PRACTICAL A . . . 170
2012 - CHEMISTRY 2A ACTUAL PRACTICAL A . . . 172
2011 - CHEMISTRY 2A ACTUAL PRACTICAL A . . . 174
2010 - CHEMISTRY 2A ALTERNATIVE A PRACTICAL . . . 176
2009 - CHEMISTRY 2A ALTERNATIVE A PRACTICAL . . . 178
2008 - CHEMISTRY 2A ALTERNATIVE A PRACTICAL . . . 180
2007 - CHEMISTRY 2A ALTERNATIVE A PRACTICAL . . . 182

Physics Past Papers **184**
2013 - PHYSICS 2A ACTUAL PRACTICAL A . . . 185
2012 - PHYSICS 2A ACTUAL PRACTICAL A . . . 187
2011 - PHYSICS 2A ACTUAL PRACTICAL A . . . 189
2010 - PHYSICS 2A ALTERNATIVE A PRACTICAL . . . 190
2009 - PHYSICS 2A ALTERNATIVE A PRACTICAL . . . 193
2008 - PHYSICS 2A ALTERNATIVE A PRACTICAL . . . 195
2007 - PHYSICS 2A ALTERNATIVE A PRACTICAL . . . 198
2006 - PHYSICS 2A ALTERNATIVE A PRACTICAL . . . 201
2005 - PHYSICS 2A ALTERNATIVE A PRACTICAL . . . 203
2004 - PHYSICS 2A ALTERNATIVE A PRACTICAL . . . 205

A Sources of Laboratory Equipment **208**

B Science Shillings **221**

C Kiswahili Laboratory Glossary **222**

Index **225**

Part I

Laboratory Development

Chapter 1

Starting School Laboratories

A science laboratory is any place where students learn science with their hands. It might be a room, or just a box. The goal is to develop a space that facilitates hands-on learning.

1.1 Benefits of a School Laboratory

There are many benefits of having a laboratory:

- Students learn more and better science
- Students get more excited about science class
- Students have to go to the lab for class, thus eliminating those too lazy to walk over
- Practical exams are easier than the alternative-to-practical exams
- Everyone thinks practicals are important, and that science without practicals is silly.

1.2 Challenges of a School Laboratory

There are some challenges with having a laboratory:

- They are places where people can get hurt
  This is true. Please see the sections on Classroom Management in the Laboratory (p. 46) and Laboratory Safety (p. 30) to mitigate this risk.
- Many teachers do not know how to use a laboratory
  Then use the lab to teach them how to use it, thus spreading skills.
- Laboratories are far too expensive for poor schools to build and stock
  This is simply incorrect. Any room will work for a lab, and any school can afford the materials required to stock it. The rest of this book is dedicated to this point.

So you want to build a laboratory?

1.3 Step one: Location

A permanent location is obviously preferable. If your school has an extra classroom, great. The only requirements of a potential room are that it be well ventilated (have windows that either open or lack glass altogether) and be secure: bars in the windows, a sturdy door, and a lock. If you plan to put fancy equipment in your lab, remember that hack saw blades are cheap and that the latch through which many pad locks pass can be cut quickly regardless of the lock it holds. But if you are just starting, there will probably not be any fancy equipment; a simple lock is enough to keep overly excited students from conducting unsupervised experiments.

If there is no extra space at all, the lab can live in a few buckets and be deployed in a classroom during class time. "There is no lab room," is no excuse for not having a lab.

1.4 Step two: Funding

Yes, some is required. But the amount is surprisingly little – in most countries a single month of a teacher's salary is enough to furnish a basic laboratory. Almost every school can find the amount required to get started, and if not the community certainly can. A single cow in most countries would pay for a basic laboratory many times over. A cow is valuable. So is science education.

We encourage you to resist the temptation to ask people outside of the school community or school system to pay for the lab. There is simply no need to encourage that sort of dependence; this can be done locally, and it should be.

Chapter 2

Specific Technical Needs of a School Laboratory

Laboratories facilitate hands-on investigation of various phenomena. Every syllabus requires different topics for study, but a core of topics provide a good foundation for each subject.

2.1 Basic Biology Laboratory

A basic biology laboratory should allow the following investigations:

- Collection, shelter, and observation of living specimens (plant, insect, fish, reptile, mammal)
- Bacterial and fungal cultures
- Preservation and dissection of dead specimens (plant, insect, fish, reptile, mammal; both whole and parts thereof)
- Assembly and observation of miniature ecosystems
- Low power microscopy
- Diffusion and osmosis
- Chemical tests of basic biological molecules ("biochemical tests" / "food tests")
- Chemical analysis of the products of animal and plant respiration
- Non-invasive investigation of human systems (nervous, sensory, circulatory, muscular, parts of the digestive)

Key materials are:

- Containers, bottles, tubes, super glue
- Plants, insects, fish, (safe) reptiles, and small mammals
- Sugar, starch, protein source, fertilizer, salt, food coloring
- Chemicals for preservation of specimens
- Scalpels and pins
- Low power microscopes (water drop microscopes, locally assembled)
- Reagents for biochemical tests
- Reagents for gas identification
- Stopwatches
- Heat sources

2.2 Basic Chemistry Laboratory

A basic chemistry laboratory should allow for the following investigations:

- Distinguishing compounds from mixtures, preparing chemical compounds, separating mixtures
- Changes in the state of matter (melting/freezing, evaporation/condensation, sublimation/deposition, dissolution/crystallization)
- Comparison of metals and non-metals
- Comparison of covalent and ionic (electrovalent) compounds
- Observing various elements and compounds and their reactivity with air, water, acids and bases
- Acid/base, oxidation/reduction, and precipitation reactions
- Energy changes from chemical reactions (thermochemistry, energetics)
- Factors affecting the rates of chemical reactions (chemical kinetics)
- Properties of gases (gas laws)
- Preparation of gases (hydrogen, oxygen, carbon dioxide)
- Electrochemical experiments (conductivity, electrolysis, electroplating, voltage generation)
- Volumetric analysis (titration)
- Identification of unknown salts ("qualitative analysis")
- Very basic organic reactions (e.g. preparation of ethanol by fermentation, oxidation of ethanol to ethanal)

Key materials are:

- Containers, bottles, tubes, balloons
- Tools for measuring volume (calibrated plastic water bottles, plastic syringes)
- Low cost balance (digital)
- Heat sources and open non-luminous flames
- Stopwatches
- Power supplies (e.g. batteries) and wires
- Wide variety of chemicals including metallic elements, non-metallic elements, solid covalent compounds, salts, acids, bases, redox reagents, indicators, and many chemicals for specific kinds of reactions

2.3 Basic Physics Laboratory

A basic physics laboratory should allow for the following investigations:

- Measuring volume, mass, and density of liquids and solid objects
- Measuring time, velocity, acceleration
- Gravitational acceleration, force, and friction
- Mechanical tools (levers, pulleys, etc.)
- Simple harmonic motion (pendulum, spring)
- Temperature, heat capacity, and heat transfer (conduction, convection, radiation)

- Waves (including water and sound)
- Optical experiments (reflection, refraction, diffraction)
- Electromagnetic experiments (conductivity, magnetic field lines, induction, motors, electrical generation)
- Simple circuits (including resistors, capacitors, and switches)

Key materials are:

- A low cost balance (digital)
- Tools for measuring volume
- Containers, misc. objects, bottles, etc.
- Stopwatches
- Heat sources
- Thermometers
- String, springs, wire
- Water, oil, sand, rocks
- Mirrors, lenses, glass blocks, diffracting surfaces
- Magnets
- Power supply (e.g. batteries)
- Inexpensive multimeters or locally made galvanometers
- Electrical components

Chapter 3

Improving an Existing School Laboratory

If there is already a laboratory at your school, the immediate tasks are to see what it has, make it safe, get it organized, make repairs, and ensure smart use with sound management.

3.1 Inventory

- Making a list of what and how much of everything is in your lab is easy, if time consuming. Difficulties arise when you find apparatus you have never seen before, or containers of chemicals without labels.
- There is no harm in unknown apparatus, they just are not useful until you know what they do. Ask around.
- Unknown chemicals, however, pose a hazard, because it is unclear how to properly store them or how to clean up spills. If a chemical is unknown, there is no safe way to responsibly dispose of it. Therefore, it is best to attempt to identify unknown chemicals. For assistance in identifying unknown chemicals, please see Identifying Unknown Chemicals (p. 21).
- Burettes and apparatus concerning electricity, for example voltmeters and ammeters, should be tested to ensure that they work. Please consult Traditional Volumetric Analysis Technique (p. 99) to learn how to use burettes and Checking Voltmeters and Ammeters/Galvanometers (p. 18) to do just that.

3.2 Organize

3.2.1 Have enough space

The key to organization is having enough space. Usually, this means building shelves. In the long term, find a carpenter to build good shelves. In the short term, boards and bricks, scrap materials, chairs, anything to provide sturdy and horizontal storage space. It should be possible to read the label of every chemical, and to see each piece of equipment

3.2.2 Apparatus

- Arrange apparatus neatly so it is easy to find each piece.
- Put similar things together.
- Beakers can be nested like Russian dolls.

3.2.3 Chemicals

- Organize chemicals alphabetically. There are more complicated schemes involving the function or the properties of the chemical but what is most important is a scheme that everyone working in the lab can follow. ABC is the easiest, and has the best chance of being used. A good alternative is to organize by chemical makeup (e.g. sodium, etc.)
- Glass bottles of liquid chemicals should be kept on the floor, unless the laboratory is prone to flooding, in which case they should be on a sufficiently elevated, broad and stable surface. What you do not want are these bottles falling and breaking open.
- Million's Reagent, benzene, and other chemicals that should never be used should be kept in a special place, ideally locked away, and labelled to prevent use. See Dangerous Chemicals (p. 37) for a list of chemicals that should never be used.

- Label plastic containers directly with a permanent pen, especially if the printed label is starting to come off.
- Replace broken or cracked containers with new ones.

3.2.4 Make a map and ledger

- Once you have labeled and organized everything in a lab, draw a map.
- Sketch the layout of your laboratory and label the benches and shelves.
- In a ledger or notebook, write down what you have and the quantity. For example, Bench 6 contains 20 test tubes, 3 test tube holders, and 4 aluminum pots.

This way, when you need something specific, you can find it easily. Further, this helps other teachers – especially new ones – better use the lab. Finally, having a continuously updated inventory will let you know what materials need to be replaced or are in short supply. Proper inventories are a critical part of maintaining a laboratory, and they really simplify things around exam time.

3.3 Repair/Improve

Once the lab is organized, it is easy to find small improvements. Here are some ideas:

3.3.1 Build more shelves

You really cannot have too many.

3.3.2 Fix broken burettes

Burettes are useful, expensive and – if glass – fragile. Broken burettes can often be made functional again. If you have broken burettes, see Repairing Burettes (p. 19).

3.3.3 Check voltage and current meters

Voltmeters, ammeters and galvanometers often get discarded or unused despite still being able to function sufficiently for use during demonstrations and practicals. Before getting rid of these meters, see the section on Checking Voltmeters and Ammeters/Galvanometers (p. 18).

3.3.4 Identify key apparatus needs

Sometimes a few pieces of apparatus can be very enabling, like enough measuring cylinders, for example. Buy plastic!

3.4 What next?

Once the lab is safe and organized, develop a system for keeping it that way. Consider the advice in Routine Cleanup and Upkeep (p. 48). Make sure students and other teachers in involved.

Then, start using the lab! Every class can be a lab class. That is the whole point.

Chapter 4

Salvaging Old Equipment

4.1 Checking Voltmeters and Ammeters/Galvanometers

Needed: Meters to check, a couple wires, some resistors and a fresh battery.

Important note: There is a wrong way to hook up the meter. The needle will try to deflect down because negative and positive are swapped. If the reading is zero, make sure that you try the opposite connection to be sure.

4.1.1 Voltmeters

Hook up the voltmeter across the battery. The battery is probably 1.5 V, but do not worry if you see 1.1, 1.2, even if using a brand new battery. Try not to use a battery that reads much below 1 V on several different meters.

Unuseable Voltmeters

- Totally dead, no deflection of the needle
- Voltage reading jumps excessively. Ensure that the connections are solid and test again.
- Measured voltage is totally wrong, not close to 1.5 V

Useable Voltmeters

Read a voltage close to 1.5. If the voltage if not 1.5 exactly, the voltmeter is probably working fine, and the battery is just off a bit.

4.1.2 Ammeters

Hook up the ammeter in series with a resistor. Because you do not necessarily know the condition of the ammeter before testing, be sure to have several different resistors on hand. An ammeter may appear not to work if resistance is too high or too low. Start testing different ammeters.

Unuseable Ammeters

- Totally dead, no deflection of the needle
- Current reading jumps excessively (but check connections)
- Totally wrong, reads much different from other ammeters

Useable Ammeters

Read a current similar to other ammeters. Hard to say exactly what current, but feel free to calculate based on your resistor using $V = IR$, although do not forget that there is some internal resistance r of battery, so $V = I(R + r)$. The resistance of the resistor is usually coded on the resistor in a series in stripes – see the instructions under Resistors in Sources of Laboratory Equipment (p. 216).

Tip: You can hold the wires onto the battery with your fingers; the current is far too low to shock you.

Other: Now that you have tested to see if your voltmeters and ammeters work, you can feel free to check all of them for accuracy, by calculating expected values and comparing between meters. Most practicals will still work alright with "somewhat" accurate meters, and most meters are either fine, or broken.

4.2 Repairing Burettes

First, if you need burettes, consider buying plastic burettes. They are widely available if you ask persistently and they tend not to break. This may be hard as many suppliers prefer to sell glass burettes. Why? As one supplier told us, "Because when people buy plastic burettes, they don't return."

The good news for every school with glass burettes is than often broken burettes can be repaired.

4.2.1 The top of the burette is broken, above the 0 mL line.

This burette is still fully functional. A student will probably need a beaker for filling the burette, but she should be using one anyway. Use a metal file (best!), stone, or piece of cement to gently grind the broken edge smooth to prevent cuts.

4.2.2 The burette is broken in the graduated section, that is, between 0 ml and 50 ml.

This burette is still slightly useful for titrations if it has most of its length. Students will just have an initial volume of 7 ml, perhaps. If it has broken around the 45 ml mark, no such luck. The burette tube however, is still quite useful as a glass pipe. Keep it around for other kinds of experiments. At the very least you have a glass rod for mixing solutions. Regardless, grind the edges smooth as in case one.

4.2.3 The burette is broken below the 50 ml but above the valve.

To fix this, you need a Biafa (fake Bic) pen and about 8cm of rubber tubing. Orange gas supply tubing is best, but hard to find. The black rubber of the inside of bicycle pump hoses also works.

- First, cut off the tip of the pen, the first 2 cm of so, and attach the non-tapered end of it to the tubing. Cutting is easiest done by scoring all the way around with a razor blade and then cleanly snapping the shaft. Remove any plastic burrs from the cut edge and then insert the wider end of the severed tip into the plastic tubing so the narrow end hangs out.

- Second, remove from the pen the little plastic end cap (the one that tells you what color ink you have) and insert it into the tubing, curved side first. Push it about half way down the tube using your fingers like esophageal peristalsis and make sure that the axis of symmetry of the pen cap stays aligned with that of the rubber tubing. That is, if the now discarded pen were still there, it would be surrounded by the tube.

- Finally, attach the other end of the tubing to the broken burette. Again, grind the sharp glass end to smooth it. What you should end up with is a burette that does not pass solution except when you press on the tubing around the pen end cap, deforming the tube to allow liquid to pass. With practice this can be easier than using a valve, and just as accurate.

Steel ball bearings are available for cheap at bicycle supply shops. These might be an alternative to the end caps of Biafa pens if you can get them in the right size. Experiment!

4.2.4 The valve is jammed

No problem! Soak it in dilute acid (not nitric) until it is free.

4.2.5 The valve is hopelessly broken.

Break the burette just above the valve and follow the instructions above. Soak a string in something flammable – kerosene, nail polish remover – and gently squeeze out the excess. Tie the string around the shaft where you want to "cut" the glass and remove the excess string. Dry up any liquid that spilled on other parts of the glass. Light the string on fire and rotate to make sure it burns evenly. After five or so seconds of burning, plunge the piece into a beaker or bucket of water. The contraction of the rapidly cooling glass should break the burette along where you tied the string. Grind the edge to smooth it.

4.2.6 The burette is broken below the valve.

This problem is mostly aesthetic, but to fix it you only need about 3 cm of rubber tubing and a clear plastic pen. Cut the tip from the pen as above and insert it into the tubing. Then stick the other end of the tubing onto the broken burette, grinding down the glass edge before you do.

4.2.7 The rubber tubing is cracking.

This usually comes from leaving clamps on the tubing during storage. To fix this, replace the rubber tubing. But while you are at it, insert a pen cap as in case three and do away with the clamps. They are more difficult to use and not as sensitive.

Chapter 5

Identifying Unknown Chemicals

Unlabelled chemicals are dangerous. If you do not know what the chemical is, then you do not know what to do if it spills, or how to safely get it out of your school.

5.1 Identifying Bottles of Unknown Liquids

Usually, these are:

- Concentrated acids (sulfuric, hydrochloric, nitric, ethanoic)
- Concentrated ammonia solution
- Organic solvents including methanol, ethanol, isobutanol, propanone (acetone), diethyl ether, ethyl ethanoate (ethyl acetate), dichloromethane, trichloromethane (chloroform), tetrachloromethane (carbon tetrachloride), trichloroethene, benzene, chlorobenzene, toluene, xylene, and petroleum spirits

Distinguishing these chemicals is important, and relatively possible. Here is a procedure:

1. **Protect yourself against whatever it might be.**
   Concentrated acids burn skin on contact and blind if they get in the eyes. Concentrated hydrochloric acid and concentrated ammonia solution release fumes that corrode the throat and lungs. Diethyl ether and propanone rapidly evaporate at room temperature and pose a significant flash fire hazard if opened near flame. Ingesting even a small amount of toxic carbon tetrachloride can be fatal, and benzene is a proven and serious carcinogen.

   Why, you might say, should I even attempt this? Because sooner or later, someone will, and better it be someone with these instructions than without. But if you do not feel comfortable, call a friend who is more excited about this process.

   Many precautions are available.

   - Tie a cloth over your mouth and nose to mitigate inhalation.
   - Find a pair of goggles or sunglasses to protect you eyes from any splash when opening the stopper in the bottle.
   - Wear gloves or at least plastic bags on your hands. Neither will protect your hands for more than a second or a few against concentrated acids or some organics, but that second can be useful in this case.
   - Thick rubber gloves are available (see Sources of Laboratory Equipment, p. 208) and offer greater protection.
   - Regardless, have at the ready a bucket of water and a box of baking soda (bicarbonate of soda) to neutralize acid burns.
   - Move the container outside and remain upwind.
   - Have a small, dry, clean beaker ready to hold a sample.

2. **Open the bottle.**
   This may be as simple as unscrewing the top or there may be an internal stopper that requires prying off.

   - Find a suitable tool, one that can pry under the cap but cut neither the cap nor you. A butter knife works well. *Do not use your fingers.*
   - When the bottle opens, look at the top. Are there white fumes? Is there an obvious smell that you can perceive from where you are standing? White fumes suggest hydrochloric acid and an intense smell could be ammonia (smells like stale urine), hydrochloric or ethanoic acid (both smell like vinegar), or an organic solvent (various odors).
   - If the contents smell obviously like ammonia, there is no need to further experimentation. Nothing else in schools smells even remotely like ammonia. Stopper that bottle and give it a good label.

- Otherwise, carefully, pour a few cubic centimeters of the liquid into your sample beaker. As you pour the liquid, observe the viscosity. Concentrated acids are all noticeably more viscous than water, especially concentrated sulfuric acid. Propanone, on the other hand, is noticeably more fluid than water. Close the bottle and take the beaker to a safe place for experimentation.
- Color is surprisingly useless in identifying unknown liquids because most readily take on color from even small amounts of contamination.
- Rest the beaker on a sturdy surface. If you have already noticed an intense smell, leave the cloth on your face. If you have not yet noticed a smell, remove it.

5.2 Test one: Add to water

- Fill a large, clean test tube half way with ordinary water. Alternatively, find the smallest beaker you have (probably 50 mL), and fill it about a quarter of the way with water. Carefully pour in a few drops of your unknown and observe what happens.
- If it does not mix with the water, instead forming a new (possibly quite small) layer on top, you have an *organic solvent less dense than water*, probably one of: isobutanol, diethyl ether, ethyl ethanoate, benzene, chlorobenzene, toluene, xylene, or petroleum spirits.
- If it does not mix with the water, instead sinking to form a distinct layer on the bottom, you have an *organic solvent more dense than water*, probably dichloromethane, chloroform, or carbon tetrachloride.
- If your unknown does not mix with water, jump down to Test four: What kind of organic? on what to do with organics.
- If the unknown seems to sink into the water but not mix completely, you probably have a *concentrated acid*. The test tube might even get a little warmer. You might also have a very concentrated solution of some other solute, left over from a previous experiment.
- If the unknown seems to mix into the water like, well, water, you probably have an *aqueous solution* that is not very concentrated. It might be dilute acid, dilute hydroxide, hydrogen peroxide solution, etc. – more work lies ahead.

5.3 Test two: Is it an acid?

This only applies to solutions that mix completely into water.

1. This is easy with a piece of blue litmus paper. Dip a corner down into the test tube or beaker.
   - If it turns bright red, you probably have an *acid*, and if your liquid was noticeably viscous, a concentrated acid.
   - If there is no change, move on to Test five: What else?.
2. Another option is universal indicator or universal pH paper.
   - Prepare a 100-fold dilution of the original acid and test with the indicator.
   - If the color is bright red, you must have a *strong acid*, like hydrochloric, sulfuric, or nitric acid.
   - If the color is instead orange or yellow, you must have a *weak acid*, like ethanoic acid.
3. If there is no universal indicator, you can show that something definitely is an acid if it causes methyl orange to turn from orange to red. However, if there is no color change, you might still have a weak acid, so you cannot use methyl orange to eliminate the possibility of an acid.
4. You also cannot use POP to show that there is an acid, as both concentrated acid and tap water have the same effect on POP: none whatsoever.
5. If you do not have any litmus paper or other indicator, find another beaker and add 10-20 mL of ordinary water and dissolve a bit of baking soda (bicarbonate of soda). Carefully, with eye protection, add a few drops of your DILUTED unknown (from test one). If there are bubbles, you have an *acid*. Adding a concentrated acid directly to baking powder can cause such vigorous effervescence as to eject acid from the test tube.

5.4 Test three: What kind of acid?

5.4.1 Sulphuric acid

- Hints: obviously viscous, significantly denser than water, noticeable heat released on dilution, no smell
- Confirmatory test: dip the wooden end of a match into the original solution. If the end appears to char, you have concentrated sulfuric acid. Another variant of this test is take some concentrated sulfuric acid and pour over some sugar in a beaker. After some time, a black color from carbon produced from the dehydration of sugar confirms sulfuric acid. Yes, the same thing happens to skin. The downside of this second test is that the beaker is almost impossible to clean.
- Alternative test: Find or prepare a 0.1 M barium nitrate, barium chloride, or lead nitrate solution. In a test tube, add about one centimeter of your diluted sample and then a few drops of one of the above solutions. An instant, white, cloudy precipitate demonstrates that sulfate is present. To confirm that this is from sulfuric acid and not, say, your tap water, test in the same way the water you used for the dilution. Not much should happen. If your tap water contains sulfates, find some distilled (e.g. rain) water and remake the dilution.

5.4.2 Hydrochloric acid

- Hints: white fumes, intense acidic smell similar to vinegar, more dense than water
- Confirmatory test: prepare a dilute potassium permanganate solution. This should be pink in color, which might require significant dilution. Fill a test tube with a couple centimeters of your dilute solution and add the potassium permanganate solution drop wise. If the pink color is rendered colorless after mixing with your diluted sample, you probably have hydrochloric acid. This reaction makes small amounts of chlorine gas, but that poses much less risk than the hydrochloric acid fumes.
- Alternative test: Find or prepare a 0.05 M or 0.1 M silver nitrate solution. Remember that this chemical is very expensive, so only make a small quantity. In a test tube, add about one centimeter of the water you used for diluting your sample and then a few drops of one of silver nitrate solution. An instant, white to gray, cloudy precipitate demonstrates that chloride is present. If this happens, your tap water contains chlorine and you will have to prepare another dilution using rain or distilled water. If the water you used for dilution lacks chlorine, add a centimeter of the diluted sample to a clean test tube and add a few drops of silver nitrate solution. The precipitate confirms that you have hydrochloric acid. Note that for this test to be effective, the hydrochloric acid must be diluted. Concentrated hydrochloride acid reacts with aqueous silver to form the $[AgCl_2]$-complex, which is soluble.

5.4.3 Ethanoic (acetic) acid

- Confirmatory Test: This acid smells strongly of vinegar. If you have a definite vingar smell, it is probably ethanoic acid, but beware that concentrated HCl can have a similar smell. To confirm ethanoic acid, use some diluted acid from test one and add a small amount of baking soda until it is just neutralized. Do not add excess baking soda - neutralization is the goal. After neutralizing, add a small amount of iron (III) chloride or nitrate. A blood red solution of iron (III) acetate proves that the acid is ethanoic. Boiling the solution should form a red brown precipitate. If you do not have iron (III) salts but do have universal indicator, use the indicator method above for confirming that your unknown is a weak acid – ethanoic is the only common weak acid that smells like vinegar.

5.4.4 Nitric acid

A concentrated acid in a school that does not smell like vinegar and is not hydrochloric or sulfuric acid is very likely to be nitric acid.

- Confirmatory Tests: Take a wooden splinter or match stick and dip it in the concentrated acid. If the splinter turns yellow, the acid is nitric. A second confirmatory test is adding copper wire or

turnings to the concentrated acid. A brown gas of nitrogen dioxide is formed. Do this confirmatory test in a well ventilated area.

- Special note: if you suspect nitric acid, dip a piece of copper wire into the solution. If it comes back with a silvery coating, you have Million's Reagent, mercury metal dissolved in nitric acid. This is highly toxic, very dangerous, and should never be used in a school. Label the bottle "Million's Reagent, Contains Hg_2^+, TOXIC, CORROSIVE, do not use, do not dump" along with similar warnings in any local language(s) and find a safe place to store it.

5.5 Test four: What kind of organic?

Let us be honest. Distinguishing between different kinds of organic solvents is hard with the resources that are probably available.

- If the chemical is more dense than water and no one at the school claims that it is chloroform (trichloromethane) for the biology lab, there is no way to show that it is not carbon tetrachloride (tetrachloromethane), a toxic organic solvent responsible for the death students in several countries. Label the bottle "Unknown organic solvent more dense than water, possibly carbon tetrachloride, TOXIC, never use, never dump," with similar warnings in any local language(s) and find a safe place to store it.

- If the chemical is less dense than water and you are familiar with organic solvents, you might try a careful smell test.

- If the unknown smells like strong booze and is soluble in water, it is probably ethanol or methanol. Do not drink it! – methanol blinds.

- If it is bright red, it is probably Sudan III solution, for biology. Label and use it.

- If it is yellow or brown it might be iodine solution, see below in test five.

- If it is light purple or green or whatever the popular color in your country, it is probably methylated spirits, a mixture of about 70% ethanol and 30% water with some impurities to make it undrinkable. Confirm this by showing that paper soaked in the chemical will burn with a blue flame but that paper soaked in a 50/50 mixture of the chemical and water will not burn.

- If it is clear and someone at the school can assure that the contents are ethanol and not methanol, label the bottle "ethanol" and use it.

- If the bottle might be methanol, a poison, pour the contents into a large bucket and leave it in a place where no one will disturb it and where the fumes will not accumulate. Let it evaporate.

- If the unknown smells like nail polish remover and is soluble in water, it is probably propanone (acetone).

- If you put a drop in a spoon it should evaporate relatively quickly. Label it "Propanone, EXTREMELY FLAMMABLE" and keep it around.

- If it is not soluble in water and smells like magic markers, it is probably diethyl ether or ethyl ethanoate (ethyl acetate). If you are familiar with organics, perhaps you can pick between these. Otherwise just label the bottle "volatile organic solvent, insoluble in water, EXTREMELY FLAMMABLE" and keep it around.

- If the unknown has a sweet sickly smell it might be toluene. It also might be benzene.

- If you cannot further identify it and no one else can, label the bottle "unknown non-volatile organic solvent less dense that water, possibly benzene, TOXIC, never use, never dump" with similar warnings in any local language(s), and find a safe place to store it.

5.6 Test five: What else?

If your unknown is not an acid, not ammonia, and soluble in water, see what it smells like.

- If it smells like booze or nail polish remover, it could be methanol, ethanol, or acetone. See the above section on organics.
- If it does not have a smell, it is probably a solution left over from an earlier lab. These are not nearly as dangerous as concentrated acids or some volatile organics. However, be sure to use proper handing methods.

  Here are some possibilities:

5.6.1 Sodium hydroxide solution

- Hints: cloudy and a jammed stopper, but not always
- Test: red litmus turns blue or POP pink.
- What to do: sodium hydroxide is cheap when bought as caustic soda. Keep it around just for neutralizing acid wastes. If its presence disturbs you, add some indicator and then cheap acid until neutralization. After complete neutralization, dump.

5.6.2 Hydrogen peroxide

- Hints: colorless liquid, in an opaque or dark bottle
- Test: add a bit of acidified potassium permanganate solution. The potassium permanganate should turn colorless on mixing and bubbles of gas should be observed.
- What to do: label and use. If you want to dispose of it for some reason, leave it in a bucket in the sun.

5.6.3 Potassium permanganate solution

- Hints: intensely purple, pink after significant dilution
- Test: to a very dilute solution, add crushed vitamin C (ascorbic acid) from a pharmacy. The solution should turn colorless.
- What to do: test the pH with litmus paper or methyl orange to see if acid has been added. Then label "(acidified) potassium permanganate" and use. If you want to dispose of it, add crushed vitamin C until the color disappears and then pour into a pit latrine.

5.6.4 Iodine solution

- Hints: brown color, smells like iodine tincture, and possibly also like ethanol
- Test: to a dilute solution, add crushed vitamin C (ascorbic acid) from a pharmacy. The solution should turn colorless.
- What to do: Put a centimeter of water in a test tube followed by a smaller quantity of cooking oil. Add a few drops of the solution, cap with your thumb and mix thoroughly for one minute. If two layers quickly separate, the iodine solution has been prepared without ethanol. If a cloudy mixture (an emulsion) forms, the iodine solution has been prepared with ethanol. Label the solution "iodine solution (with ethanol)" and use it.

5.6.5 Potassium ferrocyanide solution

- Hints: light neon green or yellow color
- Test: make a dilute solution of copper sulfate and add a few drops of the unknown. An instant, massive brown precipitate confirms potassium ferrocyanide solution.
- What to do: Label and use. Do not dump while it remains useful.

5.7 Unidentifiable Liquids

...are worthless. In order to safely dump a liquid chemical, ensure the following are true:

- The liquid is water soluble (otherwise see the organic section above)
- The liquid is neutral pH (if acid, neutralize with bicarbonate of soda, if base neutralize with acid waste, citric acid, or, carefully, battery acid)
- The liquid does not contain heavy metals (to a small sample, add dilute sulfuric acid drop-wise. A precipitate indicated lead or barium. Continue adding until additional precipitation stops. Then neutralize with bicarbonate of soda. The solids are safe for disposal in a pit latrine, but may clog sink pipes).
- The liquid does not contain mercury (to a small sample, add sodium hydroxide solution until POP turns the solution pink. A yellow precipitate indicates mercury. Label the solution "Contains Hg_2^+, TOXIC, do not use, do not dump" and store it in a safe place.)

Then, dilute the chemical in a large amount of water and dispose of it in a lab sink or pit latrine.

5.8 Deliquescent Salts

If you have a chemical in a container that seems meant for holding solids but the chemical looks like a thick liquid, you probably have a deliquescent salt that fully deliquesced. These solutions can be quite dangerous because they are maximally concentrated. Make sure that no unknown chemicals touch your skin, and wear goggles for this work. Then, do the following:

5.8.1 Test for a base

The most common deliquescent salt is sodium hydroxide.

- Fill a test tube half way with water and a few drops of the unknown syrup followed by a few drops of POP indicator. If the solution turns pink, you almost certainly have either *sodium hydroxide* or *potassium hydroxide*.
- Dilute the liquid in at least 10 times its rough volume of water and titrate a sample against 1 M acid.
- Find a plastic water bottle with a screw cap for your dilution and label it "sodium or potassium hydroxide, n M", where n is the molarity you measured in your titration.

5.8.2 Color

If the liquid is not a base, it is probably a chloride or nitrate salt of one element or another.

- If it is colorless, the cation is probably in Group IIA (Ca, Sr, or Ba) or Group IIB (Zn, etc). Group IIA compounds have distinct flame test colors:
  - Ca = orange red,
  - Sr = bright red,
  - Ba = apple green

  while Zn has no flame test color.
- If it is red or brown, it is probably iron (III) nitrate or iron (III) chloride.
- If it is intensely pink, it might be cobalt.

To identify the compound completely, you will have to perform qualitative inorganic analysis. An introduction to the art is in the Qualitative Analysis (p. 106) section of this manual, and more advanced methods are available on the internet and in some advanced chemistry books.

5.8.3 Check for mercury

- If the liquid is not a base, dilute a small sample in water and add sodium hydroxide solution until POP turns pink.
- If a bright yellow precipitate forms, you probably have a mercury salt. Transfer all of the compound to a sturdy container with a well-sealing lid, wash the original container with minimal water once and add the washings to the storage container. Then wash the original container and anything the liquid touched thoroughly. Label the new storage container "Solution of unknown mercury salt, CONTAINS Hg!!, TOXIC! Do not use, Do not dump," along with appropriate warnings in any local language(s), and find a safe and secure place for long term storage.

5.9 Identifying Unknown Solid Chemicals

This is not nearly is important as identifying unknown liquids for two reasons:

1. These chemicals are generally (though not always!) less dangerous.
2. Accidental spills are less dramatic.

The smallest containers are the most likely to hold dangerous chemicals, like mercury salts. It is best to just leave these ones alone.

What you can do is look at the solid and see if it matches any of the descriptions below. Color is much more useful for identifying solids.

- *Bright orange crystals* are likely a chromate or dichromate salt (toxic) or a ferricyanide salt (much less toxic). The later will form an intensely blue precipitate with a small amount of Fe_2^+, perhaps from iron (II) sulfate. Chromates form a yellow solution that turns orange on addition of acid while dichromates for an orange solutions that turns yellow on addition of base.
- *Bright yellow, orange, or red powders* might be lead or mercury compounds. These are poisonous, the latter very. It also might be methyl orange powder.
  - Try to dissolve a small amount in water. Methyl orange will dissolve readily to give a bright orange solution, one that turns red in acid and yellow in base. Label the powder and keep it around.
  - If the salt dissolves but does not seem to be methyl orange, add sodium hydroxide until POP changes color.
  - A yellow precipitate suggests mercury. Label as with mercury compounds encountered above.
  - Most lead compounds are not soluble, and will not form a color in solution. Other mercury compounds are also insoluble. Label a container that might be lead or mercury as "possible lead or mercury compound, POISON," and store it for the long haul.
- *A yellow powder insoluble in water* may also be sulfur. It should smell like sulfur. A small amount will dissolve in kerosene, and the dry powder will melt when heated in a spoon over a flame and then burn with a blue flame – producing sulfur dioxide, a poisonous gas. Do not heat an unknown yellow compound unless you are fairly sure it is sulfur.
- *Blue compounds* are often copper salts. These should have a green flame test.
- *Purple crystals or flakes* insoluble in water are probably iodine. Iodine will dissolve in kerosene to form a red solution.
- One of the few *green powders* is nickel carbonate.
- *Pink wet looking crystals* might be a cobalt compound. Heat them gently in a spoon and they should dehydrate to turn blue. The blue crystals should turn pink when dissolved in water. Cobalt is poisonous.
- *Crystals so purple they look brown or yellow* are probably potassium permanganate. They should form an intensely purple solution in water. Confirm as with potassium permanganate solution above.

- *White crystals and powders* are really hard to identify. Label them "unknown white powder/crystals" and move them to a safe and secure place.
- Flat dull *gray metallic ribbon* about 5 mm wide and 1 mm thick is probably magnesium metal. It should turn shiny if polished with steel wool. It will also burn with a very intense white light if lit in either a Bunsen burner or gas cigarette lighter. Hold it with tongs, and do not stare at the light.
- A *metal stored under oil* is probably sodium or potassium. If you are feeling adventurous, remove a sample and cut off a VERY small piece, perhaps 5 mm on a side. Both metals may be easily cut with a knife. Return the rest to the original container and seal it again. Then, add the piece of metal to an open container of water and stand back. Both react violently and generally send the piece of metal spinning around on a cushion of hydrogen gas. Potassium generally gets hot enough to ignite this gas which then burns with a lilac flame. If the hydrogen under sodium burns, it will be yellow. The water will become a solution of sodium or potassium hydroxide.

Part II

Laboratory Safety

Chapter 6

Guidelines for Laboratory Safety

There is no excuse for laboratory accidents. Students and teachers get hurt when they do something dangerous or when they are careless. If you do not know how to use a substance or a tool safely, do not use it. If your students do not know how to use a chemical or a tool safely, do not let them use it until they do. Adopt a zero tolerance policy towards truly unsafe behavior (running, fighting, throwing objects, etc.) – first infraction gets students kicked out of class for the day. Explain the error to everyone to make sure that it is never repeated. If the same student errs again, expel him for longer. Make it clear that you will not tolerate unsafe behavior.

Remember, the teacher is responsible for everything that happens in the lab. If a student is hurt the teacher is to blame. Either the teacher did not understand the danger present, did not adequately prepare the laboratory or the lesson, did not adequately train the student in safe behavior, or did not offer adequate supervision. As a teacher, you must know exactly the hazards of your chemicals, tools, and apparatus. Explain these hazards clearly and concisely to your students before they touch anything.

The following rules are for everyone in the lab to follow – students, teachers, and visitors alike. We recommend painting them directly on the wall as most paper signs eventually fall down.

6.1 Basic Lab Rules

1. Wear proper clothes. For every practical, wear shoes. Sandals are not acceptable lab ware. If you are pouring concentrated chemicals, you need to wear safety goggles.

2. Nothing enters the mouth in the lab. This means no eating, no drinking, and no mouth pipetting.

3. Follow the instructions from the teacher. Obey commands immediately. Only mix chemicals as instructed.

4. If you do not know how to do something or what to do, ask the teacher.

In addition to these rules, we recommend a variety of guidelines for teachers and lab managers to keep the lab a safe place.

6.2 Specific Guidelines to Reduce Risk

1. Never use the following chemicals:

   1.1. Organic liquids, including:

      1.1.1. Benzene (C_6H_6)
      1.1.2. Chlorobenzene (C_6H_5Cl)
      1.1.3. Dichloromethane (CH_2Cl_2)
      1.1.4. Tetrachloromethane/carbon tetrachloride (CCl_4)
      1.1.5. Trichloroethane (CH_3CCl_3)
      1.1.6. Trichloromethane/chloroform ($CHCl_3$)

   1.2. Anything containing mercury:

      1.2.1. Mercury metal (Hg)
      1.2.2. Mercurous/mercuric chloride ($HgCl/HgCl_2$)
      1.2.3. Million's Reagent ($Hg + HNO_3$)
      1.2.4. Nestler's Reagent ($HgCl_2$ + others)
      1.2.5. For more information about these chemicals, their risks, and what to do if you find them in your lab, see Laboratory Safety: Dangerous Chemicals (p. 37).

2. Do not make hazardous substances

2.1. Chlorine gas - electrolysis of chloride salts, oxidation of chloride salts or hydrochloric acid by oxidizing agents such as bleach or potassium permanganate

2.2. Chloroamines - ammonia with bleach. People have died mixing ammonia and bleach together when mixing cleaning agents.

2.3. Hydrogen cyanide - cyanide salts, including ferro- and ferri-cyanide, with acids.

3. Avoid hazardous substances

3.1. If you have a choice, use non-poisonous substances. To be a good teacher, the only poisons that you have to use are those required by the national exams. For all other activities, use less dangerous substances.

3.2. Only give students small quantities of required poisons.

3.3. For advice on handling the various required poisons, see Laboratory Management: Dangerous Chemicals.

4. Avoid explosions

4.1. Never heat ammonium nitrate.

4.2. Never heat nitrates in the presence of anything that burns.

4.3. Never heat a closed container.

4.4. If performing a distillation or other experiment with boiling or hot gases, make sure that there is always an unobstructed path for gases to escape.

5. Avoid fires

5.1. Be careful!

5.2. Keep all flammable materials away from flames. Never have the following very flammable chemicals in the same room as fire: propanone (acetone), ethyl ethanoate (ethyl acetate), diethyl ether.

5.3. Keep stoves clean and in good working order. Do not douse stoves with water to extinguish them because the metal will corrode much faster (think kinetics). There is never a need for this. If the stove does not extinguish on its own, you should repair it so it does.

5.4. Only use the appropriate fuel for a given stove. For example, never put petrol in a kerosene stove.

6. Avoid cuts

6.1. Only use sharp tools when required, and design activities to minimize use of sharp tools.

6.2. Keep sharp tools sharp. The only thing more dangerous than cutting with a sharp knife is cutting with a dull one.

6.3. Use the right tool for cutting.

6.4. Use as little glass as possible.

6.5. Do not use broken glass apparatus. The last thing you want to deal with during a practical is serious bleeding. It is tempting to keep using that flask with the jagged top. Do not. Do not let anyone else use it either – break it the rest of the way.

6.6. Dispose of sharp trash (glass shards, syringe needles) in a safe place, like a deep pit latrine.

7. Avoid eye injuries

7.1. Students should wear goggles during any activity with a risk of eye injury. See the Materials: Apparatus section for suggestions on goggles. If you do not have the goggles necessary to make an experiment safe, do not do the experiment.

7.2. Keep test tubes pointed away from people during heating or reactions. Never look down a test tube while using it.

7.3. Never wear contact lenses in the laboratory. They have this way of trapping harmful chemicals behind them, magnifying the damage. Besides, glasses offer decent (though incomplete) protection on their own.

8. Avoid chemical spills
   - 8.1. Teach students that if they get chemicals on their hands, they should wash them off immediately, without asking for permission first. Some students have been taught to wait for a teacher's permission before doing anything in the lab, even if concentrated acid is burning their hands. On the first day, give them permission to wash their hands if they ever spill chemicals on them.
   - 8.2. Also, teach students to tell you immediately when chemicals are spilled. Sometimes they hide chemical spills for fear of punishment. Do not punish them for spills - legitimate accidents happen. Do punish them for unsafe behavior of any kind, even if it does not result in an accident.
9. Use adequate protection with hazardous chemicals.
   - 9.1. Wear eye protection (see above). Find goggles or things that will substitute.
   - 9.2. Tie a cloth over your face when using concentrated ammonia or HCl. For the latter chemical, see below.
   - 9.3. Sulfuric Acid, H_2SO_4
     - 9.3.1. There is never any reason to ever use fully concentrated (18 M) sulfuric acid.
     - 9.3.2. For qualitative analysis, 5 M H_2SO_4 is sufficient for "concentrated sulfuric acid."
     - 9.3.3. Do not buy 18 M sulfuric acid. Battery acid will suffice for qualitative analysis and is a much safer (if still quite dangerous) source of sulfuric acid.
     - 9.3.4. If you already have 18 M sulfuric acid in your lab, just leave it. Battery acid is so cheap you can afford to get as much as you need.
   - 9.4. Hydrochloric acid, HCl
     - 9.4.1. Hydrochloric acid is never required.
     - 9.4.2. Do not buy concentrated hydrochloric acid. Use battery acid for all of its strong acid applications.
     - 9.4.3. When you need the reducing properties of HCl, for the precipitation of sulfur from thiosulfate in kinetics experiments for example, make a solution with the proper molarity of chloride and H^+ by dissolving sodium chloride in battery acid and diluting with water.
   - 9.5. Nitric acid, HNO_3
     - 9.5.1. The only time nitric acid is required is to dissolve certain carbonates in qualitative analysis. The first time you need nitric acid, prepare a large volume of dilute acid (e.g. 2.5 L) so that you do not need to handle the concentrated acid again.
     - 9.5.2. If many schools share a single bottle of concentrated acid, they should dilute it at a central location and transport only the dilute acid.
     - 9.5.3. Teach qualitative analysis of insoluble carbonates using copper, iron, or zinc carbonate – these will dissolve in dilute sulfuric acid.
10. Avoid mouth pipetting
    - 10.1. Never do it!
    - 10.2. This is a dangerous activity prohibited in every modern science laboratory.
    - 10.3. Use rubber pipette filling bulbs or plastic syringes.
    - 10.4. For more explanation, see Mouth pipetting in Dangerous Techniques (p. 43).
11. Be prepared
    - 11.1. Set aside a bucket of water for first aid.
    - 11.2. It should not be used for anything else.
    - 11.3. Have materials to fight fires and know how to use them.
    - 11.4. A bucket of sand will work for any lab fire, is available to every school, and can be used by anyone.
12. Use good habits

12.1. Hand washing

12.1.1. Students should wash their hands every time they leave the lab.

12.1.2. Always have water and soap available, ideally in buckets on a desk near the door.

12.1.3. Even if students do not touch any chemicals when they are in the lab, they should still wash their hands.

12.2. Clean all benches and chemicals

12.2.1. Stray chemicals and contaminated apparatus has the potential for danger.

12.2.2. Make sure students do not leave stray pieces of paper.

12.2.3. Ensure all students clean the apparatus they use immediately after use.

12.2.4. Have students to clean apparatus prior to use. It is not always possible to trust the students washed the apparatus after their last use.

12.3. Tasting chemicals

12.3.1. Students should never eat anything in the lab. Ever.

12.3.2. Barium nitrate looks just like sodium chloride. Lead carbonate looks like starch.

12.3.3. Do not bring food into the lab.

12.3.4. If you use domestic reagents (vinegar, salt, baking soda, etc.) in the lab, label them and leave them in the lab.

12.4. Smelling chemicals

12.4.1. Be aware that many chemicals give of fumes that can produce obnoxious odors or be irritating to the respiratory system.

12.4.2. Practicals involving nitrates, chlorides, ammonium compounds, and some sulphates produce harmful gases.

12.4.3. Open the lab windows to maximize airflow.

12.4.4. Kerosene stoves also produce noxious fumes - it is much better to use motopoa.

12.4.5. If students feel dizzy or sick from the fumes, let them go outside to recover.

12.4.6. Many lab reagents – ammonia, hydrochloric acid, nitric acid, ethanoic (acetic) acid – can cause serious damage if inhaled directly.

12.5. Keep bottles and other apparatus away from the edge of the table. Twenty centimeters is a good rule.

12.6. Cap reagent botles when they are not in use.

12.7. Do not do things you do not want your students to do. They are always watching, always learning.

13. Dispose of wastes properly

13.1. See Laboratory Management: Waste Disposal (p. 50)

Chapter 7

First Aid

In spite of taking all necessary precautions to avoid dangerous situations in the laboratory, emergencies may still arise which require the immediate use of First Aid techniques. Listed below are various types of possible emergencies, as well as some immediate treatment guidelines to follow until professional medical attention may be given to the victim.

For treatment information relating to specific chemicals, refer to the section on Dangerous Chemicals (p. 37).

7.1 Cuts

1. Immediately wash cuts with lots of water to minimize chemicals entering the blood stream.
2. Then wash with soap to kill any bacteria that may have entered the wound.
3. To stop bleeding, apply pressure to the cut and raise it above the heart. If the victim is unable to apply pressure him/herself, remember to put something (gloves, a plastic bag, etc.) between your skin and their blood.
4. If the cut is deep (might require stitches) seek medical attention. Make sure that the doctor sees how deep the wound really is – you might do such a good job cleaning the cut that the doctor will not understand how serious it is.

7.2 Eyes

1. If chemicals get in the eye, immediately wash with lots of water.
2. Keep washing for fifteen minutes.
3. Remind the victim that fifteen minutes is a short time compared to blindness for the rest of life. Even in the middle of a national exam.

7.3 First and Second Degree Burns

1. Skin red or blistered but no black char.
2. Immediately apply water.
3. Continue to keep the damaged skin in contact with water for 5-15 minutes, depending on the severity of the burn.

7.4 Third Degree Burns

1. Skin is charred; there may be no pain.
2. Do not apply water.
3. Do not apply oil.
4. Do not removed fused clothing.
5. Cover the burn with a clean cloth and go to a hospital.
6. Ensure that the victim drinks plenty of water (one or more liters) to prevent dehydration.

7.5 Chemical Burns

1. Treat chemical burns by neutralizing the chemical.
2. For acid burns, immediately apply a dilute solution of a weak base (e.g. sodium hydrogen carbonate).
3. For base burns, immediately apply a dilute solution of a weak acid (e.g. citric acid, ethanoic acid). Have these solutions prepared and waiting in bottles in the lab.

7.6 Ingestion

1. If a student ingests (eats or drinks) the following, induce vomiting.
   1.1. Barium (chloride, hydroxide, or nitrate)
   1.2. Lead (carbonate, chloride, nitrate, oxide)
   1.3. Silver (nitrate)
   1.4. Potassium hexacyanoferrate (ferr[i/o]cyanide)
   1.5. Ammonium ethandioate (oxylate)
   1.6. Anything with mercury (see list above), but mercury compounds should just never be used.
2. To induce vomiting:
   2.1. Have the student put fingers into his/her throat
   2.2. Have the student drink a strong solution of salt water (use food salt, not lab chemicals)
3. Do not induce vomiting if a student ingests any organic chemical, acid, base, or strong oxidizing agent.
   3.1. These chemicals do most of their damage to the esophagus and the only thing worse than passing once is passing twice.
   3.2. Organic chemicals may be aspirated into the lungs if vomited, causing a sometimes fatal pneumonia-like condition.

7.7 Fainting

1. If a student passes out (faints), feels dizzy, has a headache, etc., move him/her outside until fully recovered.
2. Check unconscious students for breath and a pulse.
3. Perform CPR if necessary and you know how.
4. Generally, these ailments suggest that harmful gases are present in the lab – find out what is producing them and stop it. Kerosene stoves, for example, may emit enough fumes to have this effect.
5. See Sources of Heat in the Materials section for alternatives.
6. Chemicals reacting in drain pipes can also emit harmful gases. See Waste Disposal.

7.8 Electrocution

– If someone is being electrocuted (their body is in contact with a live wire)

1. First disconnect the power source. Turn off the switch or disconnect the batteries.
2. If that is not possible, use a non-conducting object, like a wood stick or branch, to move them away from the source of electricity.
3. Unless there is a lot of water around, the sole of your shoe is non-conducting.

7.9 Seizure

1. If a student experiences a seizure, move everything away from him/her and then let the body finish moving on its own.

Chapter 8

Dangerous Chemicals

8.1 Chemicals that should never be used in a school

8.1.1 Mercury and its compounds (e.g. Million's Reagent, Nestler's Reagent)

- Hazard: Toxic
- Route: Ingestion of solutions and salts; inhalation of vapors from the liquid metal. Mercury has a very low vapor pressure, but the vapors that do form are quite poisonous – inhalation is therefore a significant hazard.
- Use: Showing off to students, Million's reagent for biology (no longer used)
- Alternatives: Use the biuret test to detect proteins (1 M NaOH followed by 1% $CuSO_4$, a purple color is a positive result)
- Precautions if it needs handling (e.g. broken thermometers): Wear gloves or plastic bags on the hands.
- First aid: If ingested, induce vomiting at once. Administer activated charcoal. Seek medical attention.
- Disposal: If you find mercury or its compounds, keep them in sealed in a bottle and locked away. Label the bottle very clearly "POISON, DO NOT OPEN, DO NOT DUMP" and also include a strong warming in the local language(s). If you spill liquid mercury, ask everyone to leave and apply powdered sulfur immediately. Put on gloves and tie a cloth on your face. Open windows to increase ventilation. Then use pieces of cardboard to gather the mercury back together so you can seal it in a bottle. Apply powdered sulfur to any mercury that cannot be reached – e.g. cracks in the floor.

8.1.2 Benzene

- Proven carcinogen, toxic. A horrific and generally fatal form of cancer is associated with benzene exposure, with tumors appearing rapidly throughout body.
- Route: Can be fatal if ingested, especially if aspirated into the lungs (e.g. if mouth pipetting); also passes through skin(!)
- Use: Multi-purpose non-polar solvent. Less dense than water.
- Alternative: kerosene
- Precautions if it needs handling: Thick rubber gloves. It will pass rapidly through latex.
- Disposal: If you find a bottle of benzene, leave it sealed and in a secure place with a stern warning label. If a bottle breaks, evacuate the room and return only wearing a cloth over your face and thick rubber gloves. Absorb the benzene with cardboard, cotton wool, saw dust, rice hulls, or flour, transfer the mass to a dry place outside, add a significant amount of kerosene and burn completely. Benzene will combust on its own, but you want to make sure it burns hot enough that none simply vaporizes without combusting.

8.1.3 Tetrachloromethane (carbon tetrachloride)

- Hazard: Proven carcinogen. The chemical has killed students in both Tanzania and Kenya.
- Route: Ingestion can be fatal. Will pass through skin. Inhalation of vapors is quite dangerous.
- Use: Multi-purpose non-polar solvent. More dense than water.

- Trichloromethane (chloroform) is another non-polar solvent more dense than water, though still dangerous (listed in category two, below). If the density is not important, use kerosene. If the solvent must not be flammable, consider a different experiment.
- Precautions if it needs handling: Thick rubber gloves. It will pass rapidly through latex.
- First Aid: Seek medical attention immediately. Ask a medical expert if you should induce vomiting (the chemical can kill if absorbed through the stomach, but also if aspirated into the lungs when vomiting)
- Disposal: If you find a bottle of carbon tetrachloride, leave it sealed and in a secure place. If a bottle breaks, absorb the chemical with cotton wool and move the cotton to a place where it can off-gas away from people and other living things. Protect from rain and from leaching into the ground. Once the cotton is completely dry, douse with kerosene and burn it. Do not burn the chemical directly – it used to be used in some fire extinguishers. On heating, it decomposes to release poisonous gases.

8.1.4 Other hazardous organic solvents

The following chemicals have hazards similar to if less severe than benzene and tetrachloromethane. None should ever be used in a school. Leave them sealed in their bottles. If a bottle breaks, follow the instructions listed with tetrachloromethane.

- Chlorobenzene
- Dichloromethane
- Trichloroethane

8.2 Dangerous chemicals that you might need to use

8.2.1 Ammonia (ammonium hydroxide solution)

- Hazard: The liquid burns skin, the fumes burn lungs, and reaction with bleach or any combination of chloride and oxidizer can form toxic chloroamine fumes.
- Use: Common bench reagent.
- Alternative: For a simple weak base, use carbonate or hydrogen carbonate.
- Precaution: Strongly prohibit mixing of different bench reagents. Neutralize waste completely before disposal. When pouring ammonia for distribution, wear cloth over your mouth and nose and work outside, upwind. To smell, waft carefully – never inhale ammonia directly from a bottle!
- First Aid: In case of skin contact, wash with plenty of water followed by a dilute weak acid (vinegar or dilute citric acid) and more water. In case of eye contact, wash with water for at least ten minutes. If ingested, do NOT induce vomiting. In case of inhalation, move victim to fresh air. Seek medical attention if the victim does not recover quickly.
- Disposal: Save unused solution for another day. If you must dispose of it, add to several liters of water and leave in an open bucket in the sun, away from people and animals. The ammonia will evaporate, leaving water behind. The process is finished when the bucket no longer smells like ammonia.

8.2.2 Barium

- Hazard: Very poisonous if ingested in a soluble form (e.g. barium chloride, hydroxide, or nitrate). Note that barium carbonate will dissolve very quickly in stomach acid.
- Use: Preparation of hydrogen peroxide, test for sulfates, flame tests.

- Alternatives: hydrogen peroxide is often sold in pharmacies, sulfates may be confirmed with soluble lead salts (also poisonous!), and boron compounds (e.g. boric acid, borax) also produce a green flame color.

- Precautions: Distribute only in small quantities in bottles clearly labeled "POISON." Also use the local word for poison, e.g. SUMU in Swahili. Collect all barium waste in a special container. This will require training students. Have a bottle of magnesium sulfate or sodium sulfate available for spills on skin or tables. Sodium sulfate can be prepared by neutralizing dilute sulfuric acid with sodium bicarbonate – err on the side of excess bicarbonate. See Sources of Chemicals for magnesium sulfate.

- First Aid: If ingested, induce vomiting and administer activated charcoal if available. Go to the hospital. The material safety data sheet for barium chloride recommends use of sodium or magnesium sulfate under a doctor's direction. Chemically, this would precipitate barium sulfate, preventing absorption of the element. Magnesium sulfate is non-toxic, though will probably cause diarrhea.

- Disposal: Collect unused solutions for another day. Collect all waste in a large container and add dilute sulfuric acid until precipitation stops. Pour off most of the liquid and treat it as dilute acid waste. Use the remaining liquid to send the precipitate to the bottom of a pit latrine.

8.2.3 Chloroform (Trichloromethane)

- Hazard: used to render mammalian specimens unconscious; it has the same effect on humans. Also toxic in ingested. Passes through skin.

- Use: Knocking out dissection specimens, sometimes as a specialty non-polar solvent.

- Alternatives: Dissect dead specimens; use safer solvents.

- Precautions: Work in a well-ventilated space, like outside. NEVER, EVER MOUTH PIPETTE!

- First Aid: If ingested, go to the hospital. Do not induce vomiting unless instructed by a medical professional. If inhaled, immediately remove the victim to fresh air and sit (but not lie) him or her down in case of fainting. If the victim loses consciousness, go to the hospital. In both cases, monitor breathing and pulse. In case of skin contact, wash off immediately, and use soap as soon as it is available.

- Disposal: For the small amounts used in preparing specimens for dissection, allow the chemical to evaporate away from people and animals. For large amounts, e.g. if a bottle spills or breaks, evacuate the room and keep everyone away for at least one day. Return carefully, allowing more time if the room still smells like trichloromethane.

8.2.4 Concentrated Acids

(sulfuric, hydrochloric, nitric, ethanoic (acetic))

- Hazard: Serious skin burns, will blind in the eyes.

- Use: Often the starting material when preparing dilute acids for titrations or food tests. Also used directly in small quantities in chemical qualitative analysis.

- Alternatives: If any acid will suffice, use a safer weak acid, e.g. citric acid (best) or ethanoic (acetic) acid. If a dilute strong acid is required, use battery acid as a starting source of sulfuric acid. Note that many experiments calling for dilute hydrochloric acid work just as well with dilute sulfuric acid. Battery acid will also work for many experiments calling for "concentrated" sulfuric acid – indeed it is about 5 M – but will not work if one requires the dehydrating action of concentrated sulfuric acid. For such cases, consider other experiments. Note that battery acid is still quite dangerous – it will burn holes in clothing and blind in the eyes.

- Precautions: Always have a full bucket of water and at least half a liter of sodium bicarbonate or other weak base solution available. Use thick rubber gloves and wear goggles. If you do not have these in your lab, go buy them. Whenever handling battery acid, wear goggles. Keep other people away when pouring concentrated acids. If you are using either concentrated hydrochloric, ethanoic (acetic), or nitric acid, work outside and stand upwind – the fumes corrode the lungs. Wear cloth over your mouth and nose. If a bottle ever falls and breaks, calmly but clearly ask everyone to stop working and leave the room. Keep everyone upwind while the fumes blow away. Most of the acid will be consumed by reacting with cement. If the damage is significant, a building engineer should inspect the structure. Always pour acid into water when diluting. The heat of solvation of sulfuric acid especially is so exothermic that it can cause water to boil. If a small quantity of water is added to concentrated acid, it can boil so vigorously that it will cause acid to splash out of the container, on skin or into eyes. Finally, pour slowly from the bottle, always allowing air to enter as you pour. Otherwise, air will enter in sudden amounts, causing acid to exit the same way. This can cause it to splash back up at you.
- First Aid: For skin burns, promptly wash the affected area with a large amount of water. Then liberally apply a sodium bicarbonate or other weak base solution to the affected area. Then wash again with a large amount of water. Repeat until the burning sensation is gone. If the chemical ever gets in the eye, immediately apply sodium hydrogen carbonate solution to neutralize the acid in the eye, but nothing stronger – not carbonates and definitely not hydroxides. Then wash continuously with large amounts of water for ten minutes, longer if the eye still burns. Seek medical immediately. If swallowed, do not induce vomiting – the damage is done on the way down.
- Disposal: Add the concentrated acid to twenty or more times its volume of water and then add ash or baking soda until the mixture stops fizzing. The gas produced is carbon dioxide. Note that containers used to measure or hold concentrated acids often have enough residual acid to be dangerous. They should be submerged in a large container of water following use.

8.2.5 Diethyl Ether (ethoxyethane)

- Hazard: Can be fatal if aspirated into lungs. Also extremely flammable and a significant flash fire hazard. May also cause unconsciousness on inhalation.
- Use: Non-polar solvent
- Alternative: for a non-polar solvent, use kerosene. For a more volatile solvent, use paint thinner or lighter fluid. To demonstrate a rapidly evaporating substance, use propanone, ethyl ethanoate, or iso-propanol - note that all are also extremely flammable.
- Precautions: Never use alone (in general, do not use the lab alone). Distribute in bottles with lids or in beakers covered with e.g. cardboard to prevent evaporation. Under no circumstances should an open container of diethyl ether be in the same room as open flame. Only use in well ventilated spaces and encourage students to go outside if they feel at all drowsy or unwell.
- Disposal: See instructions on recycling of organic solvents to minimize the need for disposal. For what cannot be recovered, place where it can evaporate without being disturbed and without anyone downwind.

8.2.6 Ethandioic acid (oxalic acid), sodium/ammonium ethandioate (oxalate)

- Hazard: Poison
- Use: Volumetric analysis, oxidation-reduction reactions, qualitative analysis
- Alternatives: For its weak acid properties, use citric acid (best) or ethanoic (acetic) acid. For its reducing properties, use ascorbic acid or sodium thiosulfate.
- First aid: If ingested, induce vomiting and administer activated charcoal. Go to the hospital.
- Disposal: Collect unused solutions for another day. To dispose, add potassium permanganate solution slowly until the ethandioic acid / ethandioate lacks the power to decolorize it. At this point the compound should have been fully converted to carbon dioxide. If you used far excess oxidizing agent, neutralize with ascorbic acid prior to disposal.

8.2.7 Lead

- Hazard: Poisonous, toxic to many organs including the brain.
- Use: Unknown salt for qualitative analysis. Thus students must treat ALL unknown salts as potential lead compounds.
- Precautions: Unequivocally prohibit taste-testing of unknown salts. This seems obvious. Unfortunately, to many students it is not. Explain the hazard clearly – there are salts in the lab (e.g. barium compounds) where even a small taste can kill. Also, make sure students wash their hands.
- First Aid: If ingested, induce vomiting and administer activated charcoal.
- Disposal: Collect unused solids for another day. If the salt is soluble, dissolve all waste in a large container and add sodium chloride solution until precipitation stops. Send the precipitate to the bottom of a pit latrine. If the salt is already insoluble, drop it down.

8.2.8 Potassium hexacyanoferrate (potassium ferr[i/o]cyanide)

- Hazard: Reaction with concentrated acid releases hydrogen cyanide, the agent used in American gas chambers for executions. On inhalation, the cyanide enters the blood stream and binds cytochrome-c oxidase with a higher affinity than oxygen. Cellular respiration halts and tissues slowly die.
- Use: Qualitative analysis bench reagent.
- Precautions: Strongly prohibit mixing of different bench reagents. There are plenty of other dangerous combinations.
- First Aid: If a student seems to have trouble breathing, bring him/her outside immediately. If breathing remains difficult, seek medical attention. If the chemical is ingested, induce vomiting.
- Disposal: Dilute with plenty of water and send down the pipe. Make sure all acid waste is also diluted and neutralized.

8.2.9 Sodium hydroxide (and potassium hydroxide)

- Hazard: Concentrated solutions corrode metal, blacken wood, and burn skin. Even solutions as dilute as 0.1 M can blind if they get in the eyes. Note that this is a common concentration for volumetric analysis. Also note that the dissolution of sodium and potassium hydroxide are highly exothermic – rapid addition, especially to acidic solutions, can cause boiling and splatter. Finally, the salts are highly deliquescent and often turn to liquid if containers are not well sealed. This liquid is maximally concentrated hydroxide – the most dangerous form; do not dispose without neutralization.
- Use: Volumetric analysis, food tests, qualitative analysis bench reagent.
- Precautions: Use weak bases (carbonates and hydrogen carbonates) for volumetric analysis, provide students with goggles.
- First Aid: Treat spills and skin burns with a dilute solution of a weak base – ethanoic (acetic) or citric acid. If it gets in the eyes, immediately wash with a large amount of water. Continue washing for at least five minutes and seek medical attention if the eye still hurts.
- Disposal: Save for future use. To dispose, neutralize with citric acid or other acid waste and dump.

8.3 Chemicals that merit warning

8.3.1 Ammonium nitrate

Can explode (and shatter glassware, sending shards into eyes) if heated. Otherwise as innocuous as any other inorganic fertilizer.

8.3.2 Ethanol

The vapors are flammable, so ethanol should never be heated directly on a stove. If it must be warmed, it should be heated in a hot water bath. If the ethanol ignites anyway, do not panic. Cover the top of the ethanol container and smother the flame. Please note that methylated spirits have chemical additives that are poisonous, causing blindness, etc. Also, alcohol prepared for laboratory or industrial use is sometimes purified by extraction with benzene and probably contains traces of this potent carcinogen. Do not even think about drinking it.

8.3.3 Ethyl acetate/ethyl ethanoate

- Hazard: Extremely flammable
- Use: Solvent
- Precautions: Never open a bottle in the same room as an open flame.
- Disposal: Save for use as a solvent. If you must dispose, allow to evaporate away from people and fire.

8.3.4 Potassium permanganate

Powerful oxidizing agent. Do not mix with random substances. If you are trying something for the first time, use small quantities. Concentrated solutions and the crystals themselves will discolor skin, though the effect lasts only a few hours. This is the same stuff they sell in the pharmacies to prevent infections of cuts and surface wounds. Do not eat!

8.3.5 Propanone (acetone)

- Hazard: Extremely flammable
- Use: Solvent
- Precautions: Never open a bottle in the same room as an open flame.
- Disposal: Save for use as a solvent. If you must dispose, allow to evaporate away from people and fire.

Chapter 9

Dangerous Techniques

Some common laboratory techniques are actually quite dangerous. Identify practices in your school that seem likely to cause harm and devise safer alternatives. Below are some examples of techniques often performed in the laboratory that can easily bring harm and alternative methods to do the same thing more safely.

9.1 Mouth Pipetting

Many schools use pipettes for titrations. Many students use their mouths to fill these pipettes. We strongly discourage this practice.

The solutions used in ordinary acid-base titrations are not particularly dangerous. A little 0.1M NaOH in the mouth does not merit a trip to the hospital. Nevertheless, there are two pressing safety issues.

1. First, there are often other solutions present on the same benches – the qualitative analysis test reagents for example – that can kill if consumed. It seems like it would be a rare event for a student to mix up the bottles, but in the panic of the exam anything is possible.

2. The second safety issue applies to the best students, those that continue on to more advanced levels. High level secondary and university students must measure volumes of the size fit for pipettes for chemicals that under no circumstances should be mouth pipetted. If a student is trained in mouth pipetting, she will continue with this habit in advanced level, especially in a moment of frustration when a pipette filling bulb seems defective, or if the school has not taught her how to use them, or if they are not supplied. Students have died in many countries from mouth pipetting toxins.

Fortunately, there is no reason to ever use a pipette in secondary school, even if rubber-filling bulbs are present. *Disposable plastic syringes are in every way superior to pipettes for the needs of students.*

- They have no risk of chemical ingestion.
- They are more accurate – plastic is much easier to make standard size than glass; the pipettes available generally vary from their true volume, but all the syringes of the same model and maker are exactly the same volume.
- Plastic syringes are easier to use
- They are faster to use
- They are much more durable
- When they do break they make no dangerous shards
- They are much less expensive, by about an order of magnitude

Schools all over are already substituting plastic syringes for glass pipettes. For information on how to use these plastic syringes, please see Laboratory Techniques: Use of a Plastic Syringe to Measure Volume (p. 58).

9.2 Shaking Separatory Funnels

Separatory funnels are useful for separating immiscible liquids. They are also made of glass, very smooth, and prone to slipping out of students' hands. The liquids often used in these funnels can be quite harmful and no one wants them splashed along with glass shards on the floor.

Much better is to add the mixture to a plastic water bottle, cap it tightly, and shake. After shaking, transfer the contents of the bottle into a narrow beaker. Either layer can be efficiently removed with a plastic syringe.

There are some cases where a separatory funnel remains essential. For secondary school, however, simply design experiments that use other equipment - and less harmful chemicals.

9.3 Looking Down into Test Tubes

May blind.

Part III

Laboratory Management

Chapter 10

Classroom Management in the Laboratory

In addition to the guidelines recommended in the Laboratory Safety section, we recommend the following strategies to keep lab work safe, productive, and efficient.

10.1 Set lab rules

Before the first practical of the year, hold a short session to teach lab rules and lab first aid. Try to set a few clear, basic rules – like the four proposed in the Laboratory Safety section – instead of a long list of rules. Post these rules in the lab, and be consistent and strict in enforcing them with students and teachers.

10.2 Train students in basic techniques

For students just beginning laboratory-based education, you can probably teach each specific skill one at a time as they come up in experiments. For more advanced students, especially when they have different backgrounds in terms of laboratory experience, it is wise to spend several sessions practicing basic techniques (e.g. titrations for chemistry, using the galvanometer for physics, etc).

10.3 Have students copy the lab instructions before entering the lab

Do not let them into the lab unless they can show you their copy of the procedure, etc. Have a class dedicated to explaining the practical activity before the actual session. Bring a demo apparatus into the classroom.

10.4 Demonstrate procedures at the beginning

Do not assume that students know how to use a syringe or measure an object with calipers. If there are many new procedures, hold a special session before the practical to teach them the procedures. For titration, for example, hold a practice session in using burettes and syringes with water and food coloring. For food tests, explain and demonstrate each step to the students before holding a practical. It will save you a lot of trouble during the actual practical.

10.5 Have enough materials available

Always prepare 25-50 percent more reagent than you think you will need. Also have spare apparatus in case they fail in use. For example with physics, have extra springs, resistors, weights, etc. That said; do not make all of what you prepare immediately available to the students. As with sugar and salt, an obvious surplus increases consumption. If there is a definite scarcity of resources, it may be necessary to distribute the exact volumes necessary to each student. If you are doing this, make sure students understand that there is no more. In an exam, you might take unique objects, such as ID cards, to ensure each student receives her/his allotment only once.

10.6 Have enough bottles of reagent available

Even if only a small quantity of a reagent is needed, divide it into several bottles and put a bottle on each bench. If the volume is sufficiently small, distribute the chemical in plastic syringes. Do not use

syringes for concentrated acids or bases – because these chemicals can degrade the rubber in the syringe, there is a risk of the syringe jamming and the student squirting chemicals into eyes. The waiting caused by shared bottles leads to frustration and quarrels between groups. The last thing you want are students wandering around the lab and crowding to get chemicals.

10.7 Designate fetchers

If students must share a single material source, designate students to fetch materials If a reagent needs to be shared among many students, explain this at the beginning, and have them come to the front of the room to get it rather than carrying it to their benches. This will help to avoid arguments and confusion over where the reagent is. If the students are in groups, have each group appoint one student to be in charge of fetching that chemical. However, it is much better to have the reagent available for each group at their workplace.

10.8 Teach students to clean up before they leave

This will save you a lot of time in preparing and cleaning the lab—and it is just a good habit. Do not let students leave the lab until their glassware is clean and the bench is free of mystery salts and scraps of paper. If they do, consider not letting them in for the next practical. This might take assigned seats if you have many students. When they perform this clean up, make sure they follow whatever guidelines you have set for proper waste disposal.

10.9 Allow more time than you think you will need

What seems like a half hour experiment to you may take an hour for your students. Add fifteen minutes to a half hour more than you think will be necessary. If you finish early, you can have them clean up and then do a bonus demonstration.

10.10 Know the laboratory policies at the school

What is the policy on replacing broken equipment at the school? As a teacher, you need to know what you are going to do when the student drops an expensive piece of glassware. It is no fun to make up procedure while a student is in tears. What criteria will you use to determine if the student is "at fault?" Of course, this is less of an issue if you do not use glass apparatus.

Chapter 11

Routine Cleanup and Upkeep

Like gardens and children, laboratories require constant attention. The Second Law of Thermodynamics does not sleep. The following advice should keep you on the winning side of the struggle against entropy.

11.1 Things to do immediately

- Remove broken glass from the floor. Use tools, like pieces of cardboard, not fingers!
- Neutralize and wash up chemical spills
- Replace chemical labels that have fallen off

The person who made the mess should clean it up. Make sure they know how before they are in a position to make a mess. If they are unable (e.g. hurt), have someone else do it. Review the incident with everyone present focusing on how to prevent similar accidents in the future. Avoid blaming other people – as the supervisor the accident is your fault; either you did not train someone well enough or your supervision of their behavior/technique was inadequate.

11.2 Things to do right after every lab use

- Return stock containers of chemicals to the store area. Only teachers should move glass bottles of corrosive or toxic chemicals. Remember to carry these with two hands!
- Transfer waste, including chemicals to be reused, into suitable storage containers
- Return apparatus to their proper places
- Put broken apparatus in a special place
- Wash off all benches / tables

The people who used the lab should do these things. If it is a lab class, the students should clean up the lab in that class period. If it is a group of teachers preparing experiments, the teachers should clean up their mess. Mess tends to grow with time, and no one wants to clean up someone else's mess.

11.3 Things to do either right after lab use or later that same day

- Transfer chemicals to be reused into more permanent and well labeled storage containers.
- Process all waste for disposal. See the instructions in Waste Disposal (p. 50).
- Remove all trash from the laboratory.

If done right after lab use, those who used the lab should do this work. If the work is done later anyone can take out the trash but waste should only be processed by someone who knows what (s)he is doing, and never working alone.

11.4 Things to do every week

- Sweep and mop the floor. Note that this should be done with brooms and buckets of water, or long handled mops, not by pushing cloth on the floor directly with hands.
- Wipe down the chemical storage area. Check for broken and leaking bottles.

- Ensure that sinks (if present) are not clogged. If a sink is clogged, either unclog it immediately or prevent use of the sink by physically obstructing the basin and also writing a sign. Signs by themselves are often insufficient. Barriers with signs tend to get moved.

You can do this work or you can train students to do it. Supervise their work while they are learning to make sure they use safe techniques. Ensure that students never work alone – even for mopping at least two students must be present at all times. Students should not work in the chemical storage area without a teacher present.

Chapter 12

Waste Disposal

12.1 Introduction to waste management

Practical work produces chemical waste. Some of these wastes may be harmful to people, property or the environment if not properly treated before disposal. Regardless of where the waste will go – a sink, a flower bed, a pit latrine – the following procedures should always be followed.

Note, often there are unused reagents at the end of a practical. These are valuable and should be stored for use on another day. When storing left over reagents, label the container with:

1. The name of the compound, e.g. "sodium hydroxide solution"
2. The concentration, e.g. 0.1 M
3. The date of preparation, e.g. 15 June 2010
4. Important hazard information, e.g. "CORROSIVE, neutralize spills with weak acid."

Sometimes, there are used reagents that may be recycled. Recycling of chemicals reduces harm to the environment and saves money. Examples of chemical recycling are:

- Regenerating silver nitrate solution from qualitative analysis waste.
- Purification for reuse of organic solvents from distribution/partition law waste

In order to recycle these compounds, students must put their waste in designated containers. Specific instructions for chemical recycling follow in another section.

Some wastes may be discarded without worry. These solutions may be poured down a sink or into a pit latrine. These include:

- The final mixture in the flask after a titration. This is neutral salt water.
- All of the wastes from food tests in biology. Note that unused reagents are not waste!

Finally, some wastes require special treatment. These wastes and their treatments follow.

12.2 Special instructions for certain wastes

12.2.1 Organic wastes

These are any substance that does not mix with water, for example kerosene, isobutanol, ether, chloroform, etc. These substances should be placed in an open container and left to evaporate down-wind from people and animals. Setting these wastes on fire is usually unnecessary and may be dangerous.

12.2.2 Strong acids

Sulfuric, hydrochloric, and nitric acid solutions will corrode sinks and pipes if not neutralized before disposal. These wastes should be collected in a special bucket during a practical. After the practical, bicarbonate of soda should be added until further addition no longer causes effervescence. The gas produced is carbon dioxide.

12.2.3 Strong bases

Sodium and potassium hydroxide solutions as well as concentrated ammonia solutions are also corrosive. These wastes should be collected in a different special bucket during a practical. After the practical, the waste should be colored with POP or a local indicator and acid waste should be added until the color changes. If there is more base waste than acid waste available to neutralize it, citric acid may be added until the color finally changes.

12.2.4 Heavy metals

Barium, lead, silver and mercury solution are highly damaging to the environment and may poison human or animal drinking water if disposed without treatment. Waste containing barium and lead, generally from qualitative analysis, should be collected in a special container during a practical. After the practical, dilute sulfuric acid should be added drop-wise until further addition no longer causes precipitation. At this point, soluble lead and barium will have been converted to insoluble lead sulfate and barium sulfate. These salts may then be disposed in a pit latrine. The waste should of course first be neutralized with bicarbonate of soda.

Waste containing silver should be collected in a different special container. Ideally, this waste will be treated to regenerate silver nitrate solution according to the instructions in the next section. If such recycling is infeasible, sodium chloride solution should be added drop-wise until further addition no longer causes precipitation. At this point, soluble silver will have been converted to insoluble silver chloride and may be disposed in a pit latrine.

There is no treatment for mercury solutions that may be safely performed in a secondary school. This fact combined with the extreme danger of using mercury compound in schools supports the recommendation that mercury compounds never be used. If mercury waste is ever discovered at the school, it should be placed in a well-sealed bottle labelled: MERCURY WASTE. TOXIC. DO NOT USE. MUST NOT ENTER THE ENVIRONMENT. SUMU KALI. USITUMIE NA USIMWAGE.

12.2.5 Strong oxidizers

Concentrated solutions of potassium permanganate, chromate, dichromate, hypochlorites (bleach), and chlorates should be reduced prior to disposal. Grind ascorbic acid (vitamin C) tablets to powder and add until the permanganate decolorizes, chromate and dichromate turn green or blue, and hypochlorites lose their smell. The resulting solutions may be safely disposed in a sink or pit latrine.

12.2.6 Solid waste

Solids clog pipes and should never be put into sinks. If the solid is soluble, dissolve it in excess waste and treat as solution waste. If the solid is insoluble, dispose into a pit latrine.

12.2.7 Unknown compounds

If you do not know what a compound is, you do not know what kind of treatment it requires prior to disposal. That solution that looks like water could be nitric acid, or mercury chloride solution. Before disposing of unknown compounds, please use the Guide to Identifying Unknown Chemicals in the appendix. Even if you cannot identify the compound with these instructions, you can use them to ensure that it is not dangerous to dispose.

Chapter 13

Recycling Materials

13.1 Recycling Silver Nitrate

In many places, silver nitrate is the most expensive chemical used in a school laboratory. Silver nitrate is often used to confirm the presence of halide ions, which form insoluble precipitates with silver cations. The result of such tests are silver halide precipitates, themselves of little value.

- To regenerate the silver nitrate from these silver halides you must first reduce the silver halides to silver metal and then dissolve the metal in nitric acid. This process is easiest and most efficient with a large amount of material, so consider accumulating silver waste in a central location for many terms and perhaps from many schools.
- To reduce the silver halides, they must be in solution.
  - Add aqueous ammonia solution to the silver halides until they dissolve. You have formed a soluble silver - ammonia complex.
  - Add to the mixture a reducing agent. We have used both glucose and steel wool. Ascorbic acid, zinc metal, and sodium thiosulfate should in theory also work.
  - Heat the mixture until a metallic silver precipitate forms. It is OK if the solution boils.
- Once you believe all of the silver has precipitated as metal, decant the liquid, ideally filtering to separate all of the silver metal. Wash the silver metal in distilled (rain) water and filter again.
- Before adding nitric acid, make sure that the silver is dry.
- Then, add concentrated nitric acid slowly. The goal is to dissolve most but not all of the silver metal. If you dissolve all of the metal, you may have residual nitric acid in your silver nitrate solution that will make it ineffective for many uses.
- Decant the solution into a dark bottle - silver nitrate decomposes in light - and save the residual silver metal for the next time you do this.

13.2 Recycling Organic Waste

Organic chemicals are often expensive to purchase and difficult to dispose. Every effort should be made to collect organic wastes and recycle them. For the purpose of this discussion, organic chemicals are liquids insoluble in water, e.g. kerosene, ether, ethyl ethanoate (ethyl acetate), etc.

- Mixtures of multiple organic wastes require fractional distillation to separate. This is difficult and dangerous without the right equipment. Generally, if none of the organic chemicals in the mixture are particularly dangerous – see the section on Dangerous Chemicals (p. 37) – it is best to label the mixture "mixed organic solvents, does not contain benzene or chlorinated hydrocarbons" and keep it for future use as a generic solvent or for solubility activities.
- Mixtures of a single organic waste and water are inherently unstable, and given enough time will separate out into two layers. If the organic layer is on the bottom, it is probably di-, tri-, or tetrachloromethane, all dangerous chemicals. Follow the instructions in Dangerous Chemicals (p. 37). If the organic layer is on the top, simply decant it off. You might do this in two steps – the first to separate only water from organic mixed with some water, and the second to separate from the latter fraction pure organic from a small volume that remains a mixture. Then the water can be discarded, the organic saved, and the small residual mixture left open to the air to evaporate.
- Often, mixtures of organic and aqueous waste contain a solute dissolved in both solvents. The solvent is said to be distributed or partitioned between these two layers. Examples of compounds that partition between an organic and an aqueous layer are organic acids, like ethanoic acid (acetic acid) and succinic acid, and iodine when the aqueous layer is rich in iodide (usually potassium iodide). To reuse the organic layer it is necessary to first remove the solute.

- If the solute is an organic acid, add a small amount of indicator to the mixture and then sodium hydroxide solution, shaking vigorously from time to time. The sodium hydroxide will react with the organic acid in the aqueous layer, converting it to the salt. As the concentration of the acid in the aqueous layer decreases, the distribution equilibrium will "push" acid dissolved in the organic layer into the aqueous layer, where it too is converted to salt. Eventually, all the organic acid will be converted to its conjugate base salt, which is only soluble in the aqueous layer, and the indicator will show that the aqueous layer is alkaline even after much shaking. Now the organic layer may be run off as above.

- If the solute is iodine, the organic layer should have a color due to the iodine, and thus it will be straightforward to know when the iodine is fully removed. If there is no color, add starch to give a black color to the aqueous layer. Then add ascorbic acid (crushed vitamin C tablets) to the mixture and shake vigorously until either the organic layer returns to its normal color or the starch-blackened aqueous layer turns colorless. At this point all of the iodine will have been reduced to iodide, soluble only in the aqueous layer. The clean organic layer may be run off as above. Sodium thiosulfate may be used instead of ascorbic acid.

If you require the final organic to be of quite high purity, repeat the treatment. A small amount of residual water may also be removed with use of a drying agent, such as anhydrous sodium sulfate or calcium chloride.

13.3 Industrial Ecology in the Laboratory

Industrial Ecology is a manufacturing design philosophy where the byproducts of one industrial operation are used as input material for another. The philosophy may be applied to a school laboratory with similar economic and environmental benefits.

The science teacher generally plans the term in advance, and thus has a good understanding of the experiments students will perform. Each experiment has input reagents and output products. Normally, each of these inputs has to be purchased, sometimes at great expense, and each of these outputs has to be disposed of properly. When the term is analyzed in aggregate, however, there should be many occasions where the outputs of one experiment may serve as the inputs for another.

For example, students learning about exothermic reactions might dissolve sodium hydroxide in water and measure the temperature increase. The students might then use this solution of sodium hydroxide for a titration against a solution of ethanoic acid. The product of this titration will be perfectly balanced sodium ethanoate solution, which may be used in qualitative analysis for detecting iron (III) salts.

To maximize the opportunities for such pairings of inputs and outputs, the teacher should identify the reagents and byproducts of all activities planned for the term. Teachers may even coordinate between subjects - the reaction between citric acid and sodium carbonate to make carbon dioxide in chemistry class produces a sodium citrate solution that may be used to prepare Benedict's solution for biology class.

Part IV

Laboratory Techniques

Laboratory Techniques in Chemistry

Chapter 14

Use of the Beam Balance for Measuring Chemicals

14.1 Measuring Mass

A common tool for measuring mass is the triple beam balance. The name comes from the three parallel beams holding sliding weights. On one side of pivot point there is either a flat metal surface or a boom suspending a weighing tray. On the other side of the fulcrum, one the three parallel beams, are weights that the user slides closer or further to the pivot point. At the far end of the three beams is some kind of level indicator showing when the balance is in equilibrium or, if not, which side is too heavy.

14.2 Calibration

Calibrate the balance prior to use. Move all the sliding masses as far as they go towards the pivot point – the zero mass mark. There are usually small groves that the sliders will fit snugly in. Make sure they are in those groves – each slider except for the smallest should "click" into place. Take off any weight on the weighing tray and clean it completely. Look at the level indicator. There are two pieces. The right side not moving, but the left side of the level will move on addition of mass. The level shows the balance is calibrated when the level forms an unbroken horizontal line. If the balance is not level, there usually is a massive screw or a dial under the weigh pan. Turn it until the balance becomes level.

14.3 Weighing Chemical Samples

Triple beam balances are very accurate at measuring masses if used properly. Do not measure the chemicals directly on the metal weighing tray; use a piece of paper or glass. Many samples will react with the metal, permanently altering its mass and ruining the balance. Because the paper of glass you put the chemical on has mass, before adding any chemical you must weigh the paper or the glass first by itself. To weigh properly, move the sliders slowly until the balance becomes level or makes a horizontal line. Start with the smallest. If you reach the end before the balance equalizes, return the mass to its zero and start moving the next larger mass, one stop at a time. When the balance "tips," move back one notch and again move the smallest slider until the balance is level. Record this mass by adding each of the slides together. The mass should be recorded to one decimal place beyond the units of the smallest lines on the balance. For example, if the lines each represent 0.1g, estimate the position of the slider to the nearest 0.01 g.

Sum the desired chemical mass with mass of the paper or glass you just measured. Move the sliders to this total mass. Now, slowly add the chemical onto the paper or glass until the beam balance becomes level. After weighing, transfer the chemical from the glass or paper into whatever will actually hold it. If you use a glass and plan for the sample to be dissolved, rinse the glass into your solution container to get every last bit of chemical into your solution. If you spill any chemical on the balance, clean it up immediately.

14.4 Simplified Procedure

- Clean and calibrate balance
- Use some paper or glass and move the sliders until level
- Sum the mass of desired chemicals to the mass of the paper or watch glass.
- Add the chemical until balance is level.
- Transfer chemical to receiving container.
- Clean up any spills

14.5 Other Important Tips

Many times, you will need to measure small masses, less than 5 grams. Unfortunately, the beam balance is not as accurate when measuring such small masses, as movements in the air can cause the balance to err. To overcome this problem, place an additional mass on the weighing tray along with the paper so that the effective mass is much larger. If you are using a glass container, this step is probably unnecessary. If you add another object to the tray, make sure that there is enough space still for your chemical!

Wind is another difficulty – find a place to weigh where the air is still, perhaps in a closed room or behind some sort of obstacle or screen.

If you need many samples each the same weight, use papers of identical size and therefore mass. This allows you to keep the sliding masses in the same place for each weighing.

If you are measuring a deliquescent chemical – one that takes in water from the air, e.g. sodium hydroxide, iron (III) chloride, etc – work efficiently, but remain careful not to spill. Close the stock chemical bottle as soon as possible after use. Measure the chemical on glass rather than paper if possible as the paper often absorbs the solution that forms as the chemical deliquesces.

Finally, make sure that the volume of substance you are measuring will physically fit on your paper or glass. For volumes greater than 20g of most substances, consider using a beaker or plastic container. For volumes 100g or greater, you almost certainly need a wide mouthed and high walled vessel to hold it all. Look at the volume of substance in a contain of known mass to have an idea of how much space your sample will occupy.

Chapter 15

Use of a Plastic Syringe to Measure Volume

15.1 Safety First

Syringes are probably the best means of transferring specific volumes. They are also very safe – if used correctly. First, many syringes come with sterile needles in the same package. If this is the case, open the packages yourself and collect the needles. Never provide students with both syringes and needles. Syringe needles are designed to inject compounds into the bloodstream. Many laboratory chemicals can be very toxic if injected into the blood, and any injection done improperly carries significant risk of serious infection. Laboratory syringes should be used without the needles. If you decide to keep the syringe needles for tools (e.g. optical pins, dissection pins), store them in a well labeled container. If you decide to not keep the syringe needles, dispose of them in a sharps bin at a health center (best) or in a pit latrine.

Laboratory syringes should never be used for anything other than work in the laboratory. They should never leave the laboratory. Do not let students play with the syringes like squirt guns or point them at students' eyes even when empty. The mantra for all gun users – treat every gun like a loaded gun – should apply to syringes. They should be held with the nozzle pointed down.

Anyone working with organic solvents or concentrated acids/bases should wear goggles, whether or not syringes are involved.

15.2 Measuring Volume

There are two ways to use a syringe. The second is superior.

15.2.1 Direct Measure

Place the syringe in the solution you want to measure. Push the plunger completely in to remove all air. Draw the plunger back beyond the desired volume. Use the front of the rubber plunger to read the volume measured. Slowly push the plunger in until the rubber reaches the desired volume. Remove the syringe from the liquid being measured and transfer the liquid to the desired receptacle.

This method is a poor way to use the syringe. First of all, it is difficult to remove all the air bubbles from syringes. You will push the plunger in and out many times and still not be free of the bubbles. Often students turn the syringe upside-down and try pushing the bubbles up and out. While this effectively removes air, this method is likely to eject chemicals out into a student's eye.

In addition, using the rubber stopper to measure is surprisingly difficult. It is hard to see the volume markings, and the curvature of the rubber can cause confusion. Also, the refractive index of water is different than air, introducing additional error. Finally, if this method is used to measure organic solvents or concentrated solvents, these chemicals will react with the rubber in the syringe. This will make the rubber sticky and difficult to draw in and out. This makes the likelihood of an accident even higher. Therefore, we do not recommend this method for measuring volume with a syringe.

15.2.2 Air Bubble Method

Before putting the syringe into the solution you want to measure, draw back the plunger so it holds about 1 mL of air. Now put the syringe in the solution. Draw the plunger back beyond the desired volume. This time, there will be a large air bubble between the rubber and the top of the solution. Hold the syringe about the liquid being measured and push down the plunger until the top of the liquid inside is at the desired volume. Make sure that the top level of the liquid is level with your eye to prevent parallax error. Hold the container of liquid up so liquid exiting the syringe does not fall a long distance and splatter. Transfer the measured volume to its receptacle.

This method is the preferred manner of using a syringe. The air bubble allows for easier and more exact volume measurements. In addition, this method can be used with concentrated chemicals and

organic solvents. The air bubble does not allow these chemicals to come in contact with the rubber, at least on the initial measure. The rubber will start to react with the residue, and without prompt cleaning this can destroy the syringe.

15.3 Cleaning Syringes After Use

Like all lab equipment, syringes need to be cleaned after use. Fill a beaker or other open mouth container with water. Draw water into the syringe and push it out. Repeat 2 or 3 times. If you used the syringe to measure an organic solvent, wash the syringe thoroughly in soapy water and then rinse in ordinary water until all the soap is removed.

15.4 A Note About Auto-Disable Syringes

Almost all currently available syringes are marketed as auto-disable or safety syringes. These syringes have two mechanisms to prevent their reuse. First, if the plunger is completely depressed it will catch and the syringe will be rendered useless. Secondly, if the plunger is pulled out of the syringe tube the plunger shaft will detach from the plunger head, thus rendering the syringe useless. To allow the syringes to be reused as laboratory equipment, first take the wooden end of a matchstick and push it into the syringe tip to destroy the hooks that catch the plunger. This allows the syringe to be completely depressed and reused multiple times. To deal with the second mechanism, care should be taken not to completely remove the plunger from the syringe tube.

It is important to inform students of the proper use and handling of these newer auto-disable syringes.

Chapter 16

Measures of Concentration

16.1 Molarity (M)

Molarity is the number of moles of substance per liter of SOLUTION. Note that molarity is not the number of moles of substance per liter of solvent (e.g. water), although practically these are very similar. A molar solution has a concentration of 1 M.

16.2 Density and percent purity

These measurements are used to find the concentration of stock acid solutions. The acid bottle should list two pieces of information: the density of the acid in $\frac{g}{cm^3}$ or $\frac{kg}{dm^3}$, and the percent purity of the acid. The percent purity tells you what portion of the density is due to the acid itself, and what portion is due to water or impurities. See the chapter on Calculating the Molarity of Bottled Reagents to see how this information is used to find molarity.

16.3 Percent by mass

The percent by mass of a solute (% or $^w/_w$ or $^m/_m$) is the grams of the solute in 100 g of solution. Now, for most practicals, solutions do not need to be very precise. Thus it is acceptable to let the percent by mass just be the grams of solute in 100 ml of water. This makes these solutions much faster to prepare.

Such approximation may not suffice for more advanced work. Consider a 1% by mass solution of copper (II) sulfate. This solution should contain 1 g of $CuSO_4$ in 100 g of solution. This means that the mass of water is $100g - 1g = 99g$. By assuming that the density of water before adding the solute is $1^g/_{mL}$, we find that 99 mL of water must be combined with 1 g of $CuSO_4$ to make the solution. This difference matters if you are making, say, a solution of iron sulfate on which students will perform a redox titration.

16.4 Percent by volume (% or $^v/_v$)

Percent by volume is used to measure concentration for a mixture of a liquid chemical and water. It is equal to the volume of the chemical divided by the volume of the solution.

Example: What volume of pure ethanol must be used to make 500 mL of a 70% ethanol solution?

70% ethanol means 70 mL ethanol per 100 mL of solution. Thus, the required volume is:

$$\text{volume of pure ethanol} = \text{total solution volume} \times \text{desired fraction ethanol}$$

$$V = 500ml \times 0.70$$

$$V = 350mL$$

16.5 Normality (N)

The normality of the solution is closely related to the molarity. For many solutions, the normality IS the molarity. Normality is generally used in older books to refer to acid and base solutions. Technically, it is the "moles of equivalent" per liter. So for an acid solution, it is the moles of H^+ per liter of solution. For a base solution, it is the moles of H^+ that may be neutralized per liter of solution. For example, 1 M HCl has one mole of H^+ per liter of solution. Thus 1 M HCl is also 1 N. However, 1 M H_2SO_4 provides TWO moles of H^+ per liter of solution, so 1 M H_2SO_4 is 2 N. In a similar vein, 1 M NaOH is 1 N, but 1 M Na_2CO_3 is 2 N.

16.6 Molality

Molarity is the number of moles of solute per liter of solution. **Molality** is the number of moles of solute per kilogram of SOLVENT. In dilute aqueous solutions, the molarity and the molality are almost the same.

16.7 Some Notes on Calculations

Many textbooks and student notebooks transcribed from them feature equations that range between novel and obtuse to the American eye. Here are two very common equations that you should be aware of, mostly because the teachers that mark exams expect students to use them.

First off, the equation that often defined molarity as $M = \frac{concentration}{molecular\ mass}$. That is, molarity is equal to the concentration in grams per liter divided by the molecular mass of the solute (in grams per mole).

Second, the central equation for titration calculations:

$$\frac{(M_A)(V_A)}{(M_B)(V_B)} = \frac{n_A}{n_B}$$

A refers to the acid, B to the base, M to molarity, V to volume, and N to the stoichiometric coefficient of the acid/base in the reaction equation.

This said, there is a strong case to be made for teaching students equations that rely on an understanding of moles rather than encouraging them to memorize antiquated methods. The above equations essentially try to circumvent the need to think about moles. If you are teaching ordinary level, teach your students moles, and then show how the molarity and titrations equations come about from this unifying concept. If students can reduce every quantitative problem to moles, they will have a better understanding of the manipulations they are performing.

Chapter 17

Calculating the Molarity of Bottled Liquids

You need three pieces of information to perform this calculation:

1. The molecular mass of the acid. This is usually written on the bottle and can be easily calculated if it is not. For concentrated acids: sulfuric acid is 98 g/mol, hydrochloric acid is 36.5 g/mol, ethanoic acid is 60 g/mol, and nitric acid is 63 g/mol

2. The percent purity of the compound. This might be expressed as a percent (e.g. 31% HCl), with the symbol $^{m}/_{m}$ (e.g. $^{m}/_{m} = 68\%$), or with the word purity ("98% pure"). If you cannot find this information, see the note at the end.

3. The density (ρ) or specific gravity (s.g.) of the acid. This should be in grams per cubic centimeter (cc or cm^3).

Then, you can calculate the molarity of your concentrated acid with this formula:

$$\text{molarity} = M = \frac{(\text{percent purity})(\text{density})(1000\frac{\text{cm}^3}{\text{L}})}{\text{molecular mass}}$$

For example, the molarity of an acid bottle labeled "H2SO4, 98%, s.g. 1.84" we would calculate:

$$\text{molarity} = M = \frac{(0.98)(1.84\frac{\text{g}}{\text{cm}^3})(1000\frac{\text{cm}^3}{\text{L}})}{98\frac{\text{g}}{\text{mol}}}$$

Note that we used 0.98 for 98%. Convert all percents to decimals.

Once you do this work, take out a permanent pen and label your stock bottle with its molarity. Then no one needs to do this calculation again.

Note: since you will correct the concentration of your solutions with relative standardization, you really just need to know the approximate molarity of your liquid stock reagent. For new bottles of concentrated acid, you may assume that sulfuric acid is about 18 M, hydrochloric acid is about 12 M, and that both nitric acid and ethanoic (acetic) acid are about 16 M. Battery acid should be 4.5 M H_2SO_4.

Chapter 18

Preparation of Solutions

For many exercises, solutions do not need to be prepared accurately. Even a 50% error in the preparation will still allow an effective experiment. For other activities, the solutions should be prepared with a great deal of accuracy. This is especially true for volumetric analysis and conductivity experiments. This section deals with the preparation of solutions when accuracy counts.

18.1 Measure the Water

- Calculate the total volume of solution you need to prepare. For example, if you are doing a practical with 100 students and each requires 150 mL of solution you should make at least 15 L of solution. Making 20 L is probably wise, to have some extra.
- Find a container large enough for the total volume. Plan ahead to ensure you have a large enough container.
- Add the required volume of ordinary water.
- If your syllabus encourages you to often practice acid-base titrations, designate a pair of suitably large buckets as your permanent ACID and BASE buckets and label them as such with a permanent pen. Then, use a 1 liter container to add water to these buckets, one liter at a time. Use the permanent pen to mark the water height after each liter. Use these marks when adding water to make solutions. Round up the volume you need to the nearest liter (e.g. 71 students $\times$ 200 mL per student = 14.2 L, so make 15 L). As long as you use relative standardization when you finish preparing the solutions, any errors you make when measuring the volume will not affect your students' results.
- Distilled water is rarely necessary. If you are preparing solutions for volumetric analysis, read the section on Relative Standardization to learn how to correct small errors caused by the composition of the tap or river water. If the water forms a precipitate when making solutions of hydroxide or carbonate, allow the precipitate to settle and decant the solution. If you are making a dilute solution, you might add hydroxide or carbonate gradually with mixing until precipitation stops and then add the amount you actually need to the liquid after decantation. If the only water supply if muddy, let the dirt settle and decant or use a cloth filter. If the particles are very fine, add a chemical like potassium aluminum sulfate (alum) or iron sulfate to precipitate the dirt. If you think that you do need distilled water, rain water is almost always sufficient.

What comes next depends on the nature of your stock chemical. In general, there are two kinds of solutions:

- Solutions prepared from solid stock chemicals, e.g. sodium hydroxide, citric acid
- Solutions prepared from liquid stock chemicals, e.g. sulfuric acid

18.2 Preparing solutions from solid stock chemicals

- Calculate the amount of solid chemical required. If the instructions give the required concentration in grams per liter (e.g. 4 $^{\mathrm{g}}/_{\mathrm{L}}$ NaOH solution), multiple the total volume by the required concentration (e.g. 4 $^{\mathrm{g}}/_{\mathrm{L}} \times 10$ L = 40 g). If the instructions give the required concentration in molarity or moles per liter (e.g. 0.1 M NaOH solution), multiple the required molarity by the molecular mass of the compound to find the required concentration in grams per liter (e.g. 0.1 $^{\mathrm{mol}}/_{\mathrm{L}} \times 40\ ^{\mathrm{g}}/_{\mathrm{mol}} = 4\ ^{\mathrm{g}}/_{\mathrm{L}}$). Then, multiple the required concentration by the total volume ($4\ ^{\mathrm{g}}/_{\mathrm{L}} \times 10$ L = 40 g).
- Use a balance to weigh the solid chemical. Remember to weigh the chemical in a plastic container or on a sheet of paper and not on the scale pan directly. Some chemicals (e.g. sodium hydroxide) react

with the metal pan. If you are unfamiliar with how to use a balance, read Use of the Beam Balance for Measuring Chemicals (p. 56). If you do not have a balance, read the section on Preparation of Solutions Without a Balance (p. 65).

- Carefully add the solid chemical to the water and stir with something unreactive (e.g. glass rod, broken burette, thick copper wire) until it has completely dissolved.

18.3 Preparing solutions from liquid stock solutions

- Calculate the amount of liquid chemical required. To do this, you need to know the molarity of your stock chemical. See the section on Calculating the Molarity of Bottled Liquids (p. 62). If the instructions give the required concentration in molarity or moles per liter, use the dilution equation to calculate the amount of concentrated required:

$$(M_{concentrated})(V_{concentrated}) = (M_{dilute})(V_{dilute})$$

rearranging

$$V_{dilute} = \frac{(M_{concentrated})(V_{concentrated})}{M_{dilute}}$$

For example, if you need 10 L of 0.1 M HCl and you have 12 M stock solution, the required volume of concentrated acid is

$$V_{dilute} = \frac{(12\ \text{M})(10\ \text{L})}{0.1\ \text{M}}$$

- If the instructions give the required concentration in grams per liter, divide this concentration by the molecular mass to get molarity (e.g. $\frac{3.65\ {}^{\text{g}}/_{\text{L}}}{36.5\ {}^{\text{g}}/_{\text{mol}}} = 0.1\ {}^{\text{mol}}/_{\text{L}}$) and then use the dilution equation as above.

- Use a DRY measuring cylinder the measure the required amount of liquid chemical. Concentrated acids may be measured in standard lab grade plastic measuring cylinders – there is no need for glass. If you do not have a measuring cylinder, you can use a plastic syringe. Be sure to use the Air Cushion Method for measuring volumes with syringes (see the section on Use of a Plastic Syringe to Measure Volume, p. 58) – concentrated acids will rapidly corrode the rubber in the syringe on contact, causing the syringe to jam and become dangerous. Also, please read the description of Concentrated Acids in Dangerous Chemicals (p. 39).

- Carefully pour the liquid chemical into the container of water. Stir with something non reactive (glass rod, broken burette, thick copper wire) for about one minute.

Then, for all volumetric analysis solutions, use the instructions in the Relative Standardization (p. 66) section to perfect the mole ratio of your solutions.

Chapter 19

Preparation of Solutions Without a Balance

The procedure in the section on Relative Standardization (p. 66) allows us to do something seemingly impossible – prepare solutions for volumetric analysis that allow students to get perfect results without using either a balance or volumetric glassware in the preparation. All that you have to do is make two solutions that are close, and then use several cycles of relative standardization to prefect the molarity ratio.

To measure volume, we can use marks on plastic water bottles as described in the entry for volumetric glassware in the Sources of Laboratory Equipment (p. 208) section.

19.1 To make 0.05 M sulphuric acid (equivalent to 0.1 M HCl) for fifty students

1. Put 9.9 liters of water into a bucket.
2. Add 110 mL of battery acid. This may be accomplished easily by filling a 10 mL plastic syringe eleven times.

19.2 To make 0.033 M citric acid (equivalent to 0.1 M HCl) for fifty students

1. Put 10 liters of water into a bucket.
2. Add 64 g of citric acid. In the absence of a balance, one can often have $^1/_8$ of a kilogram (125 g) measured in the market. Dissolve this in 20 L of water to produce a 0.033 M solution.

19.3 To make 0.1 M sodium hydroxide for fifty students

1. Put 10 liters of water into a bucket.
2. Add 40 g of caustic soda. In the absence of a balance, use a plastic syringe to find the volume of a plastic spoon. Fill the spoon with caustic soda and use it to add a total of 19 cm^3 or mL caustic soda knowing the volume of each spoonful. Please read the safety note in Dangerous Chemicals (p. 37).

19.4 To make 0.1 M sodium hydrogen carbonate for fifty students

1. Put 10 liters of water into a bucket.
2. Add 84 g of bicarbonate of soda. In the absence of a balance, find the volume of a spoon as above and add 39 cm^3 or mL of bicarbonate of soda. Alternately, if 8.33 liters of solution is sufficient, measure this volume of water and then add one whole box of bicarbonate of soda. A box is 70 g.

Chapter 20

Relative Standardization

Preparing large volumes of solution is difficult with great accuracy. Relative standardization is a technique to correct the concentration of solutions so that they give the correct results for practical exercises. Note that this technique is only useful in educational situations where the purpose is to prepare a pair of solutions for titration that give an answer known by the teacher. In scientific research, the aforementioned technique – absolute standardization – is used because the concentration of one of the solutions is truly unknown.

All schools should use relative standardization to check the concentration of the solutions they prepare for the national examinations. This ensures that the tests measure the ability of the students to perform the practical, and not the quality of the school's balance, water supply, glassware, etc. While useful for all schools, relative standardization is particularly helpful for schools with few resources, as it allows the preparation of high quality solutions with extremely low cost apparatus and chemicals.

20.1 General Theory

The principle of a titration is that the chemical in the burette is added until it exactly neutralizes the chemical in the flask. If the two chemicals react 1:1, e.g.

$$\mathrm{HCl}_{(aq)} + \mathrm{NaOH}_{(aq)} \longleftarrow \mathrm{NaCl}_{(aq)} + \mathrm{H_2O}_{(l)}$$

then exactly one mole of the burette chemical is required to neutralize one mole of the chemical in the flask. If the two chemicals react 2:1, e.g.

$$2\mathrm{HCl}_{(aq)} + \mathrm{Na_2CO_3}_{(aq)} \longleftarrow 2\mathrm{NaCl}_{(aq)} + \mathrm{H_2O}_{(l)} + \mathrm{CO_2}_{(g)}$$

then exactly two moles of the burette chemical is required to neutralize one mole of the chemical in the flask. Let us think of this reaction as a mole ratio.

$$\frac{\text{moles of } A}{\text{moles of } B} = \frac{n_A}{n_B}$$

Where n_A and n_B are the stoichiometric coefficients of A and B respectively.

$$\text{moles} = \text{molarity} \times \text{volume} = M \times V \text{ (so long as V is measured in liters)}$$

By substitution,

$$\frac{(M_A)(V_A)}{(M_B)(V_B)} = \frac{n_A}{n_B}$$

A student performing a titration might rearrange this equation to get

$$M_A = \frac{(n_a)(M_B)(V_B)}{(n_B)(V_A)}$$

or

$$M_B = \frac{(n_B)(M_A)(V_A)}{(n_A)(V_B)}$$

As teachers, however, we care with something else: making sure that our students find the required volume in the burette. Solving the equation for V_A we find that

$$V_A = \frac{(n_a)(M_B)(V_B)}{(n_B)(M_A)}$$

As n_A and n_B are both set by the reaction, as long as we use the correct chemicals there is no problem here.

V_B is measured by the students – it is the volume they transfer into the flask. As long as the students know how to use plastic syringes accurately, they should get this value almost perfectly correct.

The remaining term, $\frac{M_B}{M_A}$ is for the teacher, not the student, to make correct. If we prepare the solutions poorly, our students can do everything right but still get the wrong value for V_A. It is very important that we ensure that our solutions have the correct ratio of $\frac{M_B}{M_A}$ so that the exercise properly assesses the ability of our students.

Many people look at this ratio and decide that they therefore need to prepare both solutions perfectly, so that M_B and M_A are exactly what is required. This not true. The actual values for M_B and M_A are not important; what matters is the ratio M_B to M_A!

For example, if the titration requires 0.10 M HCl and 0.10 M NaOH, our expected mole ratio is:

$$\frac{M_{HCl}}{M_{NaOH}} = \frac{0.10}{0.10} = 1$$

Preparing 0.11 M HCl and 0.09 M NaOH will cause the students to get the wrong answer:

$$\frac{M_{HCl}}{M_{NaOH}} = \frac{0.11}{0.09} = 1.22$$

However, preparing exactly 0.05 M HCl and 0.05 M NaOH results in the same molar ratio:

$$\frac{M_{HCl}}{M_{NaOH}} = \frac{0.05}{0.05} = 1$$

Thus the students can get exactly the right answer if they use the right technique even though neither solution was actually the correct concentration.

How can we ensure that we have the correct molar ratio between our solutions? Titrate your solutions against each other. If the volume is not the expected value, one of your solutions is too concentrated relative to the other. You can calculate exactly how much too concentrated and add the exact amount of water necessary to perfect the ratio. This process is called relative standardization, because you are standardizing one solution relative to the other.

20.2 Procedure for Relative Standardization

In some titrations the acid is in the burette and in some it is the base is in the burette. So let us not use "acid" and "base" to refer to the solutions, but rather "solution 1" and "solution 2" where solution 1 is the solution measured in the burette and solution 2 is measured by pipette (syringe).

You should have prepared a bucket or so of each. The volume you have prepared is V_1 liters of solution 1 and V_2 liters of solution 2.

Titrate the solutions against each other. Call the volume you measure in the burette "actual titration volume" You know the desired molarity of each solution, so from the above student equations you can calculate the burette volume you expect, which you might call "theoretical titration volume."

After the titration, there are three possibilities. If the actual titration volume equals the theoretical titration volume, your solutions are perfect. Well done.

If the actual titration volume is smaller than the theoretical titration volume, solution 1 is too concentrated and must be diluted. Use the ratio:

$$\frac{V_1 \text{ (before dilution)}}{V_1 \text{ (after dilution)}} = \frac{\text{actual titration volume}}{\text{theoretical titration volume}}$$

If the actual titration volume is larger than the theoretical titration volume, solution 2 is too concentrated and must be diluted. Use the ratio:

$$\frac{V_2 \text{ (before dilution)}}{V_2 \text{ (after dilution)}} = \frac{\text{theoretical titration volume}}{\text{actual titration volume}}$$

After diluting one of your solutions, repeat the process. After a few cycles, the solutions should be perfect. Remember that the volume "before dilution" is the volume actually in the bucket, so the amount you made less the amount used for these test titrations.

Laboratory Techniques in Biology

Chapter 21

Preservation of Specimens

See the *Shika Express Biology* manual for information regarding collection and preservation of particular specimens covered in the O-level syllabus.

21.1 Dead Specimens

- Mosses and lichens: Wrap in paper or keep in a closed container.
- Plants and parts thereof: hang in the sun until dry. Alternatively, press the plants using absorbent material and a stack of books.
- Insects: Leave exposed to air but out of reach by other insects until bacteria eat everything except the exoskeleton. If you want to preserve the soft tissue, store under methylated spirits.
- Fish, worms, amphibians, and reptiles: Store in methylated spirits (will makes specimens brittle) or a 10% formaldehyde solution (more poisonous and more expensive).
- Parts of mammals (e.g. pig eyes, bovine reproductive organs): store in 10% formaldehyde solution.

21.2 Skeletons

Skin the animal and remove as much meat as possible. Bury the bones for several months. Exhume and assemble with wire and superglue.

21.3 Living Specimens

Be creative! Figure out what the animal will eat, who will feed it, what it will drink, where it can hide, how it can be observed, etc.

Chapter 22

Dissection

See the *Shika Express Biology* manual for information regarding dissection of particular specimens covered in the O-level syllabus.

22.1 Preparation of Specimens

Unless you want students to observe a beating heart, dead specimens are much easier to work with than unconscious ones. This also removes the problem of stunned animals waking up in the middle of their dissection.

- Flowers and other plant parts: No preparation required as long as the samples are relatively fresh. Store samples in closed plastic bags to minimize drying. If you intend to keep them for more than a day or two, submerge the bags in cold water to slow the rate of molding.

- Insects: Kill with household aerosol insecticide. Use specimens within one day of collection, unless you have refrigeration or freezer.

- Fish: Keep living until the day of the dissection. Then remove from water until they suffocate. Use immediately after death.

- Frogs: Able to breathe above and below water, frogs are hard to starve of oxygen. One option is to seal them in a container of methylated spirits and then rinse the dead bodies with water prior to dissection.

- Reptiles, birds, and mammals: For most organ systems, you can kill the animal by blunt trauma without ruining the lesson. Students can even bring animals caught and killed in homes. Snakes should be decapitated along with enough of the body to remove the fangs and venom sacks. Bury these deeply. Do not use animals killed by poison, or those that were found dead. For completely undamaged specimens, enclose the live animal in a cage (or a tin with adequate holes) and submerge in a bucket of water until drowned.

- Living specimens: If you really want to see that heart beating, use chloroform. This can be transferred from bottle to specimen jar via cotton ball, or perhaps made in situ by the reaction between propanone (acteone) and bleach. We have not yet attempted the latter – if you do, remember that the products are poisonous gases; indeed, that is the point. Note that if you use too little chloroform, the animal will feel the blade opening it up. If you use way too little, it may start squirming. If you use too much chloroform, however, you will simply kill the animal – you might as well have drowned it.

22.2 Tools

For more, see the section on Sources of Laboratory Equipment (p. 208).

- Scalpels can be made using razor blades and tongue depressors. Make sure the razors are very sharp. If the blade is dull or floppy, the students will probably push too hard, and may cut themselves when the skin finally gives and the blade slips.

- Optical Pins from new disposable syringes are an easy option.

- Dissection trays can be prepared using cardboard or by making a 1 cm thick layer of wax on the bottom of a shallow tray or bowl. This surface will readily accept pins and is easy to clean.

22.3 Procedure

This varies by species. The internet has many resources and there are many good books with very detailed instructions – alas, this manual is not yet one of them. A crude method follows:

1. Position the specimen on its back and make a clean, symmetric, and shallow incision down the full length of the underside.
2. Make additional perpendicular cuts at the top and bottom of the torso for an overall "I" shape. These cuts should only just penetrate the body cavity.
3. Open up skin "door" you have created, pinning them back onto the dissection tray.
4. Pick an organ system – circulation, digestion, nervous, etc – and, with the aid perhaps of a good drawing, remove other material to focus on the target anatomy.

You can teach many systems from one specimen – start with the most ventral (front) and move to the most dorsal (back).

Encourage students to sketch at various steps in the process. Also encourage them to identify anatomy for themselves, perhaps with the aid of thought provoking questions and discussion in groups.

22.4 Cleanup and Carcass Disposal

Wash all blades, pins, and trays with soapy water. Rinse all tools to remove the soap and then soak for about fifteen minutes in bleach water. When finished, rinse again in ordinary water.

Bury all carcasses in a deep pit, below the reach of dogs. You may also add kerosene and burn, but this smells bad and costs money.

Chapter 23

Preparation of Culture Media

23.1 Introduction

In microbiology, there are two basic types of media: solid agar media and a liquid broth media. From these, many types of media can be made. Generally, exact amounts of ingredients are not needed so if you want to make some agar plates or liquid cultures try with the resources you have. The recipes listed are a guideline to help you get started.

23.2 Media Recipes

23.2.1 Basic Agar (1.5%)

- 15 g/L agar
  it is like gelatin or if you can find seaweed you can grind it up

- 10 g/L nutrient source
  e.g.sugar, starch (potatoes), beans fruits like mango and papaya

- 1-2 g/L salts and phosphates
  this varies with what you want to grow — experiment! (table salt is usually fine)

- 1 L water

Add and mix all the ingredients together and heat until boiling. Boil for $\tilde{1}5$ minutes and make sure all the gelatin/agar is dissolved. Pour liquid into Petri plates (15-20 mL each). The plates should solidify $\tilde{4}5$°C. Cover and keep agar side up in a cool place if possible. If the plates do not solidify, try adding more gelatin or corn starch to thicken it up. You can also pour agar into test tubes/syringes to do oxygen tests (aerobic vs. anaerobic)

23.2.2 Blood Agar

- 15 g/L agar/gelatin/ground sea weed

- 10 g/L nutrient source

- 15 mL sheep's blood (other organizisms also work)

- 1 L water

Heat and boil agar, nutrient source and water for 15 minutes. After liquid has ceded ($\tilde{4}5$°C (when you can leave your hand on the flask for a few seconds) add in blood until the mixture is blood red. Swirl in and pour into plates.

23.2.3 Liquid Broths

- 10 g nutrient source

- 1 L water

- 1-2 g salts/phosphates

  Mix together, heat, and boil. Distribute in test tubes.

23.3 Things you can do after media preparation

- Agar-streaked plates! Swab something (back of throat, nose, belly button, door handle, etc) and gently rub onto the agar. Try not to gouge the agar.
- You can also do experiments to test the effects of salt concentrations, temperature, and nutrient concentrations.
- After all the plates solidify, incubate them at around 25-30°C. Ideally the temperature remains constant. Check the plates after 24 hours for growth.
- For liquid broths you can inoculate test tubes with a sample from the environment. Incubate and check like agar plates. If there is growth the liquid will be turbid instead of clear like a control tube with only broth.
- You can use liquid cultures for wet mounts under microscopes as samples for agar plates or to allow students to see the difference between growth and no growth.

23.4 What to use if you do not have plates or test tubes

- Use old water bottles or old plastic packaging for plates
- Use anything rigid and heavy for covered, e.g. building tiles
- Sealed/closed plastic syringes for test tubes
- Try to keep materials as sterile as possible but do not worry if there is contamination. Use contamination as a learning experience. Penicillin was contamination and it became a wonder drug.

23.5 Things to do once you have cultures

- Take a sample from agar plate and drop hydrogen peroxide on it. Does it bubble? (Yes, it has catalase)
- Extract DNA from *E. coli.*
- Fermentation = use a liquid broth with peptone, acid-base indicator like phenol red, and inverted tube to trap gas and 0.5 – 1.0% of carbohydrate you want to test. If fermentation occurs (phenol red), the broth will turn yellow and gas should be collected in the tube. If the tube remains red, you can test for glucose production by adding a few drops of methyl orange. If the pH is below 4.4, it will remain red. If the pH is above 6.0, it will turn yellow.

23.6 Guide to Identifying Common Microorganisms

- *Pseudomonas aeroginosa*: is green and smells like grape jelly (can grow in disinfectant)
- *Serratia marcescens*: grows pink-red between 25-32°C (will be white otherwise)
- *Escherichia coli*: pale white/yellow, smells like inole
- *Proteus spp*: swarm on plates and smell like urine and brownies
- *Bacillus subtillis*: pale beige, smells a bit sweet
- *Vibrio cholera*: smells like buttery popcorn
- Staph vs Strep: Staph is catalase (+), strep is (-)

Chapter 24

Using a Microscope

24.1 Parts of a Microscope

- Eyepiece: or ocular lens is what you look through at the top of the microscope. Typically, the eyepiece has a magnification of 10x.
- Body Tube: tube that connects the eyepiece to the objectives
- Objective Lenses: primary lenses on the microscope (low, medium, high, oil immersion) which are used to greater magnify the object being observed. A low power lens for scanning the sample, a medium power lens for normal observation and a high power lens for detailed observation. Normal groups of lens magnifications may be [4×, 10×, 20×] for low magnification work and [10×, 40×, 100×] for high magnification work. Some microscopes also use oil immersion lenses and these must be used with immersion oil between the lens and the cover slip on the slide. Oil immersion allows for a much greater magnification than air and typically ranges from 40x-100x.
- Revolving Nosepiece: houses the objectives and can be rotated to select the desired magnification.
- Coarse Adjustment Knob: a large knob used for focusing the specimen
- Fine Adjustment Knob: small knob used to fine-tune the focus of the specimen after using the coarse adjustment knob.
- Stage: where the specimen to be viewed is placed
- Stage Clips: used hold the slide in place
- Aperture: hole in the stage that allows light through to reach the specimen
- Diaphragm: controls the amount of light reaching the specimen
- Light Source: is either a mirror used to reflect light onto the specimen or a controllable light source such as a halogen lamp

24.2 How to Use a Microscope

- Always carry a microscope with two hands! One on the arm and one on the base!
- Plug the microscope into an electrical source and turn on
- Make sure the stage is lowered and the lowest power objective lens is in place
- Place the slide under the stage clips with specimen above the aperture
- Look through the eyepiece and use the coarse adjustment knob to bring the specimen into focus
- If the microscope uses a mirror as the light source, adjust the mirror so enough light is reflected through the aperture onto the specimen
- You can adjust the amount of light reaching the specimen by opening and closing the diaphragm
- Once the object is visible, use the fine adjustment knob for a more precise focus
- At this point you can increase the magnification by switching to a higher power objective lens
- Once you switch from the low power objective lens, you should no longer be using the coarse adjustment knob for focusing because it is possible to break the slide and scratch the lenses

- If you switch objectives, use the fine adjustment to fine-tune the focus of the object If the high powered objective lenses on the microscope say oil then you can place a small drop of immersion oil on the cover slip then switch to the oil immersion lens. *Only use the oil immersion lens with immersion oil and don't use oil with any other objective that does not say oil.*
- Once you have finished observing the specimen, lower the stage, remove the slide, and return to the lowest objective
- Clean the lenses with lens cleaner and lens paper (only use lens paper as other tissues will scratch the lenses)
- Wrap the cord around the base and cover the microscope for storage

24.3 Making a Wet Mount

- Collect a thin slice (one cell layer thick is optimal) of specimen and place on the slide
- Place a drop of water directly over the specimen
- Place a cover slip at a 45 degree angle over the specimen with one edge touching the drop of water then drop the cover slip over the specimen. If done correctly, the cover slip will completely cover the specimen and there will be no air bubbles present.

24.4 Staining a Slide

- Once you have completed the above process place one small drop of stain (ex. Iodine, methylene blue) on the outside edge of the cover slip
- Place the flat edge of a paper towel on the other side of the cover slip. The paper towel will draw the water out from under the cover slip and pull in the stain

24.5 Magnification

The actual power of magnification is a product of the ocular lens (usually 10x) times the objective lens.

Ocular lens (eyepiece)	Objective Lens	Total magnification
10x	4x	40x
10x	40x	400x
10x	100x	1000X

24.6 Troubleshooting

1. The Image is too dark!
   Adjust the diaphragm and make sure your light is on.
2. There's a spot in my viewing field, even when I move the slide the spot stays in the same place!
   Your lens is dirty. Use lens paper, and only lens paper to carefully clean the objective and ocular lens. The ocular lens can be removed to clean the inside.
3. I can't see anything under high power!
   Remember the steps, if you can't focus under scanning and then low power, you won't be able to focus anything under high power.
4. Only half of my viewing field is lit, it looks like there's a half-moon in there!
   You probably don't have your objective fully clicked into place.

Chapter 25

Low Tech Microscopy

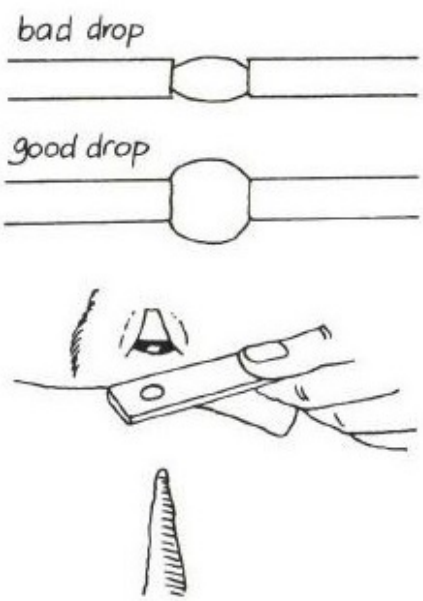
Microscopes are powerful tools for teaching biology, and many of their benefits are hard to replace with local fabrications. However, simple materials can be used to achieve sufficient magnification to greatly expands students' understanding of the very small. They may view up close the anatomy of insects and even see cells.

25.1 Water as a lens

Water refracts light much the way glass does; a water drop with perfect curvature can make a powerful lens. A simple magnifier can be made by twisting a piece of wire around a nail and dipping the loop briefly into some water. Students can observe the optical properties of the trapped drop of water.

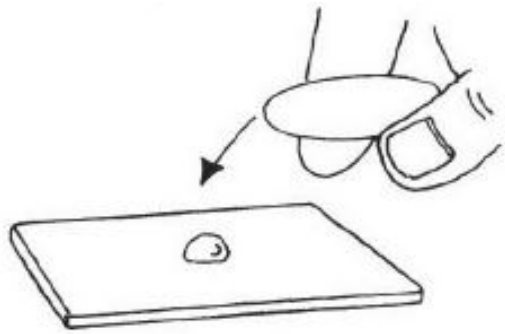
25.2 Perfect circles

Better imaging can be had if the drop is more perfect in shape – the asymmetry of the wire twisting distorts the image. Search for a piece of thin but stiff plastic – water bottles work well. Cut a small piece of this plastic, perhaps 1×2 centimeters. Near one end, make a hole, the more perfect the better. The best hole-cutting tool is a paper hole punch, available in many schools. With care, fine scissors or a pen knife will suffice; remove all burrs.



25.3 Slides

A slide and even cover slip may be made from the same plastic water bottles, although being hydrophobic they will not have the same properties of glass when making wet mounts. Improvise a method for securing the punctured plastic over the slide; ideally the vertical spacing can be closely adjusted to focus.



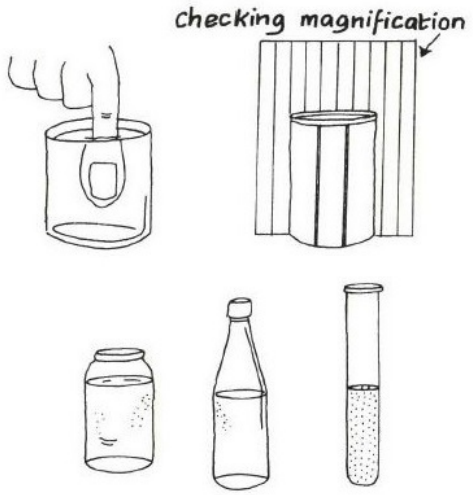
25.4 Backlighting

On a bright day, there may not be any need for additional lighting, but in most classrooms the image will be too dim to be easily seen. The sun is a powerful light source, though not always convenient. Flashlights are generally inexpensive and available; many cell phones have one built in the end. To angle the light into the slide, find either a piece of mirror glass, wrinkle-free aluminum foil, the metalized side of a biscuit wrapper, etc.

Experiment with a variety of designs to see what works best given the materials available to your school. If you use a slide of onion cells stained with iodine solution , your students should be able to see cell walls and nuclei.

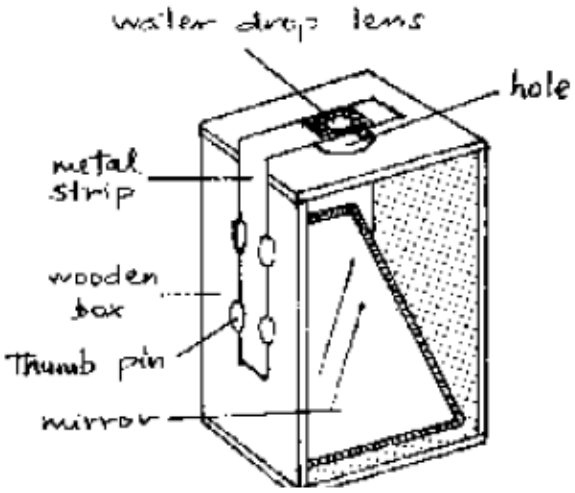
25.5 Simple Microscopes and Magnifiers

25.5.1 Clear-Container Magnifiers



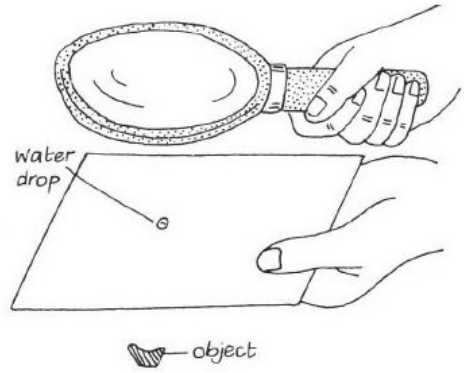
Any of these containers filled with water will make good magnifiers.

25.5.2 Simple Microscope



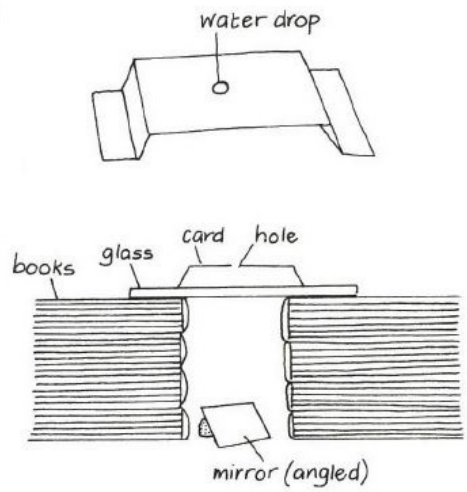
Construct a small wooden box from plywood as shown (or use a small cardboard carton such as a light bulb box). Make a round hole of 2 cm diameter, at the top. Fit a small mirror (glass or polished metal) in the box, angled to reflect light up through the hole. Make a small hole (about 6 mm) in a strip of metal. Remove the round top from a pen-torch bulb and secure it in the strip using adhesive tape. Carefully cut off the tape where it may cover the lens. Bend the strip, then fix it to the side of the box, so that it can be moved up and down. Drawing pins or nails could be used for this. The object is focused by moving this strip. Note the eye should be placed as near as possible to the lens when viewing.

25.5.3 Simple Compound Microscope



- Using 2 lenses together allows much greater magnification.
- Use a hand lens to make a water drop into a more powerful magnifier.
- Try using a hand lens with a lens from a torch bulb to make another simple compound microscope.

25.5.4 Card Bridge Microscope



- Place a water drop in the card 'bridge'.
- Place this on a sheet of glass as shown.
- Place the object you are looking at on the glass. This arrangement is most suitable for thin items, e.g. sections of leaves.
- Experiment with the angle of the mirror so that light shines up through the specimen.
- Use this arrangement with a hand lens to produce a compound microscope.

Part V

NECTA Practicals

Chapter 26

Biology Practicals

26.1 Introduction to Biology Practicals

26.1.1 Format

The format of the Biology practical exam was revised in 2011 to keep up with the 2007 updated syllabus. As such, there will be no further Alternative to Practical exams, pending approval from the Ministry of Education. Prior to 2011, there were 3 questions in the Biology practical. Question 1 was required and the student could choose to answer either Question 2 or 3.

As of now, the Biology practical has 2 questions and students must answer **both**. Question 1 can come from any of the following topics: Nutrition, Movement, Transport of Living Things, Respiration, Reproduction, Coordination, Regulation or Growth. Question 2 is on Classification of Living Things. Each question is worth 25 marks, and students have $2\frac{1}{2}$ hours to complete the exam.

Biology 1 Theory Format

The theory portion of the Biology exam comprises 100 marks, while the practical carries 50 marks. A student's final grade for Biology is thus found by taking her total marks from both exams out of 150.

The theory exam for Biology contains 13 questions over 3 sections. Section A has 2 questions worth 10 marks each. Section B has 8 short answer questions, each having two items. This section weighs a total of 60 marks, and the mark allocation for individual questions is indicated at the end of each question. Section C has 3 long answer/essay questions, though students only need to answer 1 of them. The answer must be comprehensive and include as many points as possible. It is worth 20 marks.

Note This information is current as of the time of publication of this manual. Updated information may be obtained by contacting the Ministry of Education.

26.1.2 Notes for Teachers

NECTA Advance Instructions

There are two sets of advance instructions. One set of advance instruction are given to teachers at least one month before the date of the exam. These instructions contain the list of specimen, apparatus, and other materials required for setting up the Biology practical questions. The instructions also suggest how many specimen to acquire for each candidate or group of candidates. It is imperative that the collection and storage of specimen for the practical be kept confidential.

The second set of instructions should be given 24 hours before the time of the practical. It includes how to label each specimen and which materials should be given to each candidate (or shared among candidates).

Usually if the instructions include a scalpel of some sort, this means students will be required to do some form of dissection. In most cases, the dissection is of a maize seed or bean to show whether is it a monocotyledon or dicotyledon. If the advance instructions include any form of glass apparatus or test tubes, there will most likely be a question on food tests.

26.1.3 Common Practicals

Food Tests test a food solution for starch, sugars, fats, and protein

Classification name and classify specimens, then answer questions about their characteristics

Respiration use lime water to test air from the lungs for carbon dioxide

Transport investigate osmosis by placing leaf petioles or pieces of raw potato in solutions of different solute concentrations

Photosynthesis test a variegated leaf for starch to prove that chlorophyll is necessary for photosynthesis

Coordination students look at themselves in the mirror and answer questions about the sense organs they see

Movement name bones and answer questions about their structure and position in the body

Note These are the most common practicals, but they are not necessarily the only practicals that can occur on the national exam. Food tests and Classification are by far the most common, but there are many eligible topics. Be sure to regularly look through Biology Past Papers (p. 162) to get an idea of the kind of questions that can occur.

26.2 Food Tests

In this practical, students test a solution of unknown food substances for starch, protein, reducing sugars, non-reducing sugars, and lipids. They record their procedure, observation, and conclusions, then answer questions about nutrition and the digestive system.

This section contains the following:

- Preparation of Chemical Solutions
- Preparation of Food Solutions
- Performing the Food Tests
- Examination Room
- Student Report
- Sample Food Test Practical

26.2.1 Preparation of Chemical Solutions

Always make sure the chemicals work before performing the food tests with students.

Benedict's Solution

This solution can be bought at a chemical store already prepared or you can make it yourself.

Using Sodium Carbonate:

- Add about 1 L of water to a plastic bottle.
- Add 5 spoons of sodium carbonate ($NaCO_3$).
- Add 3 spoons of citric acid.
- Add one spoon of copper sulphate.

Using Bicarbonate of Soda:

- Add 1 L of water to a cooking pot.
- Add a box (70 g) of bicarbonate of soda.
- Boil the mixture for 5-10 minutes. This makes sodium carbonate.
- Let cool and transfer to a plastic water bottle.
- Add 3 spoons of citric acid.
- Add one spoon of copper sulphate. Cap and shake to mix.

Label as: BENEDICT'S SOLUTION FOR FOOD TESTS
The solution may be stored in any plastic or glass bottle and will keep indefinitely.

Copper (II) Sulphate

- Add one spoon of copper (II) sulphate to a 1.5 L bottle.
- Add 1 L of water and shake until chemicals are fully dissolved.

Label as: 1% COPPER (II) SULPHATE SOLUTION FOR FOOD TESTS
The solution may be stored in any plastic or glass bottle and will keep indefinitely.

Iodine Solution

Make sure to use iodine tincture from a pharmacy. The tincture must not contain ethanol/alcohol/spirit.

- Add 1 part iodine tincture to 10 parts water. Example: In a 500 mL bottle, add 40 mL iodine tincture, then and 400 mL of water.
- Cap the bottle and shake.

Label as: IODINE SOLUTION FOR FOOD TESTS
The solution may be stored in any plastic or glass bottle and will keep indefinitely.

Dilute NaOH

- Using a PLASTIC teaspoon, add one level teaspoon of NaOH to a 500 mL water bottle. Caustic soda (NaOH) reacts with metal. DO NOT TOUCH.
  SAFETY NOTE: Prepare about 100 mL of citric acid or ethanoic acid solution to neutralize sodium hydroxide spills on skin or lab tables. One spoon of citric acid in 100 mL of water is suitable. Ethanoic acid solutions are sold in stores as vinegar.
- Add 250 mL of water.
  SAFETY NOTE: This reaction can cause the solution to become very warm. Avoid chemical burns by wearing gloves.
- Cap well and shake. This makes 1 M sodium hydroxide solution.

Label as: 1 M SODIUM HYDROXIDE SOLUTION FOR FOOD TESTS (CORROSIVE)
The solution will react with carbon dioxide in the air if not well sealed. Do not store in glass bottles with glass stoppers as these will stick. The solution may be stored in plastic bottles indefinitely.

Dilute Acid

Your school may have dilute hydrochloric acid or you may have to make it yourself.

Using Hydrochloric Acid (HCl):

- Add 1 part HCl to 9 parts water. Example: In a 1.5 L water bottle, add 900 mL of water, then add 100 mL of HCl.
- Shake well.

Using Citric Acid:

- Add 500 mL of water to a 1 or 1.5 L water bottle.
- Add 5 spoons of citric acid.
- Cap well and shake. This makes 0.5 M citric acid.

Label as: 0.5 M CITRIC ACID FOR FOOD TESTS
The solution may be stored in any plastic or glass bottle and will keep indefinitely.

Sudan III Solution

Using Sudan III solution takes a long time to show results. It may be replaced by iodine tincture solution for the lipids test.

- Combine 0.5 g of Sudan III powder with 100 mL of 70% ethanol solution (30 mL water and 70 mL ethanol).
- Place the solution in a warm water bath to help the Sudan III dissolve.
- Filter to remove any remaining solid.

Label as: SUDAN III SOLUTION FOR FOOD TESTS
The solution may be stored in any plastic or glass bottle and will keep indefinitely.

26.2.2 Preparation of Food Solutions

For the NECTA and mock exams you may have to set up the food test solutions. The instructions will tell you which ones you'll need to prepare in order to make 'Solution X.' Solution X consists of a mixture of at least 3 of the different food substances and is given to each student in at least 3 test tubes. Make sure food solutions are well-dissolved and colorless so that students don't know what is in the mixture. You don't need to measure the ingredients, but make sure to test the solutions before the practical.

Reducing sugar

Use glucose powder and dissolve in water. Make sure the substance is fully dissolved so that students don't know what is in the mixture.

Non-reducing sugar

Use sugar and dissolve in water. Make sure the substance is fully dissolved so that students don't know what is in the mixture.

Lipids

Mix sunflower oil with water. Shake immediately before use. Sunflower oil is best since it is liquid at room temperature.

Protein

Mix an egg white with water.

Starch

Save the water you use to boil potatoes, rice, or pasta. Make sure to remove the bits of food. You can also just mix flour in water, but it would be obvious.

26.2.3 Performing the Food Tests

Reducing Sugars Test

- Add a small amount of Benedict's solution to the food solution.
- Boil the solution and allow it to cool. Observe the colour changes from blue to green, yellow, then deep orange/brick red precipitate if reducing sugars are present.

Always do the reducing sugars test first because a non-reducing sugar will always test positive for a reducing sugar.

Non-reducing Sugars Test

- Add a small amount of dilute acid (HCl) to the solution.
- Boil the solution for about 30 seconds and allow it to cool.
- Add a small amount of NaOH to the solution and shake.
- Add a small amount of Benedict's solution and boil.
- Allow the solution to cool and observe as the solution changes from green to yellow, then to deep orange/brick red precipitate if non-reducing sugars are present.

Lipids Test

- Add a small amount of Sudan III or iodine solution to the food solution and shake.
- A red ring will form at the top of the test tube if lipids are present.

Using Sudan III colours the whole solution red whether it contains lipids or not. Use iodine solution to get a more distinct result.

Protein Test

- Add an *equal* amount of sodium hydroxide (NaOH) to the solution and shake.
- Add a small amount of copper (II) sulphate to the solution and shake.
- Observe the solution turn violet/purple in colour if protein is present.

Starch Test

- Add a small amount of iodine solution to the food solution.
- Observe the solution turn blue-black in colour if starch is present.

26.2.4 Examination Room

The NECTA practical exam is done in the school's lab or any other suitable room. Heat sources (jiko, etc.) should be spread evenly in the exam room so that students don't have to go far to heat their test tubes; this also cuts down on cheating. Spread students out and distribute supplies as you see fit.

Each student gets:

- 3 or more test tubes (to carry out 5 tests)
- A beaker containing Solution X
- A test tube rack (or a cut out water bottle with sand to hold the tubes)*

* Students may share the racks, but shouldn't share the cut out bottles

Each station should have:

- Copper II sulphate
- Water
- Dilute acid (HCl, etc.)
- Dilute base (sodium hydroxide)
- Iodine solution
- Sudan III solution (can be replaced by iodine solution)
- Benedict's solution

26.2.5 Student Report

Food test data is recorded in a table containing four columns: Test for, Procedure, Observation and Inference.

Students should write the Procedure using the passive voice in the past tense. For example, "A small amount of Benedict's solution was added to the solution. Then the solution was boiled and allowed to cool."

In the Observation column, the student should write what they observed using the past tense and passive voice. For example, "A violet colour was observed."

In the Inferences column, the students should write what they saw in the past tense and passive voice. For example, "Reducing sugars were not (or were) present."

Note that every column is worth marks on the exam. Even if students fail to do the food tests correctly, they can still get marks for writing what they are testing for and what the procedure should be.

An example of a completed food test results table is given below. Assume the solution contains proteins, reducing sugars, non-reducing sugars and starch.

Food Tested	Procedure	Observation	Inference
Lipids	A few drops of Sudan III solution (or iodine solution) were added to solution X. The solution was shaken and allowed to stand.	A red ring did not form at the surface.	Lipids were not present.
Proteins	An equal amount of NaOH was added to solution X and shaken. A few drops of copper (II) sulphate were added to solution X and shaken again.	A violet colour was observed.	Proteins were present.
Reducing sugars	A small amount of Benedict's solution was added to solution X. The solution was heated and allowed to cool.	A brick red precipitate was observed.	Reducing sugars were present.
Non-reducing sugars	A small amount of dilute acid was added to solution X. The solution was heated and allowed to cool. Then a small amount of NaOH solution was added, and the solution was shaken. Finally, a small amount of Benedict's solution was added. The solution was boiled and let cool.	The solution changed from green to yellow, then to a deep orange/brick red precipitate.	Non-reducing sugars were present.
Starch	A few drops of iodine solution were added to solution X and shaken.	A blue-black colour was observed.	Starch was present.

26.2.6 Sample Food Test Practical

You have been provided with solution **B**.

(a) Identify the food substances present in solution **B** by using the reagents provided. Tabulate your work as shown in the following Table:

Food Tested	Procedure	Observation	Inference

(b) For each food substance identified in 1(a);

(i) Name two common sources.

(ii) State their role in the body of human being.

(c) The digestion of one of the identified food substance in 1(a) starts in the mouth.

(i) Name this food substance.

(ii) Identify the enzyme responsible for its digestion in the mouth.

(d) The digestive system of human being has several parts.

(i) Name the part of digestive system in which most of digestion and absorption of food takes place.

(ii) Explain how the named part in (d) (i) is adapted for absorption of digested food substances.

Additional Food Test Questions: See Biology Past Papers (p. 162) for additional food test questions.

26.3 Classification

The classification practical requires students to identify specimens of animals, plants, and fungi. The students must write the common name, kingdom, phylum, and sometimes class of each specimen. They also answer questions about the characteristics and uses of the specimens.

This section contains the following:

- Common specimens
- Where to find specimens
- Storage of specimens
- Sample practical with solutions
- Additional classification questions

26.3.1 Common Specimens

Fungi Mushroom, yeast, bread mold

Plants Fern, moss, bean plant, bean seed, maize plant, maize seed, pine tree, cactus, sugar cane, Irish potato1, cypress tree, acacia tree, hibiscus leaf, cassava

Animals Millipede, centipede, grasshopper, lizard, tilapia (fish)3, scorpion, frog, tapeworm, liver fluke, cockroach, spider

26.3.2 Where to Find Specimens

- Start collecting specimens several months before the NECTA exams, as some specimens can be hard to find in the dry season.
- Ask your students to bring specimens! Students are especially good at finding insects and other animals. You can even find primary school children to gather insects such as grasshopper and millipedes.
- Ferns, hibiscus, pines, and cypresses are used in landscaping. Try looking near nice hotelis or guestis. Ferns should have sori (sporangia) on the underside of their leaves.
- Moss often grows near water tanks and in shady corners of courtyards. It is hard to find in the dry season.
- Sugarcane, Irish potato, cassava, tilapia, bean seeds, and maize seeds can be found at the market. Yeast is available at shops.
- Mushrooms are hard to find in the dry season. However, they are available at grocery stores in large cities, and you may be able to find dried mushrooms at the market. You can also collect mushrooms in the rainy season and dry them yourself.
- Tapeworms and liver flukes may be acquired from butchers. Find out where livestock is slaughtered and ask the butchers to look for worms (minyoo). Liver flukes are found in the bile ducts inside the liver, while tapeworms are found in the intestines. You can also try going to a livestock fair/market (mnada) or talking to the local meat inspector (mkaguzi wa nyama).
- Grow your own bread mold. Just put some bread in a plastic bag and leave it in a warm place. But do it ahead of time – it can take two weeks to obtain bread mold with visible sporangia.

26.3.3 Storage of Specimens

- Insects and mushrooms can be dried and stored in jars. However, they become brittle and break easily.
- A 10% solution of formaldehyde is the best way of storing specimens. Formaldehyde is often sold as a 40% solution. It should be stored in glass jars and out of the sun. Check specimens periodically for evaporation. Formaldehyde works because it is toxic; handle carefully.
- In a pinch, a 70% solution of ethanol can also be used to store insects, lizards, and worms. However, specimens sometimes decay in ethanol.

26.3.4 Sample Classification Practical

You have been provided with specimens L, M, N, O, and P.

1. Identify the specimens by their common names.
2. Classify each specimen to the phylum level.
3. Further classification:
   3.1. Write the classes of specimens L and M.
   3.2. List two observable differences between specimens L and M.
4. Explain why specimen P cannot grow taller.
5. Write down two distinctive characteristics of the phylum to which specimen O belongs.
6. Reproduction:
   6.1. List the modes of reproduction in specimens M and N.
   6.2. What are two differences between these modes of reproduction?

26.3.5 Sample Practical Solutions

1. Common names of specimens:
   - L: maize plant
   - M: bean plant
   - N: yeast
   - O: millipede
   - P: moss
2. Classifaction by kingdom and phylum:

Specimen	Kingdom	Phylym
L (maize plant)	Plantae	Angiospermophyta
M (bean plant)	Plantae	Angiospermophyta
N (yeast)	Fungi	Ascomycota
O (millipede)	Animalia	Arthropoda
P (moss)	Plantae	Bryophyta

3. Further classification:
   - Specimen L (maize plant): Class Monocotyledonae
   - Specimen M (bean plant): Class Dicotyledonae
   - Observable differences:

Specimen	Vein structure	Root structure
L (maize plant)	Parallel veins	Fibrous roots
M (bean plant)	Net veins	Tap roots

The answers to this question should be differences between monocots and dicots that the student can see by observing the plants with their naked eyes. Hence answers such as "vascular bundles in a ring" are not correct.

4. Specimen P (moss) cannot grow taller because it has no xylem and phloem. If it grew taller, it would not be able to transport food and water throughout the plant.

5. Characteristics of phylum Arthropoda:
   - jointed legs
   - segmented body
   - exoskeleton made of chitin

6. Reproduction

   6.1. Specimen M (bean plant) reproduces by sexual reproduction. Specimen N (yeast) reproduces by asexual reproduction.

Method	Genetic variation	Parents	Gametes
Asexual reproduction	There is no genetic variation between offspring.	Requires one parent only.	No gametes are involved.
Sexual reproduction	There is genetic variation between offspring.	Usually requires two parents.	Involves fusion of two gametes.

26.3.6 Additional Classification Questions

- Identify specimen X, Y, and Z by their common names.
- Classify specimens X, Y, and Z to the class level. (This means write the kingdom, phylum, and class.)
- Write the observable features of specimen X.
- List three observable differences/similarities between specimens X and Y.
- State the economic importance of specimen X.
- What characteristics are common among specimens X and Y?
- Why are specimens X and Y placed in different classes/phyla/kingdoms?
- Why are specimens X and Y classified under the same class/phylum/kingdom?
- What distinctive features place specimen X in its respective kingdom/phylum/class?
- How is specimen X adapted to its way of life?
- Suggest possible habitats for specimens X and Y.
- Which specimen is a primary producer/parasite/decomposer?
- For mushroom, yeast, bread mold, grasshopper, moss, tilapia, liver fluke, and tapeworm: Draw and label a diagram of specimen X.
- For tilapia: Draw a big and well-labeled diagram showing a lateral view of specimen X.
- For maize and bean:
  - Mention the type of pollination in specimen X [wind pollinated or insect pollinated].
  - How is specimen X adapted to this type of pollination?
  - Mention the type of germination [hypogeal or epigeal] in specimen X.
- For bean seed:
  - List three observable features of specimen X and state their biological importance.

  - Split specimen X into two natural halves. Draw and label the half containing the embryo.

- For fern:
  - Observe the underside of the leaves of specimen X
  - What is the name of the structures you have observed?
  - Give the function of the structures named above.
  - Draw specimen X and show the structures named above.

26.4 Respiration

The purpose of this practical is to investigate the properties of air exhaled from the lungs. This section contains the following:

- Limewater (properties and preparation)
- Apparatus
- Cautions and advice when using traditional materials
- Sample practical with solutions

26.4.1 Limewater

Limewater is a saturated solution of calcium hydroxide. It is used to test for carbon dioxide. When carbon dioxide is bubbled through limewater, the solution becomes cloudy. This is due to the precipitation of calcium carbonate by the reaction:

$CO_{2(g)} + Ca(OH)_{2(aq)} \longrightarrow CaCO_{3(s)}$

Limewater can be prepared from either calcium hydroxide or calcium oxide. Calcium oxide reacts with water to form calcium hydroxide, so either way you end up with a calcium hydroxide solution. Calcium oxide is the primary component in cement. Calcium hydroxide is available from building supply shops as *chokaa*.

To prepare lime water, add three spoons of fresh *chokaa* or cement to a bottle of water. Shake vigorously and then let stand until the suspended solids precipitate. Decant the clear solution. *Chokaa* produces a solution much faster than cement.

The exact mass of calcium hydroxide or calcium oxide used is not important. Just check whether some calcium hydroxide remains undissolved at the end – a sign that you have made a saturated solution.

Test limewater by blowing air into a sample with a straw. It should become cloudy. If it does not, then the concentration of $Ca(OH)_2$ is too low.

26.4.2 Apparatus

Many books call for delivery tubes, test tubes, and stoppers. These are totally unnecessary. Add the limewater to any small clear container and blow into it with a straw.

26.4.3 Cautions and Advice When Using Traditional Materials

If you use a delivery tube and pass it through a rubber stopper, do not use a single-holed stopper. This is what the pictures on NECTA practicals suggest, but it is a terrible idea. A single-holed stopper has no space for air to escape. So when a student blows air into the solution, the pressure in the test tube increases. The high pressure air then pushes limewater up the straw into the student's mouth. Alternatively, the student blows the stopper out of the test tube. If you use a stopper, use a double-holed stopper so that the extra air has a place to escape.

Is a glass delivery tube stuck in a rubber stopper? Do not pull hard on it. Just soak the stopper in warm water for a few minutes. The rubber will soften and the tube will come out.

Are your test tubes and delivery tubes cloudy after the practical? Clean them with dilute acid. This will dissolve any calcium carbonate that has been deposited on the glass.

26.4.4 Sample Respiration Practical

You have been provided with Solution B in a test tube. Use a delivery tube to breathe (exhale) into the solution until its color changes. (See diagram below.)

1. What is the aim of this experiment?
2. What is Solution B?
   2.1. What changes did you observe after breathing into Solution B?
   2.2. What can you conclude from these changes?
3. Breathe out over the palm of your hand. What do you observe?
4. Breathe out over a mirror. What do you observe?
5. Using your observations in the three experiments above, list three properties of exhaled air.
6. Explain why exhaled air is different from inhaled air. Where do the substances you identified in exhaled air come from?

26.4.5 Sample Practical Solutions

1. The aim of this experiment is to test exhaled air for carbon dioxide.
2. Solution B is limewater.
   2.1. Solution B became cloudy (or milky).
   2.2. Conclusion: exhaled air contains carbon dioxide.
3. Air breathed out over the palm of the hand is warm.
4. Droplets of water condense on the mirror.
5. Conclusions:
   - exhaled air contains carbon dioxide
   - exhaled air contains water
   - exhaled air is warm
6. Exhaled air contains the waste products of aerobic respiration. The carbon dioxide and water in exhaled air are products of respiration.

26.5 Transport

The purpose of this practical is to investigate osmosis by observing the changes in a leaf petiole placed in a hypotonic solution (water) and a hypertonic solution (water containing salt or sugar).

This section contains the following:

- Materials
- Sample practical with solutions
- Additional questions

26.5.1 Materials

The petiole is the stalk which attaches a leaf to a branch. The papaya leaf petioles in this practical should be soft petioles from young leaves, not stiff petioles from older leaves. Cut the petioles into pieces, and give each student two pieces of about 6 cm in length. Cylinders cut from a raw potato may be used instead of petioles.

The hypertonic solution may be made with by mixing either salt or sugar with water. The hypotonic solution is tap water.

26.5.2 Sample Transport Practical

Instructions

You have been provided with two pieces of a papaya leaf petiole, Solution A, and Solution B.

Use a razor blade to split the pieces of petiole longitudinally, up to a half of their length. You should have four strips at one end of each petiole, while the other end remains intact.

Place one petiole in solution A, and place the other petiole in solution B. Let the petiole sit for about ten minutes, then touch them to feel their hardness or softness.

Draw a sketch of each petiole after sitting in its respective solution for ten minutes.

Record your observations and explanations about the petioles in the table below.

Solution	Observation	Explanation
A		
B		

Questions

1. What was the aim of this experiment?
2. What was the biological process demonstrated by this experiment?
3. What is the importance of this process to plants?
4. Which solution contained:
   4.1. pure water
   4.2. a high concentration of solutes
5. What happened to the cells of the petioles in each solution? Illustrate your answer.
6. What would happen to the cells of the petioles in solution A if their cell walls were removed?

26.5.3 Sample Practical Solutions

(Assume Solution A is pure water, and Solution B is a concentrated solution of water and salt.)

Solution	Observation	Explanation
A	The petiole became hard (turgid)	Water diffused into the petiole cells
B	The petiole became soft (flaccid)	Water diffused out of the petiole cells

Answers to the questions

1. The aim of the experiment was to investigate the effect of osmosis on plant cells.
2. The experiment demonstrated osmosis.
3. Importance of osmosis in plants:
   3.1. Water enters plant cells by osmosis so that they become turgid. Turgor helps support the plant and hold it upright.
   3.2. Water diffuses into the xylem from the soil via osmosis.
4. Solution identification
   4.1. Pure water: Solution A.
   4.2. High concentration of solutes: Solution B
5. [Illustrations]
6. The petiole cells would burst in Solution A if their cell walls were removed.

26.5.4 Additional Questions

You can extend this experiment by giving students two pieces of meat in addition to the petioles. The piece of meat placed in pure water should expand and become soft due to the cells bursting. The piece placed in salt water should shrink and become hard due to water diffusing out of the cells. This experiment helps to teach the different effects of osmosis on plant and animal cells.

If your school has a good microscope, try observing plant cells under the microscope after letting them sit in hypotonic and hypertonic solutions.

You can add critical thinking questions to the practical that require the student to use their knowledge of osmosis. For example:

- Why does a freshwater fish die if it is placed in salt water?
- Why do merchants spray vegetables with water in the market?
- You can die if a doctor injects pure water into your bloodstream. Why?

26.6 Photosynthesis

The purpose of this practical is to prove that chlorophyll is required for photosynthesis. This is done by using iodine to test a variegated leaf for starch. The parts of the leaf containing chlorophyll are expected to contain starch, while the parts lacking chlorophyll are expected to lack starch.

This section contains the following:

- Procedure
- Cautions
- Materials and where to find them
- Sample practical with solutions
- Additional practicals

26.6.1 Procedure

1. Use iodine tincture from the pharmacy without dilution.
2. Prepare hot water bathes. The water should be boiling.
3. While the water gets hot, send the students to gather small leaves. The best have no waxy coating and are varigated (have sections without green).
4. The leaves should be boiled in the hot water bath for one minute.
5. Each group should then move its leaf into their test tube and cover it with methylated spirit.
6. Each group should then heat their test tube in a water bath. Over time, the leaf should decolorize and the methylated spirit will turn bight green. The chlorophyll has been extracted and moved to the spirit. A well chosen leaf should turn completely white, although this does not always happen.
7. After decolorization, dips the leaves briefly in the hot water.
8. For leaves that turn white, students should test them for starch with drops of iodine solution.

26.6.2 Cautions

Ethanol is flammable! It should never be heated directly on a flame. Use a hot water bath – place a test tube or beaker of ethanol in a beaker or bowl of hot water and let it heat slowly. The boiling point of ethanol is lower than the boiling point of water, so it will start boiling before the water. If the ethanol does catch fire, cover the burning test tube with a petri dish or other non-flammable container to extinguish the flame.

26.6.3 Materials and Where to Find Them

- Variegated leaf: this is a leaf that contains chlorophyll in some parts, but not in others. Often variegated leaves are green and white or green and red. Look at the flower beds around the school and at the teachers' houses – they often contain variegated leaves. Test the leaves before the practical, as some kinds are too waxy to be decolorized by ethanol. Also, check for chlorophyll by looking at the underside of the leaves; the leaves you use have at least a small section of white on their undersides, signifying a lack of chlorophyll.
- Source of heat: anything that boils water – Motopoa is best, followed by kerosene and charcoal
- Ethanol: use the least expensive strong ethanol available; this is probably methylated spirits unless your village specializes in high proof gongo.

26.6.4 Sample Photosynthsis Practical

You have been provided with specimen G.

1. Identify specimen G.
2. Make a sketch showing the color pattern of specimen G. Carry out the following experiment:
   - 2.1. Place specimen G in boiling water for one minute.
   - 2.2. Boil specimen G in ethanol using a hot water bath. Do not heat the ethanol directly on a flame.
   - 2.3. Remove specimen G from the ethanol. Dip it in hot water.
   - 2.4. Spread specimen G on a white tile and drip iodine solution onto it. Use enough iodine to cover the entire specimen.
   - 2.5. Make a sketch showing the color pattern of specimen G at the end of the experiment.
3. What was the aim of this experiment?
4. Why was specimen G
   - 4.1. Boiled in water for one minute
   - 4.2. Boiled in ethanol
   - 4.3. Dipped in hot water at the end of the experiment
5. What was the purpose of the iodine solution?
6. Why was the ethanol heated using a hot water bath?
7. What can you conclude from this experiment? Why?

26.6.5 Sample Practical Solutions

1. Specimen G is a variegated leaf.
2. Drawing: See diagram above.
3. The aim of this experiment was to investigate whether chlorophyll is required for photosynthesis.
4. Specimen G was:
   - 4.1. boiled in water to kill the cells and stop all metabolic processes.
   - 4.2. boiled in ethanol to decolorize it (to remove the chlorophyll).
   - 4.3. dipped in hot water to remove the ethanol. (If ethanol is left on the leaf it will become hard and brittle.)
5. The purpose of the iodine solution was to test for starch.
6. The ethanol was heated using a hot water bath because ethanol is flammable.
7. The experiment shows that chlorophyll is required for photosynthesis. We know this because the parts of the leaf containing chlorophyll also contained starch, which is a product of photosynthesis. Thus, the parts of the leaf containing chlorophyll performed photosynthesis. The parts of the leaf lacking chlorophyll lacked starch. Hence, these parts of the leaf did not perform photosynthesis.

26.6.6 Additional Practicals

To test if light is required for photosynthesis

Take a live plant, and leave it in the dark for 24 hours to destarch all leaves. Then, cover some of its leaves with cardboard or aluminum foil, while leaving others uncovered. Let the plant sit in bright light for several hours. Give each group of students one leaf that was covered in cardboard, and one leaf that was uncovered. Have them use the procedure above to test for starch. They should find that the covered leaf contains no starch, while the uncovered leaf contains starch.

A cool variation on this experiment is to cover leaves with pieces of cardboard that have letters or pictures cut out of them. The area where the cardboard is cut out will perform photosynthesis and produce starch. When the students do a starch test, a blue-black letter or picture will appear on the leaf.

To prove that oxygen is a product of photosynthesis

This experiment requires a water plant. Basically, place a live water plant under water*, then cover it with an inverted funnel. Place an upside-down test tube filled with water on top of the funnel. Let the plant sit in bright light until the water in the test tube is displaced and the test tube fills with gas. Use a glowing splint to test the gas – if it is oxygen, it will relight the splint.

*Note: some books suggest putting sodium bicarbonate (baking soda) in the water.

Chapter 27

Chemistry Practicals

27.1 Introduction to Chemistry Practicals

27.1.1 Format

The format of the Chemistry practical exam was revised in 2011 to keep up with the 2007 updated syllabus. As such, there will be no further Alternative to Practical exams, pending approval from the Ministry of Education. Prior to 2011, students only had to answer 2 of the 3 questions, including Question 1.

As of now, the Chemistry practical has 3 questions and students must answer **all** of them. Question 1 is on Volumetric Analysis and Laboratory Techniques and Safety. Question 2 is taken from Ionic Theory and Electrolysis/Chemical Kinetics, Equilibrium and Energy. Question 3 is on Qualitative Analysis. Question 1 is worth 20 marks, while Questions 2 and 3 carry 15 marks each. Students have $2\frac{1}{2}$ hours to complete the exam.

Students are allowed to use Qualitative Analysis guidesheet pamphlets in the examination room.

Chemistry 1 Theory Format

The theory portion of the Chemistry exam comprises 100 marks, while the practical carries 50 marks. A student's final grade for Chemistry is thus found by taking her total marks from both exams out of 150.

The theory exam for Chemistry contains 3 sections. Section A has 2 questions and is worth 20 marks - Question 1 is 10 multiple choice and Question 2 is 10 matching. Section B has 9 short answer questions, each having two items, for a total of 54 marks. Section C has 2 essay questions without items for a total of 26 marks. Students are required to answer all questions.

Note This information is current as of the time of publication of this manual. Updated information may be obtained by contacting the Ministry of Education.

27.1.2 Notes for Teachers

NECTA Advance Instructions

There are two sets of advance instructions. One set of advance instructions are given to teachers at least one month before the date of the exam. These instructions contain the list of apparatus, chemicals, and other materials required for preparing the Chemistry practical questions. The instructions also give suggestions on the amount of chemicals that should be available for each candidate to use.

The second set of instructions should be given 24 hours before the time of the practical. It includes which chemicals and apparatus should be given to each candidate (or shared among candidates) for each of the three practical questions. These instructions also state how to label each solution and/or compound.

The bottom of the 24 Hours Advance Instructions also states that the Laboratory Technician or Head of Chemistry Department should perform some of the experiments immediately after the last session of the examination. It is only required to perform the titration and chemical kinetics experiments. This is required to be done for every school and is used as a reference for the markers in case the water, chemicals, and apparatus are not the same at every school. This is enclosed and submitted together with the students' test papers and may be used as a marking scheme. It is also advised that any notes, comments or concerns for the markers be included at this time.

27.1.3 Common Practicals

Volumetric Analysis determine the concentration of a solution of a known chemical by reacting it with a known concentration of another solution

Qualitative Analysis systematically identify an unknown salt through a series of chemical tests

Chemical Kinetics and Equilibrium observe changes in chemical reaction rates by varying conditions such as temperature and concentration

Note These are the most common practicals, but they are not necessarily the only practicals that can occur on a NECTA exam. Although the updated exam format lists Questions 1 and 3 as Volumetric Analysis and Qualitative Analysis respectively, Question 2 can come from a variety of topics which may not yet have been used in older past papers. Be sure to regularly check the most recent past NECTA papers to get a good idea of the types of questions to expect.

27.2 Volumetric Analysis

This section contains the following:

- Volumetric Analysis Theory
- Substituting Chemicals in Volumetric Analysis
- Properties of Indicators
- Traditional Volumetric Analysis Technique
- Sample Practical Question

27.2.1 Volumetric Analysis Theory

Volumetric Analysis is a method to find the concentration (molarity) of a solution of a known chemical by comparing it with the known concentration of a solution of another chemical known to react with the first.

For example, to find the concentration of a solution of citric acid, one might use a 0.1 M solution of sodium hydroxide because sodium hydroxide is known to react with citric acid.

The most common kinds of volumetric analysis are for acid-base reactions and oxidation-reduction reactions. Acid-base reactions require use of an indicator, a chemical that changes color at a known pH. Some oxidation-reduction reactions require an indicator, often starch solution, although many are self-indicating, (one of the chemicals itself has a color).

See also the sections on Preparation of Solutions (p. 63), Preparation of Solutions Without a Balance (p. 65) and Relative Standardization (p. 66) in Laboratory Techniques.

The process of volumetric analysis is often called *titration.*

27.2.2 Substituting Chemicals in Volumetric Analysis

Theory

The volumetric analysis practical exercises sometimes call for expensive chemicals, for example potassium hydroxide or oxalic acid. As the purpose of exercises and exams is to train or test the ability of the students and not the resources of the school, it is possible to use different chemicals as long as the solutions are calibrated to give equivalent results. For example, if the instructions call for a potassium hydroxide solution, you can use sodium hydroxide to prepare this solution. It will not affect the results of the practical – if you make the correct calibration. How to calibrate solutions when substituting chemicals is the subject of this section.

Technically, only two chemicals are required to perform any volumetric analysis practical: one acid and one base. The least expensive options are sulfuric acid, as battery acid, and sodium hydroxide, as caustic soda. To substitute one chemical for another in volumetric analysis, the resulting solution must have the same normality (N).

- For all monoprotic acids (HCl, ethanoic acid), the normality is the molarity.
  Example: 0.1 M ethanoic acid = 0.1 N ethanoic acid
- For diprotic acids (sulfuric acid, ethandiotic acid), the normality is twice the molarity, because each molecule of diprotic acid brings two molecules of H^+.
  Example: 0.5 M sulfuric acid = 1.0 N sulfuric acid

- For the hydroxides and hydrogen carbonates used in ordinary level (NaOH, KOH, $NaHCO_3$), the normality is the molarity.
  Example: 0.08 M KOH = 0.08 N KOH

- For the carbonates most commonly used (Na_2CO_3, $Na_2CO_3 \cdot 10H_2O$, K_2CO_3), the normality is twice the molarity.
  Example: 0.4 M Na_2CO_3 = 0.8 N Na_2CO_3

Substitution Calculations

When instructions describe solutions in terms of molarity, calculating the molarity of the substitution is relatively simple. For example, suppose we want to use sulfuric acid to make a 0.2 M solution of ethanoic acid. 0.2 M ethanoic acid is 0.2 N ethanoic acid which will titrate the same as 0.2 N sulfuric acid. 0.2 N sulfuric acid is 0.1 M sulfuric acid, and thus we need to prepare 0.1 M sulfuric acid.

When instructions describe solutions in terms of concentration ($^{g}/_{L}$), we just need to add an extra conversion step. For example, suppose we want to use sodium hydroxide to make a 14.3 $^{g}/_{L}$ solution of sodium carbonate decahydrate. 14.3 $^{g}/_{L}$ sodium carboante decahydrate is 0.05 M sodium carbonate decahydrate which is 0.1 N sodium carbonate decahydrate. This will titrate the same as 0.1 N sodium hydroxide, which is 0.1 M sodium hydroxide or 4 $^{g}/_{L}$ sodium hydroxide, and thus we need to prepare 4 $^{g}/_{L}$ sodium hydroxide to have a solution that will titrate identically to 14.3 $^{g}/_{L}$ sodium carbonate decahydrate.

Common Substitutions

To simplify future calculations, we have prepared general conversions for the most common chemicals used in volumetric analysis. Remember to check all final solutions with relative standardization to ensure that they indeed give the correct results.

Required Chemical	Low Cost Alternative	Substiution Method	Molarity Example	Concentration Example
Hydrochloric Acid	Sulfuric Acid (Battery Acid)	If you are required to prepare an X molarity solution of HCl, prepane a X×0.5 molarity solution of battery acid	The instructions call for 0.12 M HCl. Instead, prepare 0.06 M sulfuric acid	
Ethanoic (Acetic) Acid	Sulfuric Acid (Battery Acid)	If you are required to prepare an M molarity solution of ethanoic acid, prepare a M×0.5 molarity solution of sulfuric acid	The instructions call for 0.10 M ethanoic acid. Prepare 0.05 M sulfuric acid.	
Ethandioic (Oxalic) Acid dihydrate ($C_2H_2O_4 \cdot 2H_2O$)	Sulfuric Acid (Battery Acid)	If you are required to prepare an M molarity solution of ethandioic acid, prepare an M molarity solution of sulfuric acid. If you are required to prepare a C concentration solution of ethandioic acid, prepare a $^{C}/_{126}$ molarity solution of sulfuric acid.	The instructions call for 0.075 M ethandioic acid. Prepare 0.075 M sulfuric acid.	The instructions call for 6.3 $^{g}/_{L}$ ethandioic acid. Prepare 0.05 M sulfuric acid.
Potassium Hydroxide	Sodium Hydroxide (Caustic Soda)	For M molarity potassium hydroxide, make M molarity sodium hydroxide. For C concentration potassium hydroxide, make C×$^{40}/_{56}$ concentration sodium hydroxide.	The instructions call for 0.1 M potassium hydroxide. Just prepare 0.1 M sodium hydroxide.	The instructions call for 2.8 $^{g}/_{L}$ potassium hydroxide. Prepare 2 $^{g}/_{L}$ sodium hydroxide.
Anhydrous Sodium Carbonate	Sodium Carbonate Decahydrate (Soda Ash)	For M molarity anhydrous sodium carbonate, make M molarity sodium carbonate decahydrate. For C concentration anhydrous sodium carbonate, make C×$^{286}/_{106}$ sodium carbonate decahydrate.	The instructions call for 0.09 M anhydrous sodium carbonate. Make 0.09 M sodium carbonate decahyrate.	The instructions call for 5.3 $^{g}/_{L}$ anhydrous sodium carbonate. Make 14.3 $^{g}/_{L}$ sodium carbonate decahydrate.
Anhydrous Sodium Carbonate	Sodium Hydroxide (caustic soda)	For M molarity anhydrous sodium carbonate, make M×2 molarity sodium hydroxide. For C concentration anhydrous sodium carbonate, make C×2×$^{40}/_{106}$ sodium hydroxide.	The instructions call for 0.09 M anhydrous sodium carbonate. Make 0.18 M sodium hydroxide.	The instructions call for 5.3 $^{g}/_{L}$ anhydrous sodium carbonate. 4.0 $^{g}/_{L}$ sodium hydroxide.
Sodium Carbonate Decahydrate ($Na_2CO_3 \cdot 10H_2O$)	sodium hydroxide (caustic soda)	For M molarity sodium carbonate ecahydrate, make M×2 molarity sodium hydroxide. For C concentration sodium carbonate decahydrate, make C×2 ×$^{40}/_{286}$ sodium hydroxide.	The instructions call for 0.09 M sodium carbonate decahydrate. Make 0.18 M sodium hydroxide.	The instructions call for 14.3 $^{g}/_{L}$ sodium carbonate decahydrate. Make 4.0 $^{g}/_{L}$ sodium hydroxide.
Anhydrous Potassium Carbonate	Sodium Carbonate decahydrate (Soda Ash)	For M molarity potassium carbonate, make M molarity sodium carbonate decahydrate. For C concentration potassium carbonate, make C×$^{286}/_{122}$ concentration sodium carbonate.	The instructions call for 0.08 M anhydrous potassium carbonate. Prepare 0.08 M sodium carbonate decahydrate.	The instructions call for 6.1 $^{g}/_{L}$ anhydrous potassium carbonate. Prepare 14.3 $^{g}/_{L}$ sodium carbonate decahydrate.
Anhydrous Potassium Carbonate	Sodium Hydroxide (caustic soda)	For M molarity potassium carbonate, make M×2 molarity sodium hydroxide. For C concentration potassium carbonate, make C×2 ×$^{40}/_{122}$ concentration sodium hydroxide.	The instructions call for 0.08 M anhydrous potassium carbonate. Prepare 0.16 M sodium hydroxide.	The instructions call for 6.1 $^{g}/_{L}$ anhydrous potassium carbonate. Prepare 4.0 $^{g}/_{L}$ sodium hydroxide.

Additional Notes

- In volumetric analysis experiments with two indicators, it is not possible to substitute one chemical for another as the acid/base dissociation constant is critical and specific for each chemical. It is still possible to substitute sodium carbonate decahydrate for anhydrous sodium carbonate with the above conversion.

- These substitutions only work for volumetric analysis. In qualitative analysis, the nature of the chemical matters. If the instructions call for sodium carbonate, you cannot provide sodium hydroxide and expect the students to get the right answer!

27.2.3 Properties of Indicators

Acid-base Indicators

These indicators are chemicals that change colors in a specific pH range, which makes them suited to use in acid-base reactions. When the pH of changes from low pH to high pH or from high to low, the color of the solution changes.

Four common acid-base indicators are methyl orange (MO), phenolphthalein (POP), bromothymol blue (BB), and universal indicator (U)

- Methyl Orange, MO, is always used when titrating a strong acid against a weak base. The pH range of MO is 4.0-6.0 and thus no color change is observed until the base is completely neutralized. If you use MO with a weak acid, the color might start to change before completely neutralizing the acid.

- Phenolphthalein, POP, is always used when titrating a weak acid against a strong base. The pH range of POP is 8.3-10.0, and thus no color change is observed until the weak acid is completely neutralized. If you use POP with a weak base, the color might start to change before completely neutralizing the base.

- Bromothymol Blue, BB, is used in the same manner as methyl orange.

- Universal indicator, U, is not suitable for volumetric analysis involving either weak acids or bases as it changes color continuously rather than in a limited pH range. It is very useful for tracking the pH continuously over a titration, perhaps by performing two titrations side by side, one with a standard indicator and another with universal indicator.

Any indicator can be used when titrating a strong acid against a strong base. Universal indicator, however, will not produce very accurate results.

No indicator is suitable for titrating a weak acid against a weak base.

In some experiments, more than one indicator may be used in the same flask, for example when titrating a mixture of strong and weak acids or bases.

Colors of Indicators The colors of the above indicators in acid and base are:

Indicator	Acid	Neutral	Base
Methyl Orange	Red	Orange	Yellow
Phenolphthalein	Colorless	Colorless	Pink
Bromothymol Blue	Yellow	Blue	Blue
Universal Indicator	Red, Orange, Yellow	Yellow/Green	Green, Blue, Indigo

Titration is finished when the indicator starts a permanent color change. For example, when methyl orange turns orange, the titration is finished. If students wait until methyl orange turns pink (or yellow) they have overshot the endpoint of the titration, and their volume will be incorrect. Likewise, POP indicates that the titration is finished when it turns light pink. If students wait until they have an intensely pink solution, they will use too much base and get the wrong answer.

Note that light pink POP solutions may turn colorless if left for a few minutes. This is due to carbon dioxide in the air reacting to neutralize bases in solution.

Note on technique Students should use as little acid-base indicator as possible. This is because some acid or base is required to react with the indicator so that it changes color. If a lot of indicator is used, students will add more acid or base than they need.

Other Indicators

Starch indicator is used in oxidation-reduction titrations involving iodine. This is because iodine forms an intense blue to black colored complex in the presence of starch. Thus starch allows a very sensitive assessment of the presence of iodine in a solution.

It is important to add the starch indicator close to the end point when there is an acid present. The acid will cleave the starch and that will prevent the starch from working properly. Students using starch should use a pilot run to get an idea when to add the starch indicator.

Preparation of Indicators

- Methyl orange (MO): if you have a balance, weigh out about 1 g of methyl orange powder and dissolve it in about 1 L of water. Store the solution in a plastic water bottle with a screw on cap and it will keep for years. If it gets thick and cloudy, add a bit more water and shake. If you do not have a balance, add half of a small tea spoon to a liter of water.
- Phenolphthalein (POP) Dissolve about 0.2 g of phenolphthalein powder in 100 mL of pure ethanol; then add 100 mL water with constant stirring. If you use much more water than ethanol, solid phenolphthalein will precipitate. Store POP in a plastic water bottle with a screw on cap. We recommend making POP in smaller quantities than MO as it does not keep as well, mostly due to the evaporation of ethanol. If the solution develops a precipitate, add a bit of ethanol and shake. We do not recommend using purple methylated spirits as a source of ethanol for making POP. You can distill purple spirits to make clear spirits. For clear methylated spirits, use 140ml of spirit and 60ml of water, as spirits generally are already 30% water.
- Starch: place about 1 g of starch in 10 mL of water in a test tube. Mix well. Pour this suspension into 100 mL of boiling water and continue to boil for one minute or so. Alternatively, use the water leftover after boiling pasta or potatoes. If this is too concentrated, dilute it with regular water.
- The authors have never prepared bromothymol blue or universal indicator from powder, but suspect their preparation is similar to methyl orange.

Note that the exact mass of indicator used is not very important. You just need to use enough so that the color is clearly visible. Students use very little indicator in each titration, and a liter of indicator solution should last you a long time.

27.2.4 Traditional Volumetric Analysis Technique

The Volumetric Analysis practical consists of an acid that is being titrated acid against a base until neutralization, in order to determine the concentration of the base.

On NECTA practical exams, titrations are done four times: a pilot followed by three trials. The pilot is done quickly and is used to determine the approximate volume needed for neutralization to speed up the following trials.

Ex: If the pilot gives an end point of 25.00 mL, then for the three subsequent trials, 20.00 mL can quickly be added from the burette. Then begin to add solution slowly until the endpoint is reached.

Results from the pilot are not accurate and are not included when doing calculations. Students should also know that not all three trials are always used in calculating the average volume used. Values of trials must be consistent and within $\pm$ 02 cm^3 of each other to be valid for average volume determination.

Volumetric Analysis Using Burettes

Preparation

1. After washing Burettes thoroughly, rinse the Burette with 3 mL of the acidic solution that will be used during the titration (Acid usually goes in the Burette).

- Cover the entire inside surface of the Burette.
- Discard 3 mL of solution properly when finished .
- Why? This prevents dilution of acid by water.

2. After washing the flask thoroughly, rinse the flask with 3 mL of solution that will be used during the titration (Base usually goes in the flask).
   - Cover the entire inside surface of the flask.
   - Discard 3 mL of solution properly when finished.

Procedure

1. Clean the burette with water.
2. Rise the burette with the acid that will be used for the titration.
3. Fill the burette with the acid. Let a little run out through the stopcock.
4. Record the initial burette reading.
5. Use a syringe to transfer the base solution into a conical flask.
6. Record the volume moved by the syringe.
7. If you are using an indicator, add a few drops to the flask.
8. Slowly add the acid from the burette to the flask. Swirl the flask as you titrate. Be careful. Avoid acid drops landing on the sides of the flask.
9. Stop titration when the slight color change become permanent. This is the end point.
10. Record final reading of the burette.
11. Repeat for remaining titrations.

Notes

- Burettes tell you the volume of solution used, not the volume present.
  Ex: Initial Reading - 4.23 mL
  Final Reading - 20.57 mL
  You used 16.34 mL of acid during the titration.
- Reading Measurements
  - Always read burettes at eye level.
  - Always read from the bottom of the meniscus. In a plastic apparatus, there is often no meniscus.
  - Burettes are accurate to 2 decimal places. Students should estimate to the nearest 0.01 mL
- For Acid-Base indicators: The less indicator used, the better. To change color, the indicator must react with fluid in the burette. If you add too much, it uses more chemical than necessary for neutralization, creating an indicator error.
- For starch indicators: use 1 mL. Starch is not titrated; indicators are, and you must use more to get a good color change.

Volumetric Analysis Without Using Burettes

Use plastic syringes instead of burettes.

As of late 2010, the most precise syringes available are the 10 mL NeoJect brand - you should use these (A titration with 2 plastic syringes is more accurate than a titration with a burette and a cheap glass pipette).

If use of these syringes is new to you, read Use of a Plastic Syringe to Measure Volume (p. 58) before continuing.

If students are using syringes in place of burettes, they require two syringes for the practical: one to use as a burette (for acid) and one to use as a pipette to transfer base into the flask. It may be useful to label the different syringes "burette" /"flask" or "acid"/"base".

Preparation (without burettes)

1. Clean the "pipette" syringe with water.
2. Rinse the "pipette" syringe with base solution that will be put into the flask.
3. Use the "pipette" syringe to transfer base into the flask. To do this accurately, first add 1 mL of air to the syringe and then suck up the base beyond the desired amount. Push back the plunger until the top of the fluid is at the required volume.
4. Record the total volume transferred (multiple transfers with the 1 syringe may be required to react the desired volume).
5. If you are using indicator, add a few drops to the flask.
6. Clean the "burette" syringe with water.
7. Rinse the "burette" syringe with the acid solution that will be used for titration.

Procedure (without burettes)

1. Add 1 mL of air to the syringe and suck up the acid beyond the 10 mL mark. Slowly push back the plunger until the top of the fluid is exactly at the 10 mL line.
2. Slowly add acid from the "burette" syringe into the flask. Swirl the flask as you titrate. Be careful. Make sure the acid lands in the base, avoid acid drops landing on the sides of the flask.
3. Stop titration when the slight color change become permanent. This is the end point.
4. Often, more than 10 mL of acid will need to be used. This is not a problem. Once 10 mL is finished in the syringe, students should just fill it up again and continue the titration.
5. Record final volume of acid transferred by the "burette" syringe.

Notes for when using syringes in place of burettes

- Students must record their results in a manner that is consistent with traditional reporting.
- On rough paper, students should calculate the volume of solutions used during titration. If they only used one syringe and the initial volume in the syringe was 10.00 mL and the final volume was 2.55 mL, the student used 7.45 mL of solution. If they used two full syringes and then part of a third (which had the initial reading of 10.00 mL and a final reading of 4.65 mL), the student used 5.35 mL + 10.00 mL + 10.00 mL = 25.35 mL total.
- In the table of results, the student should write 25.35 mL for Volume Used. If they had used a burette, the initial reading would have been 0.00 mL and the final reading would have been 25.35 mL. This is what they should write in their table of results.
- When using a syringe as a burette, students should write 0.00 mL as the Initial Volume and then, for the Final Volume, they should write the number they calculated for the total volume used.

27.2.5 Common Calculations in Titration Experiments

All NECTA practical experiments require students to determine some unknown in the titration procedure. Common calculations that the problem statement will ask for include:

- Concentration (molarity) of an acid or base
- Relative atomic mass of unknown elements in an acid or base
- Percentage purity of a substance
- Amount of water of crystallization in a substance

Concentration of an Acid or Base

The problem statement may have the student find either the unknown molarity (moles per litre) or concentration (grams per litre) of the acid or the base. As an example, the following steps are used to calculate the unknown concentration of an acid:

1. *Calculate the average volume of acid used.*
   Remember to not use the pilot trial or any trials that are not within $\pm$ 0.2 cm^3 of each other.
2. *Calculate the number of moles of the base used.*

$$\text{Molarity} = \frac{\text{number of moles}}{\text{volume of solution}}$$

   These values can usually be taken from the solutions listed on the test paper. Also be sure that the units of volume of solution are in litres or dm^3.
3. *Write a balanced chemical equation for the reaction.*
   The chemical equation can also be written as an ionic equation.
4. *Calculate the number of moles of acid used from the mole ratio taken from the balanced chemical equation.*
   Both ionic and full formulae equations give the same mole ratio.
5. *Work out the molar concentration of the acid.*
   The molar concentration can be determined using the calculated number of moles of acid (found in the previous step) and the average volume of acid used (found in step 1), using the equation in step 2.
   Alternatively, the following equation can be used:

$$\frac{C_A V_A}{C_B V_B} = \frac{n_A}{n_B}$$

   where:
   C_A is the molar concentration of the acid,
   V_A is the volume of the acid used,
   n_A is the number of moles of the acid used,
   C_B is the molar concentration of the base,
   V_B is the volume of the base used, and
   n_B is the number of moles of the base used.

Similar steps are used to calculate the unknown concentration of a base.
Repeat steps 1 through 5, but with the following changes:

- Step 2: Calculate the moles of the acid used.
- Step 4: Calculate the moles of the base from the mole ratio.
- Step 5: Find the molar concentration of the base, either using the molarity calculation or the equation above.

Relative Atomic Mass of Unknown Elements

Atomic mass of unknown elements, as well as molecular mass of compounds with unknown elements may need to be calculated in the problem statement. Most unknown elements will be a metal of a basic compound. As an example, the following steps are used to calculate the relative atomic mass of an unknown metal element of a metal carbonate:

1. *Calculate the average volume of acid used.*
   Remember to not use the pilot trial or any trials that are not within $\pm$ 0.2 cm^3 of each other.

2. *Calculate the number of moles of the acid used.*

$$\text{Molarity} = \frac{\text{number of moles}}{\text{volume of solution}}$$

   These values can usually be taken from the solutions listed on the test paper. Also be sure that the units of volume of solution are in litres or dm^3.

3. *Write a balanced chemical equation for the reaction to get the mole ratio.*

4. *Determine the number of moles of the metal carbonate used.*
   This can be taken from the balanced chemical equation.

5. *Work out the molecular concentration of the metal carbonate solution.*
   Use the formula as shown in step 2.

6. *Calculate the mass of the metal carbonate in one litre of solution.*
   This can be done using the following ratio:

$$\frac{\text{mass given in problem statement}}{\text{volume given in problem statement}} = \frac{\text{mass of unknown metal}}{\text{one litre}}$$

   Make sure the units correspond because sometimes the problem statement will be expressed in dm^3 or cm^3.

7. *Using the molarity of the solution and the mass of the metal carbonate per litre of solution, work out the relative molecular mass of the metal carbonate.*
   The following equation can be used to calculate molar mass:

$$\text{molar mass} = \frac{\text{mass per litre}}{\text{molarity}}$$

8. *Calculate the relative atomic mass of the metal based on the formula of the carbonate.*
   Use the total molar mass of the compound found in step 7 and the molar mass of each element in the compound to find the molar mass of the unknown element.

Some problem statements may require the student to identify the unknown element from its molecular mass.
Similar steps should be followed if the unknown element is of an acidic compound. Just replace the steps that include the metal carbonate solution with the acid solution.

Percentage Purity of a Substance

Problem statements that require the student to find percentage purity will usually contain one solution in the list provided that specifically states it is impure or that it is a hydrated compound (seems very low in concentration). Again, it is possible to determine percentage purity of an acid or a base. As an example, the following steps are used to calculate the percentage purity of a base:

1. *Determine the average volume of the acid used.*
   Remember to not use the pilot trial or any trials that are not within $\pm$ 0.2 cm^3 of each other.

2. *Calculate the number of moles of the acid used.*

$$\text{Molarity} = \frac{\text{number of moles}}{\text{volume of solution}}$$

   These values can usually be taken from the solutions listed on the test paper. Also be sure that the units of volume of solution are in litres or dm^3.

3. *Write a balanced chemical equation for the reaction to get the mole ratio.*

4. *Determine the number of moles of base used in the reaction.*
   This can be taken from the mole ratio from the previous step.

5. *Calculate the mass of the base used in the reaction.*
   The mass can be determined by the number of moles calculated and the following relationship:

$$\text{mass} = \text{number of moles} \times \text{molar mass}$$

6. *Work out the percentage purity of the base solution sample.*
   The following equation for percentage purity should be used:

$$\text{percentage purity} = \frac{\text{mass of pure substance in sample}}{\text{mass of the impure sample}} \times 100\%$$

   It is very important to note that when calculating percentage purity, the amount of volume in the concentration of base must be equal to the volume of concentration of acid used. For example, if there was 0.424 g of sodium carbonate in 25 cm^3 of solution reacting with a 250 cm^3 solution of acid, the mass of sodium carbonate must be converted to know the mass in 250 cm^3. Therefore, 250 cm^3 of base solution will contain 4.24 g, not 0.424 g.
   The value for the mass of the impure sample comes from the list of provided solutions and the mass of the pure sample will come from the calculations.

Similar steps can be followed to find the percentage purity of a the acid solution sample. Instead of finding the mass of the base, use the calculated moles of acid used to find the mass of acid in the actual reaction.

Amount of Water of Crystallization

Water of crystallization is the water that is bound within crystals of substances. Most hydrated substances and solutions contain water of crystallization. Problem statements that ask students to determine the amount of water of crystallization will have a solution with a formula similar to [base]·xH_2O, and they have to solve for x. As an example, the following steps are used to determine the number of molecules of water of crystallization in a hydrated base compound sample:

1. *Calculate the average volume of the acid used.*
   Remember to not use the pilot trial or any trials that are not within ± 0.2 cm^3 of each other.

2. *Calculate the number of moles of the acid used.*

$$\text{Molarity} = \frac{\text{number of moles}}{\text{volume of solution}}$$

   These values can usually be taken from the solutions listed on the test paper. Also be sure that the units of volume of solution are in litres or dm^3.

3. *Write a balanced chemical equation for the reaction to get the mole ratio.*

4. *Calculate the number of moles of the base used.*
   This can be determined from the mole ratio in the previous step.

5. *Determine the molar concentration of the base.*
   The molarity can be calculated using the volume of base used in the experiment and the equation from step 2.

6. *Calculate the relative molecular mass (R.M.M.) of the base compound.*
   The following equation can be used to calculate molar mass:

$$\text{molar mass} = \frac{\text{mass per litre}}{\text{molarity}}$$

7. *Determine the number of molecules of water of crystallization in the sample.*
   Using the relative atomic masses of the various atoms in the base compound, subtract the mass of the compound from the total mass of the hydrated compound. Water molecules always have a total molecular mass of 18 g/mol, so the remaining mass will be composed of multiples of 18. For example, if a hydrated carbonate ($Na_2CO_3 \cdot xH_2O$) has a total mass of 286 g, the molecules of water can be determined as follows:
   $2\text{Na} + \text{C} + 3\text{O} + x(2\text{H} + \text{O}) = 286$
   $(2 \times 23) + 12 + (3 \times 16) + x[(2 \times 1) + 16] = 286$
   $106 + 18x = 286$
   $18x = 180$
   $x = 10$
   Therefore, in this example, there are 10 molecules of water of crystallization in the hydrated sodium carbonate ($Na_2CO_3 \cdot 10H_2O$) sample.

27.2.6 Sample Practical Question

The following is a sample practical question from 2012.

You are provided with the following solution:

TZ: Containing 3.5 g of impure sulphuric acid in 500 cm^3 of solution;
LO: Containing 4 g of sodium hydroxide in 1000 cm^3 of solution;
Phenolphthalein and Methyl indicators.

Questions:

(a) (i) What is the suitable indicator for the titration of the given solutions?
Give a reason for your answer.
(ii) Write a balanced chemical equation for the reaction between **TZ** and **LO**.
(iii) Why is it important to swirl or shake the contents of the flask during the addition of the acid?

(b) Titrate the acid (in a burette) against the base (in a conical flask) using two drops of your indicator and obtain three titre values.

(c) (i) ____ cm^3 of acid required ____ cm^3 of base for complete reaction.
(ii) Showing your procedures clearly, calculate the percentage purity of **TZ**. **(20 marks)**

Discussion

This practical question requires students to know and understand how to use volumetric analysis apparatus and technique. Since this question involves the titration of sulphuric acid (strong acid) and sodium hydroxide (strong base), either phenolphthalein or methyl orange are acceptable indicators to use. An explanation of suitable indicators can be found in Acid-base Indicators (p. 98).

Make sure that students create a table for the first pilot titration and three titre values, for a total of four titrations. Only the titre values (not the pilot) that are within ±0.02 ml of each other will be used to calculate the average titrated volume. Students should also be swirling the contents of the volumetric flask in order to thoroughly mix the acid and base together. The titration is complete only when there is permanent color change in the indicator.

Note that although this procedure states the number of drops of indicator and how many number of titre values, it does not indicate what volume to use in the flask. The typical volume is 25 ml, but students can use any volume as long as they are consistent for each trial.

The practical question for volumetric analysis will always ask students to either determine percentage purity, molecules of crystallization of water, unknown concentration of one of the solutions, or molar mass of one of the solutions. See Common Calculations in Titration Experiments (p. 102) for more explanation on various volumetric analysis calculations.

27.3 Qualitative Analysis

This section contains the following:

- Overview of Qualitative Analysis
- Local Materials in Qualitative Analysis
- The Steps of Qualitative Analysis
- Hazards and Cleanliness
- Sample Practical: Preparation of Copper Carbonate for Qualitative Analysis

27.3.1 Overview of Qualitative Analysis

The salts requiring identification have one cation and one anion. Generally, these are identified separately although often knowing one helps interpret the results of tests for the other. For ordinary level in Tanzania, students are confronted with binary salts made from the following ions:

- Cations: NH_4^+, Ca^{2+}, Fe^{2+}, Fe^{3+}, Cu^{2+}, Zn^{2+}, Pb^{2+}, Na^+
- Anions: CO_3^{2-}, HCO_3^-, NO_3^-, SO_4^{2-}, Cl^-

At present, ordinary level students receive only one salt at a time. The teacher may also make use of qualitative analysis to identify unlabeled salts.

The ions are identified by following a series of ten steps, divided into three stages. These are:

- Preliminary tests: These tests use the solid salt. They are: appearance, action of heat, action of dilute H_2SO_4, action of concentrated H_2SO_4, flame test, and solubility.
- Tests in solution: The compound should be dissolved in water before carrying out these tests. If it is not soluble in water, use dilute acid (ideally HNO_3) to dissolve the compound. The tests in solution involve addition of NaOH and NH_3.
- Confirmatory tests: These tests confirm the conclusions students draw from the previous steps. By the time your students start the confirmatory tests, they should have a good idea which cation and which anion are present. Have students do one confirmatory test for the cation they believe is present, and one for the anion you believe is present. Even if several confirmatory tests are listed, students only need to do one. When identifying an unlabelled container, however, you might be moved to try several, especially if you are new to this process.

27.3.2 Local Materials in Qualitative Analysis

For all low-cost local material substitutes, consult the section on Sources of Laboratory Equipment (p. 208).

- Heat Sources: Motopoa burners cost nothing to make (soda bottle caps) and consume only a small amount of fuel. They give a non-luminous flame ideal for flame tests and still produce enough heat for the other tests.
- Test Tubes: Plastic test tubes suffice.
- Litmus paper: Rosella flowers give very good results.
- Low-cost sources of chemicals (see *Shika Express Chemistry* companion manual).

Share expensive chemicals among many schools. A single container of potassium ferrocyanide, for example, can supply ten or even twenty schools for several years. Schools should consider bartering 10 g of one chemical for 10 g of another. Another alternative is for all of the schools in a district or town to pool money to buy one container of each required imported reagent, and then divide the chemicals evenly.

27.3.3 The Steps of Qualitative Analysis

27.3.3.1 Appearance

Three properties of the salt may be observed directly: colour, texture, and smell.

Colour While most salts are white, salts of transition metals are often colored. Thus colour is an easy way to identify iron and copper cations in salts.

Texture Carbonates and hydrogen carbonates generally form powders although sometimes they can form crystals. Sulphate, nitrates, and chlorides are almost always founds as crystals.

Smell Some ammonium salts smell distinctly like ammonia. Some, however, have no smell. Therefore the smell of ammonia can confirm the presence of ammonium cations, but its absent can not be used to prove the absence of ammonium.

Materials soda bottle caps, table salt, bicarbonate of soda, soda ash (sodium carbonate), copper (II) sulphate*, ammonium sulphate*, locally manufactured iron (II) sulphate*, locally manufactured iron (III) sulphate*, locally manufactured copper (II) carbonate

Preparation

1. Place a small amount of each sample in a different soda bottle cap for observation.

Activity Steps

1. Look at the samples. Describe their colour, texture, and smell. Do not touch or inhale the salts.

Results and Conclusion

- Colour

  White Copper and iron absent

  Blue Copper cation present

  Green Iron (II) or copper present

  Light green Iron (II) present

  Yellow or red-brown Iron (III) present

- Texture

  Powder Carbonate or hydrogen carbonate anion present

  Crystals Sulphate, chloride, or nitrate anion probably present

  Wet crystals Chloride or nitrate anion present

- Smell

  Smell of ammonia Ammonium cation present

  No smell of ammonia Inconclusive – some ammonium compounds have no smell

Clean Up

1. Collect salts for use another day. Do not mix.
2. Wash and return soda bottle caps.

Notes Wet crystals are the result of the salt absorbing water from the atmosphere. Qualitative analysis salts with this property are not locally available. However, caustic soda (sodium hydroxide) has this property, so samples of caustic soda can be used to show the absorption of water from the air and how this changes the appearance of the salt. Note that caustic soda burns skin, blinds in eyes and corrodes metal, so care is required.

27.3.3.2 Action of heat

Many salts thermally decompose when heated. When these salts decompose, they produce gases that may be identified to identify the anion of the salt. After decomposition, many salts also leave a residue that may identify the cation.

Materials soda bottle caps, motopoa, matches, long handled metal spoons, steel wool, sand, beaker*, water, table salt, copper (II) sulphate*, bicarbonate of soda, locally prepared copper (II) carbonate*, soda ash (sodium carbonate), locally prepared zinc carbonate*

Hazards and Safety

- Ammonium nitrate explodes when heated. For this reason, ammonium nitrate should never be used in qualitative analysis when the Action of Heat test is used.
- Test tubes should be pointed away from the student holding them and from other students by holding them at an angle. This will prevent injuries due to splashing chemicals, and will also minimize inhalation of any gases produced. Teach students to never to fill test tubes or any other container more than half.

Preparation

1. Fill a beaker with water.
2. Make a small pile of sand on the table for resting the hot spoon.
3. Place a small amount of each sample in a different soda bottle cap.
4. Add motopoa to another soda bottle cap to use as a burner.

Activity Steps

1. Light the motopoa. Note that the flame will be invisible.
2. Place a very small amount of a sample on the spoon. Generally, the smallest amounts of sample give the best results because they are easier to heat to a hotter temperature.
3. Heat the sample strongly, observing all changes.
4. Place the hot spoon on the sand to cool.
5. Once the spoon has mostly cooled, dip it in the beaker of water to remove the rest of the heat.
6. Use the steel wool to remove all residue from the spoon.
7. Repeat these steps with each sample.

Results and Conclusion

- Gas released

  Brown gas Nitrogen dioxide, nitrates present, confirmed

  Colourless gas with smell of ammonia Ammonia, ammonium present, confirmed

  Colourless gas with no smell Very likely carbon dioxide, especially if the compound decomposes near the start of heating, carbonate or hydrogen carbonate present

  No change Salt probably a chloride, sulphate (very high temperatures are required to decompose many sulphates), or sodium carbonate

- Residue

  No residue Ammonium cation present

  Black residue Copper cation probably present

Red residue when hot, dark when cool Iron cation present

Yellow residue when hot, white when cool Zinc cation present

Red residue when hot, yellow when cool Lead cation present

- Sound

  Cracking sound Sodium chloride or lead nitrate present

Clean Up

1. Thoroughly remove all residues from the spoons.

Notes Sodium carbonate is the only carbonate used in qualitative analysis that does not thermally decompose. Therefore a white powder that does not decompose when heated is probably sodium carbonate.

27.3.3.3 Action of dilute H_2SO_4

Carbonates and hydrogen carbonates react with dilute acid. Sulphates, chlorides and nitrates do not. Therefore reaction with dilute acid is useful test to help identify the anion. Sulphuric acid is used because it is the least expensive.

Materials dilute sulphuric acid*, droppers*, bicarbonate of soda, table salt

Hazards and Safety

- Use only a few drops of acid. These are all that are necessary and using more can be dangerous.

Preparation

1. Place a small amount of each sample in a different soda bottle cap.
2. Fill droppers with 1-2 mL dilute acid.

Activity Steps

1. Add a few drops of acid to each sample. Observe the results.

Results and Conclusion

Bubbles of gas Carbon dioxide produces; carbonate or hydrogen carbonate anion present

No bubbles of gas Carbonate and hydrogen carbonate absent

Clean Up

1. Neutralize spills of dilute sulphuric acid with bicarbonate of soda.
2. Mix the remains from the reactions together so the extra bicarbonate of soda can neutralize the acid used to test table salt. Dilute the resulting mixture with a large amount of water and dispose down a sink, into a waste storage tank, or into a pit latrine.

Notes You can confirm that the gas produced is carbon dioxide by testing to see if it extinguishes a glowing splint. To do this, light a match, use about 0.5 mL of acid (rather than a few drops), and see if the gas released will extinguish the match.

27.3.3.4 Action of concentrated H_2SO_4

Concentrated sulphuric acid can convert chloride anions to hydrogen chloride gas and some nitrates to nitrogen dioxide. Because both of these gases are easy to detect, the addition of concentrated acid is used to distinguish between nitrates, chlorides, and sulphates. The concentrated acid used in this experiment should be about 5 M, similar to battery acid.

Materials battery acid, droppers*, spoons, test tubes*, test tube rack*, test tube holder*, heat source*, hot water bath*, table salt (sodium chloride), gypsum (calcium sulphate)*, ammonium sulphate*, blue litmus paper*, beaker*, water

Hazards and Safety

- Use battery acid or another source of 5 M sulphuric acid for this experiment. Do not use fully concentrated 18 M sulphuric acid directly from either industry or laboratory supply. 18 M is too concentrated and very dangerous to use.
- Concentrate acid reacts violently with carbonates and hydrogen carbonates. The previous test – the addition of dilute acid – will detect carbonates and hydrogen carbonates. If that test is positive, do not test the sample with concentrated sulphuric acid.
- Test tubes should be pointed away from the student holding them and from other students by holding them at an angle. This will prevent injuries due to splashing chemicals, and will also minimize inhalation of any gases produced. Teach students to never to fill test tubes or any other container more than half.

Preparation

1. Place a small amount of each sample in a different soda bottle cap.
2. Add about 1 mL of air to each dropper syringe (no needle!) and then 2 mL of battery acid. Distribute the dropper syringes in the test tube racks so they stand with the outlet pointing down. The goal is to prevent the battery acid from reacting with the rubber plunger.

Activity Steps

1. Light the heat source and start heating the hot water bath. The water in the hot water bath should boil.
2. Use the spoon to add a small amount of a sample to a test tube.
3. Add two drops of battery acid to the sample to make sure there is no violent reaction.
4. Add just enough battery acid to cover the sample. Avoid spilling drops of acid on the inside walls of the test tube.
5. If a brown gas is released, stop at this step.
6. Moisten the blue litmus paper by quickly dipping it in the water of the hot water bath.
7. Place the litmus paper over the mouth of the test tube to receive any gases produces. If the litmus paper changes colour, stop at this step.
8. Hold the test tube in the hot water bath and heat for a while. Stop heating before the acid in the test tube boils. If the litmus paper changes colour before the acid boils, this is a useful result. If the acid boils, fumes from the acid itself will change the colour of the litmus paper – this result is not useful, and acid fumes are dangerous.

Results and Conclusion

Bubbles with a few drops of acid Carbonate or hydrogen carbonate anion present

Brown gas produced Nitrate anion present

Litmus changes to red Hydrogen chloride gas produced; chloride anion present

No effect observed Sulphate anion probably present

Clean Up

1. Fill a large beaker half way with room temperature water. This will be the waste beaker.
2. Pour waste from the test tubes into the waste beaker.
3. Fill each test tube half way with water and add this water to the waste beaker.
4. Return unused battery acid from the droppers to a well-labelled storage container for future use. Immediately fill each dropper (syringe) with water and transfer this water to the waste beaker.
5. Slowly add bicarbonate of soda to the waste beaker until addition no longer causes bubbling. This is to neutralize the acid in the waste.
6. Dilute the resulting mixture with a large amount of water and dispose down a sink, into a waste storage tank, or into a pit latrine.
7. Thoroughly wash all apparatus, including the test tubes and droppers, and return them to the proper places.

27.3.3.5 Flame test

Some metal ions produce a characteristically coloured flame when added to fire.

Materials soda bottle caps, motopoa, metal spoons, beaker*, steel wool, water, table salt (sodium chloride), gypsum (calcium sulphate)*, copper (II) sulphate*, ammonium sulphate*

Preparation

1. Fill a beaker with water.
2. Place a small amount of each sample in a different soda bottle cap.
3. Add motopoa to another soda bottle cap to use as a burner.

Activity Steps

1. Light the motopoa. Note that the flame will be invisible.
2. Place a small amount of sample on the edge of the spoon. For some spoons, it is better to hold the spoon by the wide part and to place the sample on the end of the handle.
3. Hold the sample into the hottest part of the flame, 1-2 cm above the motopoa. If necessary, tilt the spoon so that the sample touches the flame directly. Do not spill the sample into the flame.
4. Dip the hot end of the spoon into the beaker of water to cool it and remove the sample. If necessary, clean the spoon with steel wool.
5. Repeat these steps with each sample.

Results and Conclusion

Blue or green flame Copper present, confirmed

Golden yellow flame Sodium present, confirmed

Brick red flame Calcium present

Bluish white flame Lead present

No flame colour Copper and sodium absent; calcium and lead probably absent; cation is probably ammonia, iron, or zinc

Clean Up

1. Collect unused samples for use another day.
2. Wash and return all apparatus.

27.3.3.6 Solubility

Materials soda bottle caps, two spoons, test tubes*, test tube rack*, hot water bath*, heat source*, distilled (rain) water*, table salt (sodium chloride), soda ash (sodium carbonate)*, gypsum (calcium sulphate)*, powdered coral rock (calcium carbonate)* or locally manufactured calcium carbonate* or locally manufactured copper (II) carbonate*

Preparation

1. Fill a beaker with water.
2. Place a small amount of each sample in a different soda bottle cap.

Activity Steps

1. Light the heat source and start heating the hot water bath. The water in the hot water bath should boil.
2. Decide which spoon will be used for transferring samples and which will be used for stirring.
3. Use the transfer spoon to transfer a very small amount of a sample to a test tube.
4. Add 3-5 mL of distilled water to the test tube.
5. Use the handle of the stirring spoon to thoroughly mix the contents of the test tube.
6. If the sample does not dissolve, heat the test tube in the water bath until the contents of the test tube are almost boiling (small bubbles rise from the bottom). Mix.
7. Repeat these steps with each sample.

Results and Conclusion

Sample dissolves in room temperature water Soluble salt present

Sample dissolves only in hot water Calcium sulphate or lead chloride present

Sample does not dissolve in even hot water Insoluble salt present

Solubility Rules

- All Group I (sodium, potassium, etc) and ammonium salts are soluble (sodium borate is an exception but not relevant to qualitative analysis)
- All nitrates and hydrogen carbonates are soluble
- Most chlorides are soluble (silver and lead chlorides are exceptions, although the latter is soluble in hot water)
- Carbonates of metals outside of Group I are generally insoluble (note that aluminum and iron (III) carbonate do not exist)
- Lead sulphate is insoluble and calcium sulphate is soluble only in hot water. Magnesium sulphate is completely soluble while sulphates of the Group II metals heavier than calcium (strontium and barium) are insoluble. All other sulphates used in qualitative analysis are soluble]

Table of Solubility for Qualitative Analysis

	ammonium	sodium	copper	iron	zinc	calcium	lead
nitrate	O	O	O	O	O	O	O
chloride	O	O	O	O	O	O	Δ
sulphate	O	O	O	O	O	Δ	X
carbonate	O	O	X	X	X	X	X
hydrogen carbonate	O	O	–	–	–	–	–

KEY:

- O = soluble at room temperature
- Δ = soluble only when heated
- X = insoluble in water
- – = salt does not exist

Clean Up

1. Collect all unused (dry) samples for use another day.
2. Unless copper carbonate is used, none of the salts listed in the materials section of this activity are harmful to the environment.
3. Dispose of solutions in a sink, waste tank, or pit latrine.
4. Dispose of solids and liquid wastes with precipitates in a waste tank or pit latrine – never dispose of solids in sinks.
5. If using copper carbonate, collect all waste containing copper carbonate and filter to recover the copper carbonate. Save for use another day.
6. If you do this activity with a lead nitrate or lead chloride, collect these wastes in a separate container. Add dilute sulphuric acid dropwise until no further precipitation is observed. Neutralize with bicarbonate of soda. Dispose this mixture in a waste tank or a pit latrine. The lead sulphate precipitate is highly insoluble will not enter the environment.
7. Wash and return all apparatus.

Notes Calcium carbonate or copper carbonate are recommended qualitative analysis salts to use as examples of insoluble salts. If these are difficult to get, other insoluble compounds may be used for teaching this specific step (but not for other parts of qualitative analysis). Examples of other insoluble compounds include sulphur power, manganese (IV) oxide from batteries, and chokaa (calcium hydroxide, which is only slightly soluble so a significant precipitate will remain).

27.3.3.7 Addition of NaOH solution

Materials soda bottle caps, two spoons, test tubes*, test tube rack*, beakers*, medium droppers (5 mL syringes without needles)*, large droppers (10 mL syringes without needles), caustic soda (sodium hydroxide)*, table salt (sodium chloride), ammonium sulphate*, copper (II) sulphate*, locally manufactured iron (II) sulphate*, locally manufactured iron (III) sulphate*, locally manufactured zinc sulphate*, distilled (rain) water

Preparation

1. Fill a 500 mL water bottle about half way with distilled (rain) water.
2. Add one level tea spoon of caustic soda and then wash the spoon.
3. Label the bottle “1 M sodium hydroxide – corrosive”
4. Place a small amount of each sample in a different soda bottle cap.
5. Pour some of the sodium hydroxide solution into a clean beaker.

6. For each small dropper syringe, suck in about 1 mL of air and then add about 4 mL of sodium hydroxide solution. Distribute the dropper syringes in the test tube racks so they stand with the outlet pointing down. The goal is to prevent the sodium hydroxide from reacting with the rubber plunger.

Activity Steps

1. Decide which spoon will be used for transferring samples and which will be used for stirring.
2. Use the transfer spoon to transfer a very small amount of a sample to a test tube.
3. Use the large dropper syringe to add 3-5 mL of distilled water to the test tube.
4. Use the handle of the stirring spoon to thoroughly mix the contents of the test tube.
5. Use the small dropper to add a few drops of sodium hydroxide solution to the test tube.
6. Observe the colour of any precipitate formed. Also waft the air from the top of the test tube towards your nose to test for smell.
7. If a white precipitate forms, use the stir spoon to transfer a very small quantity of the precipitate to a clean test tube. Add 1-2 mL of sodium hydroxide directly to this sample to see if the precipitate is soluble in excess sodium hydroxide solution.

Results and Conclusion

No precipitate and smell of ammonia Ammonium cation present, confirmed

No precipitate and no smell Sodium cation probably present

Blue precipitate Copper (II) cation present

Green precipitate Iron (II) cation present

Red-brown precipitate Iron (III) cation present

White precipitate not soluble in excess NaOH Calcium cation present

White precipitate soluble in excess NaOH Lead or zinc cation present

Clean Up

1. Save all waste from this experiment, labeling it "basic qualitative analysis waste, no heavy metals" and leave it in an open container. Over time atmospheric carbon dioxide will react with the sodium hydroxide to make less harmful carbonates. After 2-3 days, dispose of the waste in a waste tank or a pit latrine.

27.3.3.8 Addition of NH_3 solution

This test is very similar to the addition of sodium hydroxide solution. The useful difference is that zinc forms a precipitate in ammonia that is soluble in excess ammonia whereas lead forms a precipitate in ammonia that is not soluble in excess ammonia. Therefore, this test is mainly used to separate lead and zinc. Neither lead salts nor ammonia are locally available in Tanzania. Because the process of this test is the same as the addition of NaOH and the results so similar, students can adequately learn about the Addition of NH_3 test by practicing the Addition of NaOH. For the national exam, a small amount of ammonia solution can be obtained.

Note also that the addition of ammonia to a solution of copper (II) will produce a blue precipitate that dissolves in excess ammonia to form a deep blue solution. This is a useful conformation of the presence of copper, but such conformation is generally unnecessary because the flame test for copper is so reliable.

If you have ammonia solution, store it in a well-sealed container to prevent the ammonia from escaping. A good container for this is a well labeled plastic water bottle with a screw on cap.

27.3.3.9 Confirmatory Tests

Every cation and anion has at least one specific test that can be used to prove its presence. Not all of these tests are possible with local materials, but many of them are. The following list shows how to confirm each possible cation and anion.

27.3.3.9.1 Confirmatory Tests for the Cation

Ammonium

- Example salt: ammonium sulphate*
- Procedure: add sodium hydroxide solution and heat in a water bath
- Confirming result: smell of ammonia
- Reagents: NaOH solution as used above

Calcium

- Example salt: calcium sulphate
- Procedure: Two options
  1. flame test
  2. addition of NaOH solution
- Confirming results:
  1. flame test: brick red flame
  2. addition of NaOH: white precipitate insoluble in excess
- Reagents:
  1. none
  2. NaOH solution

Copper

- Example salt: copper sulphate
- Procedure: flame test
- Confirming result: blue/green flame
- Reagents: none

Iron (II)

- Example salt: locally manufactured iron sulphate (keep away from water and air)
- Procedure: addition of sodium hydroxide solution and then transfer of precipitate to the table surface
- Confirming result: green precipitate that oxidizes to brown when exposed to air
- Reagent: sodium hydroxide solution from above

Iron (III)

- Example salt: locally manufactured iron sulphate (oxidized by water and air)
- Procedure: addition of sodium ethanoate solution
- Confirming result: yellow to red solution
- Reagent: slowly add bicarbonate of soda to vinegar; stop adding when further addition does not cause bubbles; label the solution "sodium ethanoate for detection of iron (III)"

Lead

- Example salt: no local sources for safe manufacture, consider purchasing lead nitrate
- Procedure: Three options
  1. flame test
  2. addition of dilute sulphuric acid
  3. addition of potassium iodide solution
- Confirming results:
  1. flame test: blue/white flame
  2. addition of dilute sulphuric acid: white precipitate
  3. addition of KI solution: yellow precipitate that dissolves when heated and reforms when cold
- Reagents:
  1. none but a very hot flame, e.g. Bunsen burner, is required
  2. dilute sulphuric acid used in Step 5 above
  3. obtain pure potassium iodide by evaporating iodine tincture until only white crystals remain; do this outside and do not breathe the fumes; it might also be possible to use the KI solution prepared for electrolysis in the chapter on ionic theory

Sodium

- Example salts: sodium chloride, sodium carbonate, sodium hydrogen carbonate
- Procedure: flame test
- Confirming result: golden yellow flame
- Reagents: none

Zinc

- Example salt: locally manufactured zinc carbonate or zinc sulphate
- Procedure: addition of 0.1 M potassium ferrocyanide solution
- Confirming result: gelatinous gray precipitate
- Reagents: no local source of potassium ferrocyanide – consider collaborating with many schools to share a container; only a very small quantity is required

27.3.3.9.2 Confirmatory Tests for the Anion

Hydrogen carbonate

- Example salt: sodium hydrogen carbonate
- Procedure: add magnesium sulphate solution and then boil in a water bath
- Confirming result: white precipitate forms only after boiling
- Reagent: dissolve Epsom salts (magnesium sulphate)* in distilled (rain) water*

Carbonate

- Example salt: sodium carbonate
- Procedure for soluble salts: addition of magnesium sulphate solution
- Confirming result: white precipitate forming in cold solution
- Reagent: dissolve Epsom salts (magnesium sulphate)* in distilled (rain) water*
- Note that insoluble salts that effervesce with dilute acid are likely carbonates. None of the other anions described here produce gas with dilute acid. Note also that all hydrogen carbonates are soluble.

Chloride

- Example salt: sodium chloride
- Procedure: Three Options
  1. addition of silver nitrate solution
  2. addition of manganese (IV) oxide and concentrated sulphuric acid followed by heating in a water bath
  3. addition of weak acidified potassium permanganate solution followed by heating in a water bath
- Confirming results:
  1. silver nitrate: white precipitate of silver chloride
  2. manganese (IV) oxide: production of chlorine gas that bleaches litmus
  3. acidified permanganate: decolourization of permanganate
- Reagents:
  1. Silver nitrate has no local source but may be shared among many schools as only a very small amount is required.
  2. Manganese dioxide may be purified from used batteries and battery acid is concentrated sulphuric acid. Note that careful purification is required to remove all chlorides from the battery powder. This method is useful because of its low cost, but remember that chlorine gas is poisonous! Students should use very little sample salt in this test.
  3. Prepare a solution of potassium permanganate, dilute with distilled water until the colour is light pink, and then add about 1 percent of the solution's volume in battery acid. Note that this solution will cause lead to precipitate, and will also be decolourized by iron II, so it is not a perfect substitute for silver nitrate. This final option is also not yet recognized by examination boards, i.e. NECTA

Sulphate

- Example salt: copper sulphate, calcium sulphate, iron sulphate
- Procedure: addition of a few drops of a solution of lead nitrate, barium nitrate, or barium chloride
- Confirming result: white precipitate
- Reagents: none of these chemicals have local sources. Because lead nitrate is also an example salt, it is the most useful and the best to buy. The ideal strategy is to share one of these chemicals among many schools. Remember that all are quite toxic.

Notes Emphasize to students that they need to carry out only one confirmatory test for the cation, and one for the anion. If the test gives the expected result, then they can be sure that the ion they have identified is present. If the test does not give the expected result, they have probably made a mistake, and they should revisit the results of their previous tests and choose a different possibility to confirm.

27.3.4 Hazards and Cleanliness

Qualitative analysis practicals are full of hazards, from open flames to concentrated acids. Here are some ways to reduce the risk of accidents:

- Teach students how to use their flame source before the day of the practical.
- Have students hold their test tubes at an angle pointed away from them and other students to prevent splashing chemicals and minimize inhalation of any gases produced.
- Teach students never to fill test tubes or any other container more than halfway in order to minimize spills and boiling over of chemicals during heating.
- Teach students that if they get chemicals on their hands, they should wash them off immediately, without asking for permission first.
- Teach students to tell you immediately when chemicals are spilled. Sometimes they hide chemical spills for fear of punishment. Do not punish them for spills – legitimate accidents happen. Do punish them for unsafe behavior of any kind, even if it does not result in an accident.
- Practicals involving nitrates, chlorides, ammonium compounds, and some sulphates produce harmful gases. Open the lab windows to maximize airflow.
- Make absolutely sure that students clean their tables and glassware before they leave.

27.3.5 Sample Practical Question

The following is a sample practical question from 2012.

Substance **V** is a simple salt which contains one cation and one anion. Carry our the experiments described below. Record carefully your observations and make appropriate inferences and hence identify the anion and cation present in sample **V**.

S/n	**Experiment**	**Observation**	**Inference**
1	Observe the appearance of sample **V**.		
2	Put a little amount of sample **V** in a test tube then add water and shake.		
3	Heat a little amount of **V** in a dry test tube.		
4	To a little sample **V** in a test tube add dilute Hydrochloric acid. Add more of the acid until the test tube is half full. Divide the resulting solution into three portions and add the following:		
	a) To the one portion add NaOH solution drop wise then excess.		
	b) To the second portion add ammonia solution drop wise then in excess.		
	c) To the third portion add ammonium oxalate solution.		
5	Perform flame test.		

Conclusion

(i) The cation in sample **V** is ____ .

(ii) The anion in sample **V** is ____ .

(iii) The chemical formula of **V** is ____ .

(iv) The name of compound **V** is ____ . **(20 marks)**

Discussion

This particular example was to identify calcium carbonate; however, the above procedure follows the same format commonly used for other unknown salts. The only differences may be the specific solutions used in some of the steps.

At times, the procedure may not be explicit or indicated whatsoever and the student is required to write the detailed procedure in addition to the observations and inferences.

Emphasize to students that in addition to the qualitative analysis procedure they have to do only one confirmatory test for cations and one for anions.

27.4 Chemical Kinetics and Equilibrium

27.4.1 Theory

Compared to the other two NECTA chemistry practicals - Acid/Base Titration (i.e. Volumetric Analysis) and Qualitative Analysis - Chemical Kinetics has few alternative chemicals that can be used.

The chemical reaction in the NECTA exam is a precipitation of sulphur.

$$Na_2S_2O_{3(aq)} + 2HCl_{(aq)} \longrightarrow 2NaCl_{(aq)} + H_2O_{(l)} + SO_{2(g)} + S_{(s)}$$

The preparation and procedure for Chemical Kinetics is very simple.

27.4.2 Preparation

1. Prepare solution of Sodium Thiosulphate by placing appropriate mass of crystals (in accordance with desired concentration) in a water bottle and shake vigorously until completely dissolved.

2. Prepare acid solution using hydrochloric acid (preferable) or sulphuric acid. (Remember for calculations that sulphuric acid will have two H^+ ions to hydrochloric acid's one.)

3. Take a white piece of paper and draw a large black X on it

27.4.3 Procedure

The procedure can change depending on whether the variable of the reaction rate is temperature or concentration. The procedure outlined here will be focused on concentration. The sample NECTA question below is based on having a temperature variable.

1. Place in a beaker or glass the amount of sodium thiosulphate solution and clean water prescribed in the table.

2. Ready the stopwatch and pour in the prescribed amount of hydrochloric acid, starting the stopwatch as you do so.

3. Gently swirl the glass and place it over the paper with an X on it.

4. When the solution has precipitated enough sulphur to the point where the X is no longer visible, stop the stopwatch. Record the time.

5. The resulting solution is already neutralized, and can be merely diluted and disposed of as is. (Unless you wish to obtain the salt/sulphur mixture by evaporation for another day.)
6. Thoroughly rinse the glass and repeat, changing the amounts of sodium thiosulphate and water as prescribed by the problem.
7. Graph the results to calculated the desired variables

This practical is consistent and easy to practice, but it requires sodium thiosulphate which can be expensive and hard to get a hold of. (As mentioned, the hydrochloric acid can be replaced with sulphuric (battery) acid.) An alternative reaction to demonstrate chemical kinetics (that can be performed very cheaply) is the iodization of acetone, seen in the reaction below.

$$CH_3COCH_{3(aq)} + I_2 \longrightarrow CH_3COCH_2I_{(aq)} + H^+_{(aq)} + I^-_{(aq)}$$

This reaction starts as a dark opaque solution and eventually proceeds to a colorless, transparent solution. It requires an acidic environment to occur (both hydrochloric or sulphuric acid suffice). The amount of time it takes to reach the point of colorlessness varies depending on the concentration of the acetone. It is easy to demonstrate the relationship between concentration and rate of reaction. (By varying the temperature or amount of acid catalyst the reaction visibly proceeds at differing rates.)

27.4.4 Performing the Practical

Materials

7 beakers, 3 syringes, stopwatch

Chemicals

Nail polish remover, iodine tincture, sulphuric (battery) acid, water

Preparation

1. Prepare an iodine solution using iodine tincture. Solutions purchased in the local drugstores are often 0.2 M Iodine (with several other chemicals as well). Create a 0.02 M solution by adding 9 parts water to one part tincture. Put this solution in the first beaker.
2. Prepare an acidic acetone solution by mixing one part nail polish remover to one part 1 M Sulphuric acid solution. Put this solution in the second beaker. This is roughly a 5 M solution of acetone.
3. In the third beaker place clean water.

Procedure

1. Pour 8 ml of the acetone solution in a clean beaker. Clear a stopwatch and add 8 ml of iodine to the beaker, swirl the solution, and stop the watch when the solution becomes colorless.
2. Record the value and then perform the experiment again. This time start with 6 ml of acetone solution and 2 ml of clean water in a clean beaker. Clear a stopwatch and add 8 ml of iodine to the beaker, swirl the solution, and stop the watch when the solution becomes colorless.
3. Repeat the above steps with 4 ml of acetone solution, 4 ml of water, and 8 ml of iodine, and then finally one more time with 2 ml of acetone solution, 6 ml of water and 8 ml of iodine.

Test	Volume of Iodine Solution (mL)	Volume of Acetone Solution (mL)	Volume of Water (mL)	Molarity of Acetone Solution ($^{mol}/_{L}$)	Time for color to disappear (s)	Reciprocal of time ($^{1}/_{s}$)
1	8	8	0	5 M		
2	8	6	2	3.75 M		
3	8	4	4	2.5 M		
4	8	2	6	1.25 M		

This data can then be used to plot a graph of concentration of acetone against the rate of reaction.

Notes

- Patience is required for the tests using lower concentrations as they can take over 4 minute to complete.
- Concentrations may vary depending on where the tincture and remover are purchased.
- The endpoint of this reaction can be somewhat ambiguous depending on the color of the nail polish remover. Criteria for determining the endpoint may vary.
- It should be noted that if the nail polish remover is already a specific color it will affect the final color of the solution. Some solutions may never become fully colorless.
- The reaction used in the NECTA exams goes from transparent to opaque while this alternative goes from opaque to transparent. Make sure students understand this difference.
- The reaction used in NECTA exams is a neutralization reaction so there is little that needs to be done to process the waste. This alternative reaction is very acidic when finished so be prepared to neutralize it before disposal.

27.4.5 Sample Practical Question

The following is a sample practical question from 2012.

Your are provided with the following materials:

ZO: A solution of 0.13 M $Na_2S_2O_3$ (sodium thiosulphate);

UU: A solution of 2 M HCl;

Thermometer;

Heat source/burner;

Stopwatch.

Procedure:

(i) Place 500 cm^3 beaker, which is half-filled with water, on the heat source as a water bath.

(ii) Measure 10 cm^3 of **ZO** and 10 cm^3 of **UU** into two separate test tubes.

(iii) Put the two test tubes containing **ZO** and **UU** solutions into a water bath.

(iv) When the solutions attain a temperature of 60^oC, remove the test tubes from the water bath and pour both solutions into 100 cm^3 empty beaker and immediately start the stop watch.

(v) Place the beaker with the contents on top of a piece of paper marked **X**.

(vi) Note the time taken for the mark **X** to disappear.

(vii) Repeat step (i) to (vi) at temperature 70^oC, 80^oC and 90^oC.

(viii) Record your results as in Table 1.

Table 1

Experiment	**Temperature**	**Time (s)**
1	60^oC	
2	70^oC	
3	80^oC	
4	90^oC	

Questions:

1. Write a balanced chemical equation for reaction between **UU** and **ZO**.
2. What is the product which causes the solution to cloud the letter **X**?
3. Plot a graph of temperature against time (s).
4. What conclusion can you draw from you graph? **(20 marks)**

Discussion

This particular example was to investigate how temperature affects the rate of a chemical reaction. Other experiments for chemical kinetics involve concentration, surface area, and a catalyst; however, the most common problem statement is how concentration affects the rate of reaction.
For all scenarios, make sure to tell students to start the stopwatch *immediately* after combining the reactants.

Chapter 28

Physics Practicals

28.1 Introduction to Physics Practicals

28.1.1 Format

The format of the Physics practical exam was revised in 2011 to keep up with the 2007 updated syllabus. As such, there will be no further Alternative to Practical exams, pending approval from the Ministry of Education. Prior to 2011, there were 3 questions in the Physics practical, and students chose 2 questions to do.

As of now, the Physics practical has 2 questions and students must answer **both**. Question 1 comes from Mechanics, and Question 2 can come from Heat, Light, Waves or Electricity. Each question is worth 25 marks, and students have $2\frac{1}{2}$ hours to complete the exam.

Physics 1 Theory Format

The theory portion of the Physics exam comprises 100 marks, while the practical carries 50 marks. A student's final grade for Physics is thus found by taking her total marks from both exams out of 150.

The theory exam for Physics contains 3 sections. Section A has 3 questions and is worth 30 marks - Question 1 is 10 multiple choice, Question 2 is 10 matching, and Question 3 is 10 fill-in-the-blank. Section B has 6 long answer questions and is worth 60 marks. Section C has 2 questions regarding the use of apparatus and simple technological appliances in everyday life, though students must answer **only 1** of these questions. It is worth 10 marks.

Note This information is current as of the time of publication of this manual. Updated information may be obtained by contacting the Ministry of Education.

28.1.2 Notes for Teachers

NECTA Advance Instructions

There is only one set of advance instructions for the Physics practical. Advance instructions are given to teachers at least one month before the date of the exam. Unlike Biology and Chemistry, **there are no 24 hour advance instructions given for Physics**. The advance instructions will state exactly how many and which apparatus candidates should be provided with for each practical question.

Words of Advice

The Physics practical is different from the Chemistry and Biology practicals in that the exams feature a greater variety of questions. That means we need to teach it all, even if the teacher before you never found his or her way into the classroom and you realize at the end of Form Four that the students still have not studied the Form Two syllabus. If you are teaching all forms, do not wait to start practicals until later forms; always do a practical when the corresponding topic comes up. In addition, train the students well in the general principles of collecting data, graphing data, and writing up experimental points. These skills are required in every physics practical, and carry most of the points.

The practical section of the exam is a third of a student's total score, and fully half of that is graphing and labeling experimental data. Though the practical is varied, a student does not necessarily need a deep understanding of the concept in question. If they are familiar with the apparatus and the process of drawing and interpreting a graph, the practical should be quite simple. Whenever possible, allow the students to play and experiment with the apparatus, whether it is a metre bridge, mirror, pendulum, etc. If they have done each of these experiments several times, they will be confident in their ability.

The more familiar your students are with these techniques, the better they will do. Perform these practicals as often as possible: when the topic comes up, when preparing for the mock and NECTA exams, and any time you can get them to come in for an evening session or a weekend. They will make many, many mistakes the first couple of times through but that is exactly what you want as they will learn from their mistakes and remember them.

28.1.3 Common Practicals

Mathematics a brief overview of some of the mathematical and graphing skills required to perform many of the common physics practicals

Mechanics

Hooke's Law (Form 1) use a spring and various masses to determine either the spring constant or the value of an unknown mass graphically

Simple Pendulum (Form 2) find the acceleration due to gravity using a pendulum and stopwatch

Principle of Moments (Form 2) verify the Principle of Moments using masses and a ruler on a knife edge, or calculate the mass of a metre rule

Light

Plane Mirror (Reflection) (Form 1) generally 3 varieties: find image distance, verify the Laws of Reflection, or find the number of images produced by two plane mirrors placed at different relative angles

Rectangular Prism (Refraction) (Form 3) find the refractive index or critical angle for a glass block by varying the angle of incidence and measuring the corresponding angles of refraction

Electricity

Potentiometers find the drop in potential along a length of resistance wire

Metre Bridges determine the value of an unknown resistor using a known resistor and a galvanometer to find a point of equal potential along a resistance wire

Ohm's Law (Form 2) verify Ohm's Law or determine the internal resistance of a cell

Note These are the most common practicals, but they are not necessarily the only practicals that can occur on a NECTA exam. Physics practical questions can come from a variety of topics which may not yet have been used in older past papers. Be sure to regularly check the most recent Physics Past Papers to get a good idea of the types of questions to expect.

28.1.4 Recent Practicals

Given below is an attempt to characterize the Physics practical questions from recent years according to category and topic. Each question is given with its corresponding topic on top and the objective, or what is to be solved for, on bottom. This information can be used to try and find trends in what exam writers like to test students on, and what topics are most likely to occur on an exam. However, nothing should be assumed to be a guarantee, and students should be well-prepared in all practicals so that they can take on any question they may face on the NECTA.

Note a few things from the table below:

- Beginning in 2011, the format of the exam changed to consist of only 2 problems, one of which must come from Mechanics.
- The exam committee has a tendency to repeat some problems. For example, the Mechanics questions from 2004 and 2011 are nearly identical, as are the Light questions from 2004 and 2008.

Year	Mechanics	Light	Electricity
2013	Principle of Moments *verify*	Plane Mirror *object, image dist.*	—
2012	Density *verify* $\rho = \frac{m}{V}$	Plane Mirror *number of images*	—
2011	Principle of Moments *mass of "AA" battery*	—	Ohm's Law *e.m.f., int. resistance*
2010	Elasticity *unknown mass, k*	Refraction *refr. index of water*	Ohm's Law *ρ of a wire*
2009	Simple Pendulum *relate* $l \to T$	Refraction *refr. index of glass*	Ohm's Law *verify* $V = IR$
2008	Elasticity *verify* $F = kx$*, find k*	Refraction $\frac{d\cos r}{\sin(i-r)}$	Ohm's Law *verify* $V = IR$
2007	Elasticity *unknown mass, k*	Refraction *refr. index of glass*	Potentiometer ΔV *along wire*
2006	Elasticity *unknown mass,* $F = kx$	Plane Mirror *object, image dist.*	Metre Bridge *unknown R*
2005	Elasticity *unknown mass, k*	Refraction *critical angle*	Ohm's Law *e.m.f., int. resistance*
2004	Principle of Moments *mass of "AA" battery*	Refraction $\frac{d\cos r}{\sin(i-r)}$	Potentiometer *ρ of wire, int. res.*

28.2 Mathematics

No physics experiment is complete without a healthy dose of graphing and formulas. As math is typically the worst subject for most students, it is often upon the physics teacher to drive home the understanding of how to draw and interpret graphs, as well as how to apply formulas to those graphs. It comes down to a few simple things: correctly setting up a graph (scales, units, labels, etc.), plotting points from a table of data, and fitting a best-fit line. After this, the students need to find the slope of this line and its y-intercept.

28.2.1 Graphing

Most of the graphs will be linear, meaning the slope is constant, so we apply the standard equation for a line

$$y = mx + c$$

where y represents the vertical axis, x represents the horizontal axis, m is the slope of the line and c is the point on the vertical axis where the line crosses. Almost every practical will make use of this equation, so be sure that your students understand it inside and out. It often helps to do repetitive practice using just the mathematical symbols before introducing physics concepts. Note that very rarely non-linear graphs appear, e.g. cooling curves in heat practicals. In this case students will not have to find a mathematical relationship, just describe and explain the trend in the data.

28.2.2 Formulas

Now comes the physics; all the practicals will involve an equation that can be rewritten in this linear form. The exam question will dictate which variable is independent (x) and which is dependent (y). It is up to the student to simply rewrite the formula with each variable on its respective side and then infer what m, the slope, and c, the y-intercept, must be.

The most common formulas used for mechanics, light and electricity are as follows:

Hooke's Law	$F = ke$ or $F = ke - B$
Period of a Pendulum	$T = 2\pi\sqrt{\frac{l}{g}}$
Principle of Moments	$(F \times d)_{\text{clockwise}} = (F \times d)_{\text{anticlockwise}}$
Snell's Law	$n_1 \times \sin i = n_2 \times \sin r$
Ohm's Law	$V = IR$ or $V = I(R + r)$
Resistance of a Wire	$R = \frac{\rho l}{A}$
Wheatstone Bridge	$\frac{R_1}{L_1} = \frac{R_2}{L_2}$

In each case, one quantity will be changed (independent) and another will be measured (dependent) over the course of the experiment. The student will therefore need to rearrange the equation so that the dependent variable is the subject in the form

$$y = mx + c$$

Example Problem - Snell's Law

As an example, in an experiment to measure the index of refraction of a glass block, a student will be measuring angles of incidence and refraction. This means we need to use Snell's Law

$$n_1 \times \sin i = n_2 \times \sin r$$

The question will typically ask students to plot a graph of their measurements, with $\sin i$ on the y-axis and $\sin r$ on the x-axis, or vice-versa. To rewrite Snell's law in the form of $y = mx + c$ is simple; we get

$$\sin i = \sin r \times \frac{n_2}{n_1}$$

and we can see that the value corresponding to m is the ratio $\frac{n_1}{n_2}$, and c must be zero.

Since we are trying to find n_2 (the refractive index of glass), and we know n_1 is 1.0, we simply measure the slope and solve to find n_2.

The approach itself is relatively simple, but students will need lots of practice with graphing, rewriting equations in linear form, and determining what corresponds to m and c in each case. The same approach is used to find the quantities in each of the equations above.

A complete list of each equation in its most commonly found $y = mx + c$ form, along with its corresponding dependent variable y, independent variable x, slope m and y-intercept c, is given below. Note that these equations are not always used in the given form on practicals. It is up to the student to determine how each equation must be analyzed during an exam. The variables used and methods of graphing change from year to year, and so the following table should by no means be memorized or assumed to be applicable for a given problem.

Name of Law	**Equation**	$\boldsymbol{y = mx + c}$	$\boldsymbol{y}$	$\boldsymbol{x}$	$\boldsymbol{m}$	$\boldsymbol{c}$
Hooke's Law	$F = ke$	$F = ke - B$	F	e	k	$-B$
Period of a Pendulum	$T = 2\pi\sqrt{\frac{l}{g}}$	$T^2 = 4\pi\frac{l}{g}$	T^2	l	$\frac{4\pi}{g}$	0
Principle of Moments	$(F \times d)_{\text{cw}} = (F \times d)_{\text{acw}}$	$a = \left(\frac{m_2+x}{m_1}\right) b$	a	b	$\left(\frac{m_2+x}{m_1}\right)$	0
Snell's Law	$n_1 \times \sin i = n_2 \times \sin r$	$\sin i = \frac{n_2}{n_1} \sin r$	$\sin i$	$\sin r$	$\frac{n_2}{n_1}$	0
Ohm's Law	$V = IR$ $V = I(R + r)$	$V = IR$ $R = \frac{E}{I} - r$	V R	I $\frac{1}{I}$	R E	0 $-r$
Resistance of a Wire	$R = \frac{\rho l}{A}$	$V = \frac{I\rho l}{A}$	V V	l I	$\frac{I\rho}{A}$ $\frac{\rho l}{A}$	0 0
Wheatstone Bridge	$\frac{R_1}{L_1} = \frac{R_2}{L_2}$	$R_2 = \frac{100R_1}{L_1} - R_1$	R_2	$\frac{100}{L_1}$	R_1	$-R_1$

28.2.3 Units

Be sure to always stress the importance of units when performing practicals, especially in graphing. Using the wrong units can lead to inaccurate interpretations of graphical data. For example, when using Hooke's Law, a spring constant is typically given in units of $^{\text{N}}/_{\text{m}}$. But if a student is graphing mass (in

g) against extension (in cm), then the slope will be in units of $^{\mathrm{g}}/_{\mathrm{cm}}$. In order to get units of $^{N}/_{\mathrm{m}}$, one would have to first convert the slope into units of $^{\mathrm{kg}}/_{\mathrm{m}}$ and then multiply by the acceleration of gravity ($g = 10\ ^{N}/_{\mathrm{kg}}$). A problem may or may not ask for specific units in its answers, but regardless, students should always be conscious of what units they are using when making calculations.

Thinking of units can also help students to understand a problem they are struggling with. If they can remember that slope is *change in y over change in x*, then they may be able to deduce the meaning of the slope of a graph by looking at the units of the y and x axes.

The most important part of any experiment, though, is following directions. If a student can follow directions, which usually are clearly provided by the exam, and can graph data, they can easily perform any experiment. If anything, the practical exam is a test in a student's ability to follow instructions.

28.3 Mechanics

The mechanics section is mandatory on every exam and typically falls into three categories: Hooke's Law, Simple Pendulum, and Principle of Moments. However, other topics are possible, as evidenced by a question on Archimedes' Principle on the 2012 exam. These experiments use the following materials:

Metre Rules If unavailable, go to a local fundi to mass produce them.

Masses See Sources of Laboratory Equipment for local varieties.

Springs Can be bought at lab stores, or can use substitutes such as rubber bands or strips of elastic from a tailor.

Retort Stands May be available or can be made using a filled 1.5 L water bottle with a bamboo stick taped at the top and extending to one side.

Eureka Can Cut off the bottom of a 500 mL water bottle and cut a slit at the top that can be folded downward to make a curved spout.

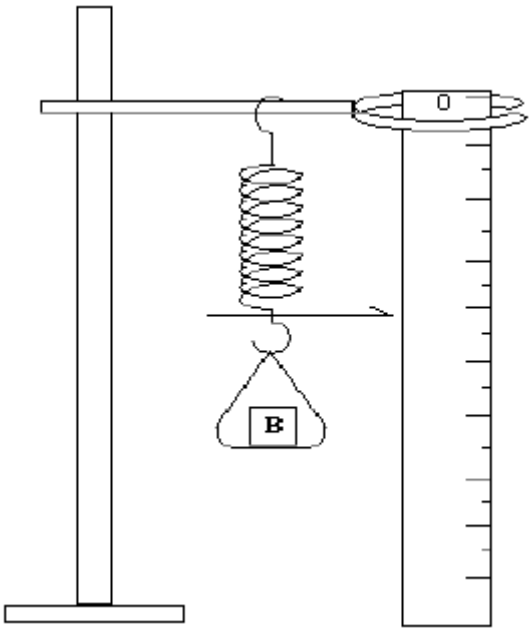
28.3.1 Hooke's Law

This is the most common practical, usually involving a spring but sometimes a rubber band or piece of string. This experiment can be tricky simply because NECTA likes to switch it up every year; try to give your students as much practice with different variations. It is likely that NECTA will require a spring of known spring constant, and you will need known masses. Either can be bought at a laboratory supply store in town, but it is possible to make your own. The practical is simple to perform, but there are some common mistakes: be sure the students understand that the extension is the change in length, not the ultimate length shown on the ruler. Also, do not confuse mass and weight, as is common.

An example question from the 2007 NECTA is shown below. After reviewing the topic with your students, let them try this on their own. You will need to repeat it several times before they are comfortable, using different springs and masses each time.

Sample Practical Question

The aim of this experiment is to determine the mass of a given object "B", and the constant of the spring provided.



(i) Set up the apparatus as shown in the figure with zero mark of the metre-rule at the top of the rule and record the scale reading by the pointer, S_0.

(ii) Place the object "B" and standard weight (mass) W equal to 20 g in the pan and record the new pointer reading S_1. Calculate the extension, $e = S_1 - S_0$ in cm.

(iii) Repeat the procedure in (ii) above with W = 40 g, 60 g, 80 g and 100 g.

(a) Record your results in tabular form as shown below:
Table of Results:

$S_0 =$			
Mass (kg)	Force, F (N)	Pointer reading S_1 (cm)	Extension $= S_1 - S_0$ (cm)
0			
0.02			
0.04			
0.06			
0.08			
0.10			

(b) Plot graph of Force F (vertical axis) against extension e (horizontal axis).

(c) Use your graph to evaluate

(i) mass of B

(ii) spring constant, K, given that force, extension, constant and weight of B are related as follows: F = Ke - B

Discussion

This practical has two parts: the first is to find the spring constant k, the second is to find the mass of an unknown object B. By looking at the equation above, we can see that F is the dependent variable, e is the independent variable, K is the slope and $-B$ is the intercept. When the graph is drawn, K and B can be found easily. Note that the intercept on the graph will be negative.

The procedure is simply to start from a certain point on the metre rule (it does not need to be a specific number) and to add masses one at a time, measuring the distance from your starting point to the new position. This distance is called the extension, e. Be sure that you are not simply reading the metre rule, but are measuring the distance from the starting point.

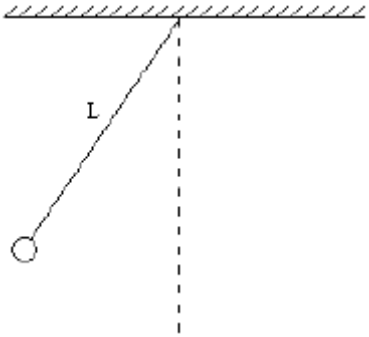
28.3.2 Simple Pendulum

With some practice, this experiment should be simple for anyone to perform. The trick comes with the math and graphing (again, an example is shown below). The materials can all be local (string, stones, ruler) except for the stopwatches (for which you should consult the materials section).

The practical usually has one objective: to find the acceleration due to gravity, g. We know that the mass of a pendulum and its angle of deflection (for small angles) do not affect its period. Therefore we vary only the length L of the pendulum and measure its period, as shown in the following example question.

Sample Practical Question

The aim of this experiment is to determine the magnitude of the acceleration due to gravity, g. Proceed as follows:



1. Make a simple pendulum by suspending a weight on a string 10 cm long from a retort stand.
2. Allow the pendulum to swing for twenty oscillations, using a stopwatch to record the time. Repeat this procedure for pendulum lengths of 20 cm, 30 cm, 40 cm, and 50 cm.
3. Record your results in tabular form as shown below

Pendulum Length l (m)	Time for 20 oscillations (s)	Period $T = \frac{t}{20}$ (s)	T^2 (s^2)
0.1			
0.2			
0.3			
0.4			
0.5			

4. Plot a graph of T^2 (vertical axis) against Pendulum Length (horizontal axis).
5. Calculate the slope of the graph.
6. Use the slope to calculate the value of g.
7. What are possible sources of error in this experiment?

Discussion

The period of a pendulum can be calculated using

$$T = 2\pi\sqrt{\frac{l}{g}}$$

where l is the length of the pendulum, T is the period and g is the acceleration due to gravity. By squaring both sides, we get a much easier equation to graph:

$$T^2 = 4\pi\frac{l}{g}$$

In this equation we see that T^2 is the dependent variable (y-axis) and l is the independent variable (x-axis), so the slope must be

$$\text{slope} = \frac{4\pi}{g}$$

When the graph is complete, the value of g can be calculated easily.

Many students are confused by the difference between the time for many oscillations and the period, which is the time for one oscillation. Be sure that they can change between the two easily.

Note that pendulum practicals do not always require students to find g. Sometimes they are just required to find the relationship between l and t. Again, it is essential that students read and understand the examination question, rather than memorize past solutions, and that they have lots of practice in collecting, organizing, and graphing data from a variety of experiments.

28.3.3 Principle of Moments

This experiment is used to verify the Principle of Moments, or equilibrium, by balancing a meter rule on a knife-edge with masses at various distances. For this experiment, students need a solid understanding of the Center of Gravity, the Moment of a force, and equilibrium. Questions can range from finding the mass of an object to asking for the mass of the metre rule. They are all variations on the same practical: using the condition of equilibrium to find mass.

The following example is from the 2011 NECTA exam and asks students to find the mass of a battery using the Principle of Moments. Following is a brief explanation of the alternative practical of finding the mass of a metre rule.

Finding the Mass of an Object

Sample Practical Question

The aim of this experiment is to determine the mass of a given dry cell size "AA". Proceed as follows:

(a) Locate and note the centre of gravity C of the metre rule by balancing it on the knife edge.

(b) Suspend the 50 g mass at length 'a' cm on one side of the metre rule and the 20 g mass together with the dry cell at length 'b' cm on the other side of the metre rule. Fix the 50 g mass at length 30 cm from the fulcrum and adjust the position of the 20 g mass together with the dry cell until the metre rule balances horizontally. Read and record the values of a and b as a_0 and b_0 respectively.

(c) Draw the diagram for this experiment.

(d) By fixing $a = 5$ cm from fulcrum C, find its corresponding length b.

(e) Repeat the procedure in (d) above for $a = 10$ cm, 15 cm, 20 cm and 25 cm. Tabulate your results.

(f) Draw a graph of 'a' against 'b' and calculate its slope G.

(g) Calculate X from the equation $50 = \frac{b_0}{a_0}(20 + X)$.

(h) Comment on the value of $\frac{b_0}{a_0}$.

(i) Sate the principle governing this experiment.

Discussion

This practical utilizes the Principle of Moments to find the mass of a "AA" battery. Initially, a known mass of 50 g is balanced with the (battery + 20 g mass) system. Note that 'a' and 'b' are measured *from the fulcrum* and so students should be careful not to just read the cm mark on the ruler where each object is suspended.

Also note that students are required to actually find the centre of gravity C of the ruler rather than assuming it to be the 50 cm mark. This measured value of C is to be used as the starting point for all future measurements of a and b.

In part (g), students should recognize the equation $50 = \frac{b_0}{a_0}(20 + X)$ as coming from the Principle of Moments. Starting with

$$F_{\text{clockwise}} \times d_{\text{clockwise}} = F_{\text{anticlockwise}} \times d_{\text{anticlockwise}}$$

we get

$$(mg)_{\text{clockwise}} \times d_{\text{clockwise}} = (mg)_{\text{anticlockwise}} \times d_{\text{anticlockwise}}$$

or

$$(50\text{g})(g) \times a_0 \text{ cm} = (20\text{g} + X\text{g})(g) \times b_0 \text{ cm}$$

Canceling g and dividing by a_0 reveals

$$50 = \frac{b_0}{a_0}(20 + X)$$

where $\frac{b_0}{a_0}$ is the ratio of the lever arm distances for the two weights being used. If the mass of the battery X is less than 30 g, this ratio should be greater than 1, but if the mass is greater than 30 g, the ratio should be less than 1.

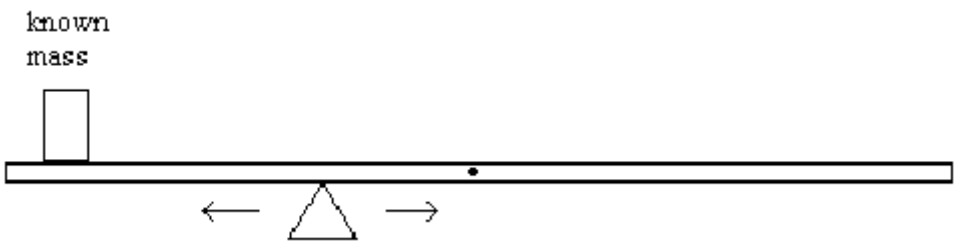
Finding the Mass of a Metre Rule

This question is less frequently seen on NECTA exams as compared to finding an unknown mass. However, it utilizes the same principles of equilibrium and balancing moments, and therefore is a useful alternative practical to ensure that students understand the concept rather than memorizing solutions to one version of the problem.

The mass of a uniform solid object, like a metre rule, is assumed to be at the center of the object. In the case of the metre rule, we can say that the center of mass is at the 50 cm mark, directly in the center. If we want it to be in equilibrium, the moments on either side of a pivot must be equal, or

$$\text{Clockwise moment} = \text{Anticlockwise moment}$$

To find the mass of the metre rule itself, we begin by placing a known mass at one point on the metre rule. We then move the pivot to one side or another until the metre rule is perfectly balanced in equilibrium. As shown in the diagram below, the pivot will not be at the 50 cm mark.



If the metre rule is in equilibrium, we know that the moments must be equal, or that

$$F_{\text{clockwise}} \times d_{\text{clockwise}} = F_{\text{anticlockwise}} \times d_{\text{anticlockwise}}$$

In this case, the anticlockwise force is the weight of the object, and the distance is that from the pivot to the object. The clockwise force is the weight of the metre rule, and the distance is that from the 50 cm mark (center of mass) to the pivot. Therefore our equation is:

$$W_{\text{rule}} \times d_{\text{rule}} = W_{\text{object}} \times d_{\text{object}}$$

Because the weight of the object is known, and the two distances can be measured, we can easily calculate the mass and therefore the weight of the metre rule:

$$W_{\text{rule}} = \frac{W_{\text{object}} \times d_{\text{object}}}{d_{\text{rule}}}$$

From this we can calculate mass of the metre rule using $F = mg$.

28.4 Light

The light practical typically involves plane mirrors or glass blocks (rectangular prisms). Presumably you will have already done these practicals with the students in Form three (refraction) and Form 1 (plane mirrors), but a little practice will make the theory and execution clear, especially if they can work in groups. The materials you will need are as follows:

Cork Board Use cardboard for this, about 0.5 to 0.75 cm thick.

Optical pins Use sewing pins or syringe needles. If using syringe needs, be sure to crimp the ends so students do not prick themselves.

Protractors These are cheap and students are supposed to have them anyway. Small ones come in local mathematical sets.

Glass Block / Rectangular Prism A simple rectangular piece of 6 mm glass, about 8 cm by 12 cm, will work.

Plane Mirror You can buy mirror glass in town in small sections for 200/= or less; it should be available in villages through the local craftsmen if they work on windows. Alternately, you can smoke one side of a piece of glass to make the other side like a mirror.

28.4.1 Plane Mirror (Reflection)

These are not as common as the rectangular prism, but they come in a variety of questions:

- Placing pins in front of a mirror at different distances and finding the distance of the image.
- Verifying the Law of Reflection at plane mirrors.
- Placing two mirrors at different relative angles to find the number of images produced.

These are not overly complicated, but you should definitely practice with your students creating images in mirrors – they are not as accustomed to playing with mirrors as you might be.

Given below is an example practical from the 2006 NECTA exam which asks students to find a relationship between object distance and image distance in a plane mirror.

Sample Practical Question

Set up the experiment as shown in the diagram below using plane mirror, soft board, three pins and a white sheet of paper.

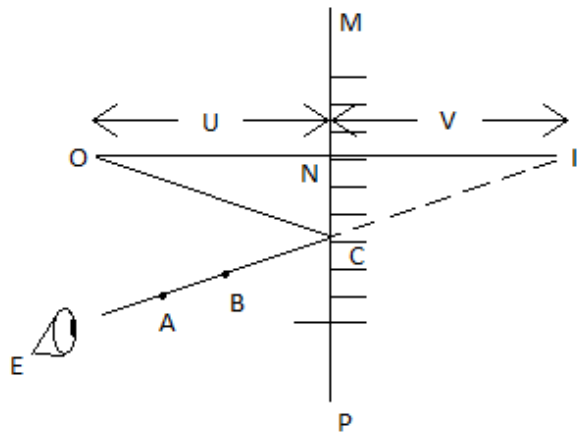


Fig. 2

Fix a white sheet of paper on the soft board. Draw a line across the width at about the middle of the white sheet (MP). Draw line ONI perpendicular to MP.

Fix optical pin O to make ON = U = 3 cm. By using plasticine or otherwise, fix plane mirror along portion of MP with O in front of the mirror. With convenient position of eye, E, look into the mirror and fix optical pins A and B to be in line with image, I, of pin O.

Measure and record NI = V. Repeat procedure for U = 6 cm, 9 cm and 12 cm.

(a) Tabulate your results as follows:

U (cm)	3	6	9	12
V (cm)				

(b) Plot graph of U against V.

(c) Calculate slope, m of the graph to the nearest whole number.

(d) State relationship between U and V.

(e) Write equation connecting U and V using numerical value of m with symbols U and V.

(f) From your equation give position of the image when object is touching the face of the mirror.

Discussion

For a plane mirror, object distance and image distance are equal. That is, U and V should be approximately equal values for this practical. Note that students should extend line ONI far behind the mirror since they don't know where exactly the image I is going to be. The location of I is found at the intersection of the extended line ONI and the extended line AB connecting the two optical pins.

From the graph, the slope should be found to be 1 after rounding to the nearest whole number. From this, we can see that U = V. When the object is touching the face of the mirror, the object distance U is 0, and so the image distance V will also be 0.

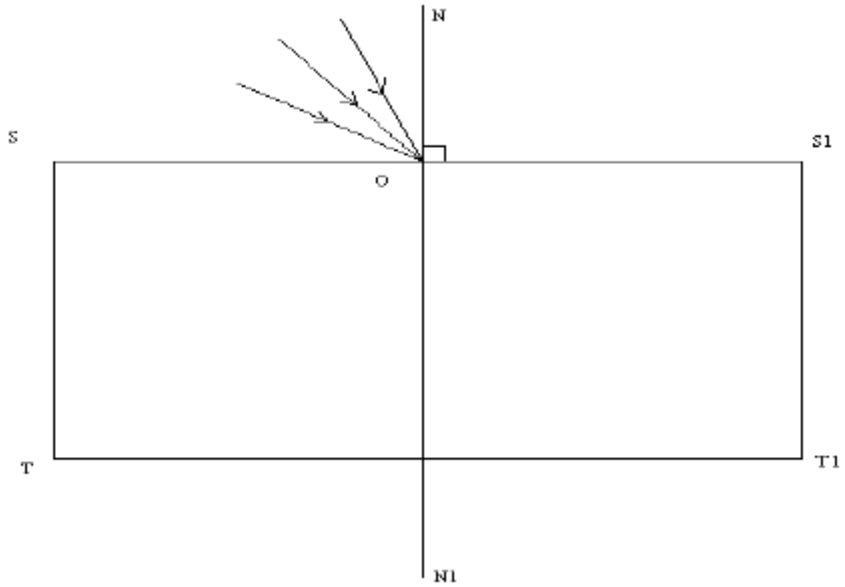
28.4.2 Rectangular Prism (Refraction)

Students will be asked to find the refractive index and/or critical angle of the glass block by varying the angle of incidence i and measuring the corresponding angles of refraction r as described in the Mathematics section earlier. They will do this by placing two pins in front of the prism, which together form a 'ray' (the light ray), and then placing two more pins on the other side of the prism so that, when observed through the prism from either side, the four pins line up exactly. By drawing the lines that the pins make on the paper, the refracted ray inside the prism can be easily traced, and the refracted angle measured. An example question from the 2007 NECTA is given below.

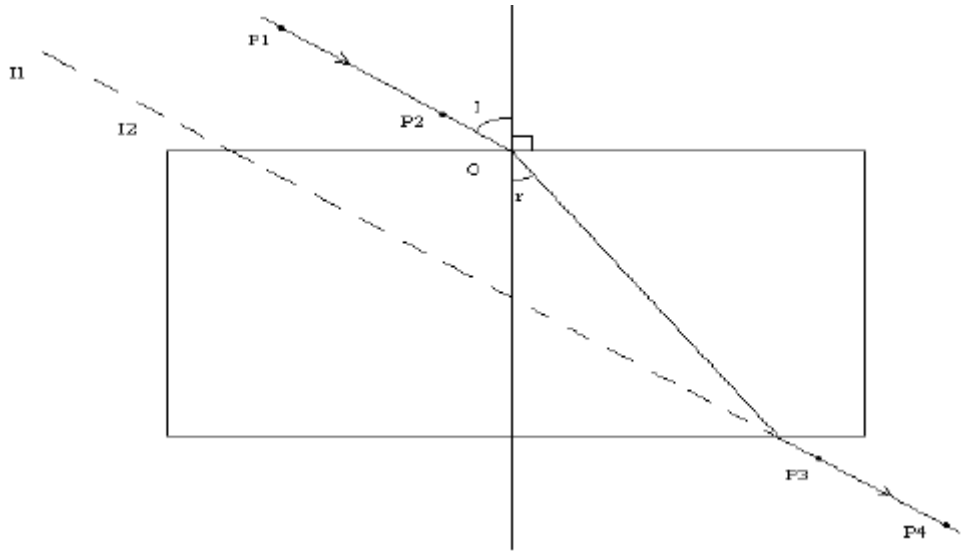
Sample Practical Question

The aim of this experiment is to find the refractive index of a glass block. Proceed as follows:

Place the given glass block in the middle of the drawing paper on the drawing board. Draw lines along the upper and lower edges of the glass block. Remove the glass block and extend the lines you have drawn. Represent the ends of these line segments as SS_1 and TT_1. Draw the normal NN_1 to the parallel lines SS_1 and TT_1 as shown in the figure below:



Draw five evenly spaced lines from O to represent incident rays at different angles of incidence (10°, 20°, 30°, 40°, and 50° from the normal). Replace the glass block carefully between SS_1 and TT_1. Stick two pins P_1 and P_2 as shown in the figure as far apart as possible along one of the lines drawn to represent an incident ray. Locate an emergent ray by looking through the block and stick pins P_3 and P_4 exactly in line with images I_1 and I_2 of pins P_1 and P_2. Draw the emergent ray and repeat the procedure for all the incident rays you have drawn. Finally draw in the corresponding refracted rays.



(a) Record the angles of incidence i and the measured corresponding angles of refraction r in a table. Your table of results should include the values of $\sin i$ and $\sin r$.

(b) Plot the graph of $\sin i$ (vertical axis) against $\sin r$ (horizontal axis).

(c) Determine the slope of the graph.

(d) What is the refractive index of the glass block used?

(e) Mention any sources of errors in this experiment.

Discussion

In this experiment, pins are used to simulate a ray of light. If all of the pins are aligned as you look through the block, they act as a single ray. It takes practice to be able to align the pins while looking through the block, so practice often with your students.

Light slows down as it enters a denser medium, so in order to minimize the time required to pass through that medium, it changes direction until it moves back into its original medium. In this case, light is moving from air into glass and then back into air, so its direction changes while inside the glass, then returns to its original direction when passing back into air. This effect is called refraction and it depends on the nature of the media, in this case air and glass. Snell's law gives us the relationship between the nature of the media and the resulting angles of incidence and refraction:

$$n_1 \times \sin i = n_2 \times \sin r$$

In this experiment, the incident angle i is being changed and the refracted angle r is being measured. The refractive index of medium 1 (air) is known as 1.0, so we can use these three to find the refractive index of medium 2 (glass). On the graph, $\sin i$ is the dependent variable and $\sin r$ is the independent variable, so the equation becomes

$$\sin i = \sin r \frac{n_2}{n_1}$$

In this case the slope must be $\frac{n_2}{n_1}$.

The refractive index of medium air is simply 1.0, so the slope is the refractive index of medium 2.

This practical is one of the easiest to perform with students because it does not require much preparation. Syringe needles should be readily available and glass blocks are cheap, so it is possible to have every students try this themselves many times before taking the exam.

Finding the Critical Angle

Some questions may ask students to find the critical angle of a glass block in addition to its refractive index. The relationship between critical angle, C, and refractive index, n, for a particular medium is given by

$$n = \frac{1}{\sin C}$$

or

$$\frac{\sin i}{\sin r} = \frac{1}{\sin C}$$

.

Thus a graph of $\sin i$ against $\sin r$ can be used to find the critical angle. However, take care to note that we must first take the reciprocal of the slope, i.e.

$$\sin C = \frac{1}{n}$$

.

This gives us $\sin C$, so to get C by itself, we need to use mathematical tables. Turn to the page for Natural Sines and search the table for the 4-figure value you obtained above by taking $\frac{1}{\text{slope}}$. The corresponding row gives the angle in degrees and the column gives the additional minutes of the angle.

For example, say we plot our graph of $\sin i$ (y-axis) against $\sin r$ (x-axis) and we calculate the slope to be 1.43. This is the refractive index of the glass block (since we can remember that glass has a refractive index of 1.5, we can do a quick mental check to make sure this makes sense). Then $\sin C = \frac{1}{1.43} = 0.6993$. From the mathematical tables, we get $C = 44°22'$ [6993 falls between 6984 ($18'$) and 6997 ($24'$), so we use the Mean Differences table to add on $4'$ giving a total of $22'$].

Note that the method of finding n and C changes if we are instead told to graph $\sin r$ (y-axis) against $\sin i$ (x-axis). Be sure to practice both versions with students to ensure their understanding.

28.5 Electricity

This is by far the least attempted practical on the exam, but not because it is difficult. The electricity practical, if properly set up, is one of the easiest to perform. It can appear in many different forms but will typically involve a simple circuit and some kind of variable resistor in order to measure current or EMF for different resistances. The materials you will probably need are as follows:

Connecting Wires Use speaker wire; it is cheap and available in most villages and towns.

Voltmeters, Ammeters, Galvanometers This is unavoidable; you can get full digital multimeters in town for about 10,000/=, galvanometers can be found in any lab store or can be made using a compass and insulated copper wire.

Batteries Two to four D-size batteries should easily be enough for these experiments. Try to avoid Tiger brand if possible. Panasonic is highly superior in quality for roughly the same price.

Resistance Wire These are used to make small resistors for the metre bridge or potentiometer. The most common type of wire to use is nichrome, which can be found in a hardware or lab store. Steel will also work, though it is less resistant and therefore harder to measure.

Metre Bridges See the activity that describes the construction of a metre bridge and potentiometer. It is best to make both together as the construction is almost identical and both are used frequently.

Variable Resistor (Rheostat) This is optional as it is typically only used to set a level that can be easily read by the voltmeter. However, if you are using a multimeter, you can simply change the magnitude setting on the multimeter to account for unusually low or high resistances.

Soldering Iron Not required, but may be a good investment for making reliable battery connections. Using electrical tape can lead to inconsistencies. Check large towns.

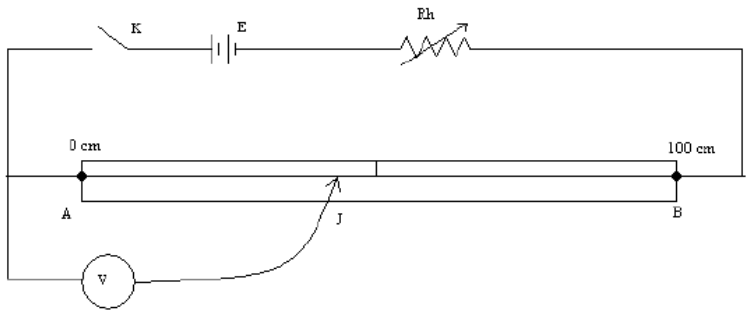
28.5.1 Potentiometers

This experiment is very simple but requires the correct materials, namely the meter bridge/potentiometer described above. A complete circuit is created with a switch (optional), power source, variable resistor and 1 m of bare resistance wire, all in series.

The potentiometer itself is simple to construct; all preparation is done by the teacher, so the student simply follows the instructions as shown in the following example from the 2007 NECTA.

Sample Practical Question

The aim of this experiment is to determine the potential fall along a uniform resistance wire carrying a steady current. Proceed as follows:



Connect up the circuit as shown in the figure. Adjust the rheostat so that when the sliding contact J is near B and the key is closed the voltmeter V indicates an almost full scale of deflection. Do not alter the rheostat again.

Close key K and make contact with J, so that AJ = 10 cm. Record the potential difference V volts between A and J as registered on the voltmeter.

Repeat this procedure for AJ = 20 cm, 30 cm, 50 cm, and 70 cm.

(a) Tabulate your results for the values of AJ and V.

(b) Plot a graph of V (vertical axis) against AJ (horizontal axis).

(c) Calculate the slope of the graph.

(d) What is your comment on the slope?

(e) State any precautions on the experiment.

Discussion

This is a simple test of the relationship between the length of a wire and its resistance, which we know is

$$R = \frac{\rho l}{A}$$

Where l is the length of the wire, ρ is the resistivity of the wire, and A is the cross-sectional area of the wire. We expect that as the length of wire increases, its potential difference will also increase. This is because the resistance (and therefore potential difference) of a wire is directly related to its length. The voltmeter in this experiment is measuring just the potential difference over the length of wire (10 cm, 20 cm, etc.), so if we use Ohm's Law to say that $V = IR$, we can write:

$$V = \frac{I\rho l}{A}$$

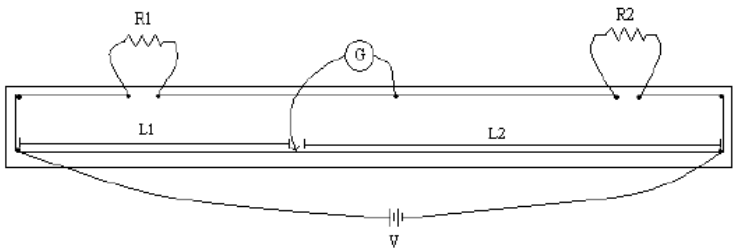
In this experiment, I, ρ and A are all constant, so the slope is

$$\text{slope} = \frac{I\rho}{A}$$

Though it is not asked for directly in the question, we can find the resistivity, ρ, by measuring I with an ammeter/galvanometer and A with vernier calipers or a micrometer screw gauge.

28.5.2 Metre Bridges

A metre bridge resembles a potentiometer, except that it uses a galvanometer to measure the difference in current between two points on the circuit, hence the name "bridge." The same materials can be used as with the potentiometer, though it is best to use small coils of resistance wire for the small resistors (between 3Ω and 20Ω is a good resistance). A galvanometer can be made easily if one is not available.



Resistors R_1 and R_2 have different resistances, but they should be somehow similar so that one resistor does not take all of the current (this will make it difficult to measure the length to the galvanometer). About 5Ω and 10Ω, for example, would work well.

However, for the sake of the practical, one resistor should not be known; the objective of the practical is to find the unknown resistance. The long wire along the bottom edge is a metre of nichrome wire or other resistance wire. One terminal of the galvanometer is connected between the two resistors, and the other terminal is connected to a flying wire (or jockey) that is free to move along the length of the nichrome wire.

The practical instructs you to move the galvanometer's flying wire back and forth along the nichrome wire until it reads zero. At this point, we know that no current is passing through the galvanometer, so the potential difference across it is zero. This means that the current flowing through R_1 is the same as that current flowing through R_2, and the current flowing through the nichrome wire is constant. From this we can conclude that

$$\frac{R_1}{L_1} = \frac{R_2}{L_2}$$

or that the ratio of the two resistors is equal to the ratio of distances from the flying wire to either end of the nichrome wire. The resistance of one resistor (say, R_1) is known and the lengths L_1 and L_2 can be measured from the flying wire to either side of the nichrome wire. Using the ratio above, we can easily calculate the unknown resistance R_2.

An example is given below from the 2006 NECTA exam.

Sample Practical Question

You are required to determine the unknown resistance labeled X using a metre bridge circuit. Connect your circuit as shown below, where R is a resistance box, G is a galvanometer, J is a jockey and others are common circuit components.

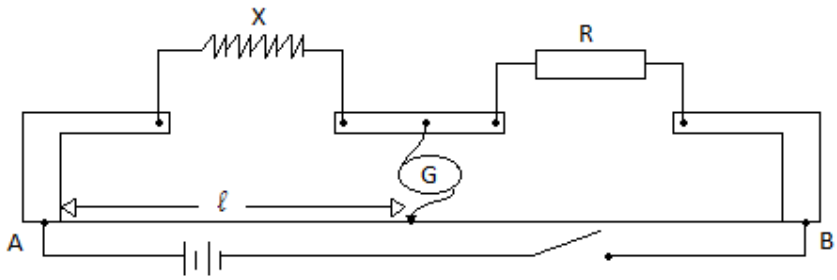


Fig. 3

Procedure:

With $R = 1\ \Omega$, obtain a balance point on a metre bridge wire AB using a jockey J. Note the length l in centimetres. Repeat the experiment with R equal to 2 Ω, 4 Ω, 7 Ω and 10 Ω.

Tabulate your results for R, l and $^1/_l$.

(a) (i) Plot a graph of R (vertical axis) against $^1/_l$ (horizontal axis).
(ii) Determine the slope S of your graph.
(iii) Using your graph, find the value of R for which $^1/_l = 0.02$.

(b) Read and record the intercept R_0 on the vertical axis.

(c) Given that,
$R = \dfrac{100\text{X}}{l} - \text{X}$
Use the equation and your graph to determine the value of X.

(d) Comment on your results in (a)(iii), (b) and (c) above.

Discussion

The procedure for this question is similar to most other wheatstone bridge problems: vary a known resistor and see how it affects the relative lengths in resistance wire required to balance the potential difference and give no current through the galvanometer. It may not be obvious at first, however, where the equation $R = \frac{100\text{X}}{l}$ - X comes from.

Starting with the balancing ratio for a wheatstone bridge,

$$\frac{R_1}{L_1} = \frac{R_2}{L_2}$$

we can solve for the unknown resistor

$$R_2 = R_1\left(\frac{L_2}{L_1}\right)$$

.

Recall that L_1 and L_2 are the corresponding lengths from either end of the metre rule to the jockey (in cm), and so taking them together, we get

$$L_1 + L_2 = 100$$

. Dividing both sides by L_1 gives

$$1 + \frac{L_2}{L_1} = \frac{100}{L_1}$$

. Then solving for $\frac{L_2}{L_1}$,

$$\frac{L_2}{L_1} = \frac{100}{L_1} - 1$$

. Now substitute this into our previous equation for R_2:

$$R_2 = R_1\left(\frac{100}{L_1} - 1\right)$$

. Replacing R_2 with R, L_1 with l and R_1 with X for this problem and distributing gives,

$$R = \frac{100\text{X}}{l} - \text{X}$$

.

From this equation, we have dependent variable R, independent variable $\frac{1}{l}$, slope 100X and y-intercept $-$X. So we can obtain the value of the unknown resistor X either by using the y-intercept (note the resistance is a positive value) or by taking the slope divided by 100.

28.5.3 Ohm's Law

The practical may give any kind of experiment to use or verify Ohm's Law in a simple circuit. Finding the e.m.f. and internal resistance of cells appears frequently. Students should be very familiar with the law, as well as the factors that determine resistance in a wire and the effect of internal resistance of a cell on a circuit. Given below is an example problem taken from the 2011 NECTA exam.

Sample Practical Question

You are provided with an ammeter, A, resistance box, R, dry cell, D, a key, K and connecting wires. Proceed as follows:

(a) Connect the circuit in series.

(b) Put $R = 1\ \Omega$ and quickly read the value of current I on the ammeter.

(c) Repeat procedure (b) above for $R = 2\ \Omega$, $3\ \Omega$, $4\ \Omega$ and $5\ \Omega$. Record your results in a tabular form.

(d) Draw the circuit diagram for this experiment.

(e) Plot the graph of R against $\frac{1}{I}$.

(f) Determine the slope of the graph.

(g) If the graph obeys the equation $R = \frac{E}{I} - r$, then

  (i) suggest how E and r may be evaluated from your graph.

  (ii) compute E.

  (iii) compute r.

(h) State one source of error and suggest one way of minimizing it.

(i) Suggest the aim of this experiment.

Discussion

To see where the equation $R = \frac{E}{I} - r$ comes from, first start with Ohm's Law, $V = IR$. Accounting for the internal resistance of the cell, r, this becomes

$$V = I(R + r)$$

. To solve for resistance, we divide both sides by I, which gives

$$(R + r) = \frac{V}{I}$$

. From this we can see that, using E as e.m.f. for this problem,

$$R = \frac{E}{I} - r$$

.

In this form, the equation resembles the classic $y = mx + c$, where R is the dependent variable, $\frac{1}{I}$ the independent variable, E is the slope and $-r$ is the y-intercept (note the internal resistance is a positive value).

Part VI

Science Activities and Competitions

Chapter 29

Hosting Science Events

There are many great ways to promote math and science education through engaging activities for students and teachers alike. These can be done regularly through extracurricular clubs, but can also be organized together as part of a larger *Science Day* event or multi-day *Math and Science Conference*. What follow are some general tips and suggestions for hosting some of these various activities.

29.1 Box of Fun

The *Box of Fun* can be used as a teacher training exercise or as a student challenge.

- Gather an assortment of everyday materials (see Sources of Laboratory Equipment, p. 208) and arrange them randomly across a table or in a large box.
- Ask participants to use the materials given to demonstrate some topic or principle in a subject of their choice (Biology, Chemistry, Physics or Math).
- You may choose to put participants into groups and designate a specific subject for each group.
- After at least 30 minutes, have groups come up to present their idea.
- Additionally, you may ask groups to fill out an activity template (see *Shika Express* companion manuals) to document their ideas.

The *Box of Fun* is intended to foster participants' creativity and encourage them to see science in the world around them. Rather than thinking first of a topic and then deciding what materials are needed to show it, this activity encourages teachers to *first look around and see what is available to them*, and then to think about how those things might be used to demonstrate some concept in science.

CAUTION: It may not be wise to assign specific topics to participants, as this can limit their creativity and may lead them to *"destroy science"* by pretending a local substitute gives the same result as a traditional lab material when it really doesn't (e.g. pretending food colour is iodine because they have similar appearances).

The purpose of using locally available materials is that they help to connect students to their everyday environment *while still achieving the same results as expensive lab equipment*. However, if a local material does not give the same result, it should NOT be substituted merely for the sake of using local materials.

29.2 Shika Express Gallery Walk

This activity can be used to share ideas of science demonstrations among students and/or teachers.

- Choose 4-6 activities or demonstrations for each subject (see *Shika Express* companion manuals for Biology, Chemistry, Physics and Math).
- Prepare the demonstrations and arrange them across a set of tables, 1-2 tables per subject.
- Spread the tables out evenly around a large empty room (e.g. dining hall).
- Divide participants into equal groups based on the number of subjects being presented.
- Have groups rotate among the different subject tables so that they are able to observe all demonstrations for each subject (approx. 15-20 mins per subject).
- Following the rotations, give 20-30 mins to allow participants to return to a demonstration of their choice for further investigation or to construct it themselves.

Suggestions:

- You may wish to have 1-2 student or teacher leaders for each subject to help explain the demonstrations during the rotations.
- Make copies of activity write-ups from the *Shika Express* manuals for each demonstration that participants can read as they walk around.
- Allow participants to perform the demonstrations themselves as much as possible, and then ask them to explain what they see.

29.3 Science Fair

Science Fair projects provide a great opportunity for students to apply their knowledge and investigate their interests in science.

- Have interested students form groups of 2-3.
- Groups select a project idea based on a shared interest or question/problem to address. Encourage students to think about what problems or issues are faced in their own communities.
- Review the steps of The Scientific Procedure (see example activities on p. 152). Have students identify the problem and form a hypothesis for their project.
- Allow several weeks for groups to work on their projects (provide additional books or computer resources if available).
- When completed, allow students to set up and explain the various projects around the school for all students to see.
- Encourage students to apply to participate in the national Young Scientists Tanzania (YST) competition (`www.youngscientists.co.tz`) in Dar es Salaam.

29.4 Science Competitions

Perform individual competitions or many strung together over the course of a day or weekend.

For more, see the section on Science Competitions (p. 145).

29.5 Science Day

Engage the entire school (or multiple schools) by combining several activities into a *Science Day* event.

- Invite a guest speaker to speak on career opportunities in math and science (e.g. accountant, engineer, doctor, nurse, carpenter, mechanic, store owner, etc.)
- Explain applications of math and science in all walks of life (e.g. farming, buying/selling, health/disease, transport, weather, drinking water, football, etc.)
- Incorporate *Science Competitions* - elect 1 or 2 teams from each Form to compete, with the rest of the school as an audience.
- Incorporate *Box of Fun* and *Gallery Walk* activities. It may be helpful to do the *Gallery Walk* first to provide examples to participants.
- Encourage girls' empowerment wherever possible.
- Give out a survey to gauge students' perception of science.

A *Science Day* event may not guarantee immediate improvements in test scores, but it shows students that *the school and its teachers are not willing to give up on math and science, and neither should they!* Continued promotion of math and science will help to change students' perception of the subjects that they may initially write off as being too difficult. Excitement and interest is the first step in changing that perception.

29.6 Math and Science Conferences

Gather students from several nearby schools to hold a special week-long *Math and Science Conference* in a nearby town or at a host school.

In addition to those ideas presented for a *Science Day* event,

- Incorporate HIV/AIDS and malaria into science demonstrations/activities (see *Shika Express* companion manuals for ideas).
- Have students prepare *Science Fair* projects over the course of the conference and present on the final day.
- Give award certificates for participation and prizes for individual/team competitions.

Math and Science Conferences encourage leadership among their participants. Students attending such events are likely to be good *ambassadors of science*, sharing what they have done and learned with fellow students back at school. Those students may then try to improve their own performance in those subjects so that they can attend a similar event later on.

29.7 Teacher Trainings

Conferences can also be directed towards improving teacher performance by conducting the suggested activities at a nearby Teacher's College.

Chapter 30

Science Competitions

Students, just like nearly all other people, enjoy competing against one another. Likening math and science-related activities to the competition of a football league can be a wonderful motivator for students. Given below are some suggestions for utilizing competitions while teaching about science.

- Combine students of various abilities together on a team. This will allow the bright students to develop leadership skills and help to bring up the slow learners.
- Limit teams to 3-5 students. Balance the number of boys and girls on a team, or choose to have all-boys teams compete against all-girls teams.
- Allow students to pick their own team name, and possibly draw a team flag if time allows.
- Create a standings board for the competition. For example:

	Egg Drop	Jenga Jengo	Raft Rally	Drop Zone	Bridge Challenge	**Total**
Big Stars						
Chelsea						
Simba						
Arsenal						
Kings						
Manchester United						

- Have teams present and explain their designs to the audience where applicable. Allow other students to ask questions/provide criticisms.
- Follow up each activity with a short lesson about a concept illustrated by the competition (e.g. Archimedes' Principle for *Raft Rally*; see competition write-ups for more).
- Ask students how they would revise their designs or make improvements if they could do the activity again.
- Explain how the activities can be applied to solve real-life problems (e.g. *Jenga Jengo* for Civil Engineers).

Egg Drop

Time: 1 hour

How It Works

Students must build a device to transport an egg through a given drop distance without cracking.

What You Need (per team)

- ☐ Plain paper – 4 sheets
- ☐ Plastic bag - 2
- ☐ String – 1 metre
- ☐ Masking tape – 1 roll
- ☐ Balloons – 4
- ☐ Straws – 10
- ☐ Toothpicks – 10
- ☐ Tongue depressors – 4
- ☐ Rubber bands – 4
- ☐ Index cards – 4
- ☐ Paper clips - 10
- ☐ Toilet paper – 1 roll
- ☐ Bottle (500 ml) - 1
- ☐ Newspaper – 1 sheet
- ☐ Egg – 1

Rules

- 45 minute time limit for construction.
- Devices dropped from a height of 3-5 metres.
- Teams can use only the materials given, but do not need to use everything.
- Egg is placed at time of testing. It must be possible to place and remove egg freely without altering the device.
- Once egg is placed, no further adjustments may be made. This means the egg cannot have any kind of "seat belt" or strap fastened after placing the egg.

Points

Egg Survives - 50 pts Egg Cracks - 0 pts

Additional Materials

- Ladder / chair
- Scissors for community use

Notes

- Do not give eggs to teams until time of testing.
- If possible, increase drop height for surviving eggs and give bonus 25 pts for each additional successful drop.

Science Applications

Air Resistance (Physics Form I): Air resistance provides a frictional force which opposes the object's motion as gravity attracts it towards the centre of the earth. This upward force reduces the speed of the object as it falls, allowing it to land more softly and protect the egg. Thus, we want to maximize the air resistance on the object (e.g. by using a large parachute).

Pressure (Physics Form I): The force of impact on the device when it hits the ground can be reduced by increasing the surface area which contacts the ground. Constructing a wide base (e.g. using balloons) reduces the impact on the egg and thus helps to protect it.

Taking It Further

Did students utilize the parachute concept? If not, show a brief example. How does this help to protect the egg? Is it better to have a large parachute or a small one?

Jenga Jengo

Time: 30-45 minutes

How It Works

Students must build the tallest structure possible, using only paper and tape, as quickly as possible and while ensuring good stability.

What You Need (per team)

☐ Plain paper – 25 sheets
☐ Masking tape – 1 roll

Rules

- 20-minute time limit for construction.
- Cannot use tape roll as weight inside structure.
- Stability tested by waving book at structure (Wind Test).

Points

1st to finish – 50 pts
Tallest structure – 50 pts
Passes Wind Test – 50 pts

Additional Materials

- Tape Measure
- Stopwatch
- Book / waving device

Notes

- All structures that pass the Wind Test are awarded 50 pts.
- Alternate Wind Test: place structures outside on a windy day. Those standing after 1 minute pass.

Science Applications

Centre of Gravity (Physics Form I): Civil engineers construct buildings with a low centre of gravity, making them less likely to fall over due to wind forces. To maintain *stable equilibrium*, a building should have a wide base with a large mass, while the top of the building should have a small area and less mass.

Taking It Further

- Show students pictures of buildings and structures from around the world after the competition. Did the students' structures resemble any of them?
- Try variations, giving students index cards, straws or matches instead of plain paper.

Raft Rally

Time: 30-45 minutes

How It Works

Students must build a raft using only aluminum foil that can support the heaviest load before sinking.

What You Need (per team)

☐ Aluminum foil – 20 cm × 20 cm sheet
☐ Straws - 4 (optional)

Rules

- 10-minute time limit for construction.
- Replacement sheet may be given in case of rips/tears, at a 20 pt deduction.

Points

1st Place – 100 pts
2nd Place – 75 pts
3rd Place – 50 pts
4th Place – 25 pts
Others – 0 pts

Additional Materials

- Large container or bucket (clear if possible) filled with water
- Nails (× 200) / Bottle caps (× 200) / Other small weights for testing

Notes

- As raft approaches the point of sinking, add weights more slowly.
- Raft is finished when water begins to enter, and total number of weights is recorded.

Science Applications

Archimedes' Principle (Physics Form I): *Archimedes' Principle* states that

$$\text{Upthrust} = \text{Weight of displaced fluid}$$

Here, we want to maximize the force of upthrust to avoid sinking. So that means maximize the Weight of the displaced water: Weight = mass × acceleration due to gravity, or

$$W = mg$$

Gravity is a constant , but mass depends on 2 things: density (ρ) and volume (V). We know that $\rho = \frac{m}{V}$, so that means

$$m = \rho V$$

The density of the water is constant, so the only thing we can change is the *Volume* of water displaced. Thus to get the most upthrust and prevent sinking, we need to displace a large volume of water, i.e. build a raft with a large base.

Taking It Further

- Ask students how they would revise their designs if they could do it again.
- Try variations, giving students straws, toothpicks, tongue depressors or index cards.

Drop Zone

Time: 30-45 minutes

How It Works

Students must build a parachute using limited materials to carry a paper clip passenger as close as possible to a target, while maximizing hang time.

What You Need (per team)

- ☐ Paper clip – 1
- ☐ String – 1 metre
- ☐ Masking tape – 15 cm
- ☐ Plastic bag - 2
- ☐ Plain paper – 2 sheets
- ☐ Scorecard (see example below)

Rules

- 10-15 minute time limit for construction.
- Parachutes dropped from a height of 3-5 metres.
- Average hang time and distance from target taken over 3 trials for each team.

Points

Hang Time (Longest):	1st – 50 pts,	2nd – 35 pts,	3rd – 20 pts,	4th – 5 pts,	Others – 0 pts
Distance (Shortest):	1st – 50 pts,	2nd – 35 pts,	3rd – 20 pts,	4th – 5 pts,	Others – 0 pts

Additional Materials

- Tape measure
- Stopwatch
- Flipchart target
- Ladder / chair

Notes

- Scorecard:

Team:	**Trial 1**	**Trial 2**	**Trial 3**	**Average**	**Points**
Hang Time (s)					
Distance from Target (cm)					

- Measure distance from paper clip to centre of target.

Science Applications

Air Resistance (Physics Form I): Air resistance provides an upward force on the parachute, which acts against the force of gravity and causes the object to fall more slowly. The larger the surface area of the parachute, the more slowly it will fall.

Taking It Further

- Students may not be familiar with parachutes. Prepare a simple example to explain the concept and function.
- Ask students questions: Why does the parachute slow the object down? To maximize hang time, do we want a very large or very small parachute? Would a parachute work on the moon?
- Drop parachute side-by-side with a paper clip having no parachute. Which one made it safely?

Bridge Challenge

Time: 1 hour 30 minutes

How It Works

Students must build a bridge that can support the most weight, while using a limited budget of *Science Shillings* to purchase construction materials.

What You Need (per team)

- ☐ Straws - 20
- ☐ Bamboo skewers - 20
- ☐ Bamboo stick (fimbo) – 1
- ☐ Tongue depressors – 10
- ☐ Toothpicks – 2 small cans
- ☐ String – 3 m
- ☐ Office glue – 1 tube
- ☐ Rubber bands - 10
- ☐ Pencils - 2
- ☐ Index cards – 10
- ☐ Masking tape – 1 roll
- ☐ Duct tape – 1 roll (total)
- ☐ Ruler - 1
- ☐ Scissors - 1
- ☐ Science Shillings – 20 (see p. 221)

Rules

- Approximately 45 minute time limit for construction.
- Teams begin with only a ruler, scissors and 20 *Science Shillings*. These items may NOT be used in construction of bridge.
- All building materials must be purchased from a science shop. Suggested prices are as follows:

Straws (bundle of 10)	1 /=	Office glue (1 tube)	3 /=
Skewers (bundle of 10)	2 /=	Rubber bands (× 5)	1 /=
Fimbo	3 /=	Pencil	1 /=
Tongue depressors (× 5)	2 /=	Index cards (× 5)	1 /=
Toothpicks (2 small cans)	1 /=	Masking tape (1 roll)	3 /=
String (1 metre)	1 /=	Duct tape (30 cm)	1 /=

- Bridges will be loaded by placing rocks or other weights into a small bucket that must rest on top of the bridge.
- Bridge must span a 30 cm gap between two chairs / tables.
- 1 student from each team must be designated as team accountant. Only this student may purchase items from the shop.

Points

(Based on number of rocks / weights placed before bridge fails)
1st – 100 pts, 2nd – 75 pts, 3rd – 50 pts, 4th – 25 pts, Others – 0 pts
BONUS: 5 pts per *Science Shilling* remaining after construction

Additional Materials

- Small bucket
- Several large rocks (20) / other large weights
- 2 chairs / tables 30 cm apart
- Science shop table
- Extra bamboo sticks (fimbos) – 2-4
- Index Cards for price signs – 12

Notes

- Student accountants may only purchase 1 of each item at a time. They must allow other students to make purchases before buying another of that item.
- At some point 15-20 minutes into construction time, shopkeeper may announce a newly received shipment of bamboo sticks (fimbos). However, due to demand, the price has increased to 5 /=.
- Bridge has failed when it either collapses / breaks or when the bucket can no longer be balanced on top of it. Record largest number of weights successfully added.

Taking It Further

- Ask students to present their bridges and describe how they decided to manage their money. What materials did they purchase and why?
- What would they do differently if they could start again?
- Which team finished with the most total points after the **BONUS**? Did the most money spent result in the strongest bridge? Who was the most *efficient* with their money?
- Show students pictures of bridges from around the world after the competition. Did the students' bridges resemble any of them?

Chapter 31

The Scientific Procedure

The following activities can be used as a method of introducing students to the scientific method. Rather than just performing the activities, first identify the question or problem with the students, then have them form a hypothesis for each step of the experiment. Students should record observations and data accordingly and use them to draw a conclusion about the activity.

Prepare an activity sheet for each student or have them copy it into their notebooks before performing the activities. Set up stations for the various activities and have students rotate among them in small groups.

After performing several of the experiments, ask students to come up with their own. Ask them to think about problems they face in their daily lives. Encourage interested students to turn their ideas into a science fair project to display for the school or community.

Biology

Hand Washing

Materials: Soap, water, bottle, basin/bucket, chalk, charcoal, food colour, stopwatch
Setup: Prepare a large amount of soapy water. Grind the chalk and charcoal into separate powders.

Problem: How long should we wash our hands?

Material	**Hypothesis (Seconds)**	**Experimental Result**
Chalk powder		
Charcoal powder		
Food colour		

Hypothesis: Predict how much time it will take to completely clean your hands and record in the table.
Procedure: Start a stopwatch and have a student or teacher slowly pour soapy water over a basin while the student washes his or her hands. Stop the clock when the student's hands are completely clean.
Observations: Record the time taken to completely wash your hands in the table.
Questions:

1. Why is it important to wash our hands?

2. When do we need to wash our hands?

Theory: Washing our hands with soap and water helps to kill harmful bacteria that can cause us to become sick if allowed into our bodies. It is very important to wash our hands before eating and after using the bathroom.

Lung Capacity

Materials: 1.5 L bottle, basin, water, plastic tubes/straws, soap, marker, ruler
Setup: Make a scale on the bottle using a marker and ruler (e.g. 100 mL increments). Prepare a soap solution for washing the tubes/straws

Problem: How much air can your lungs hold?

Breath	**Hypothesis (Volume of air in mL)**	**Experimental Result**
Normal breath		
Full breath		
After holding breath for 10 seconds		

Hypothesis: Record the volume of air that you think the lungs can hold for each case in the table.
Procedure: Fill a basin with water. Fill a 1.5 L bottle with water and invert it in the basin so that the mouth of the bottle is underneath the water. Place one end of the tube/straw inside the bottle under water. For each breath, blow into the tube to displace the water.
Observations: Note the reading on the scale before and after blowing into the tube and record the *difference* to give the amount of water displaced.
Questions:

1. Which breath produces the largest amount of air? Which give the smallest amount?

2. How long can you hold your breath?

   Hypothesis: I can hold my breath for _____ seconds.

   Experimental Result: I can hold my breath for _____ seconds.

Theory: When we breath in air, our bodies use the oxygen and produce carbon dioxide in a process called *respiration*. Oxygen is transported in our blood throughout our bodies. When we hold our breath, oxygen is not circulated throughout our bodies and we begin to feel lightheaded.

Chemistry

Acids and Bases

Acids or Bases?

Materials: Bottles, bottle caps, water, vinegar, lemons, baking soda, soda, soap, antacid tablets, rosella leaves, straws/syringes

Setup: Prepare solutions for each of the items above in separate bottles. Prepare indicator by placing rosella leaves in hot water.

Problem: What differences can we observe among acids and bases?

Solutions	**Hypothesis (Which is different?)**	**Experimental Result**
Vinegar, lemon, baking soda		
Vinegar, baking soda, soap		
Baking soda, antacid, soda		
Soda, soap, vinegar		

Hypothesis: For each set of solutions, which one will reveal a colour different from the others? Record your predictions in the table.

Procedure: Place small amounts of 3 different solutions in separate bottle caps according to the table. Add a few drops of rosella indicator to each.

Observations: Record observations of colour change under *Experimental Result* in the table.

Questions:

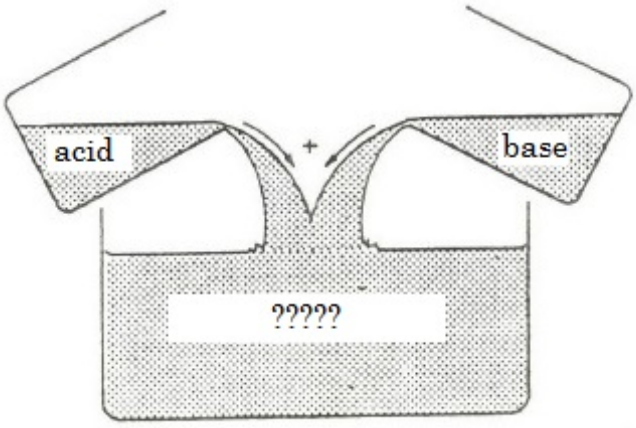
1. Which solutions have similar properties?

2. Which solutions are acids? What colour do they show?

3. Which solutions are bases? What colour do they show?

Theory: Coloured leaves such as rosella act as indicators for identifying acids and bases. Adding rosella indicator reveals a red colour for acids and a blue colour for bases. Students do not need to understand the differences between acids and bases in order to observe their different behaviours. Locally available examples of acids include sour milk, citrus fruits and soda. Local bases include ammonia, toothpaste and detergent.

Mixing Acids and Bases



Problem: What happens when acids and bases are mixed together?

Solutions to Mix	**Hypothesis (What colour?)**	**Experimental Result**
Mix vinegar and lemon		
Mix baking soda and soap		
Mix vinegar and baking soda		

Hypothesis: Predict any colour changes or observations when pairs of solutions are mixed together. Record in the table.
Procedure: Mix small amounts of solutions together according to the table.
Observations: Record observations (colour changes, etc.) in the table.
Questions:

1. What happens when an acid is mixed with an acid?

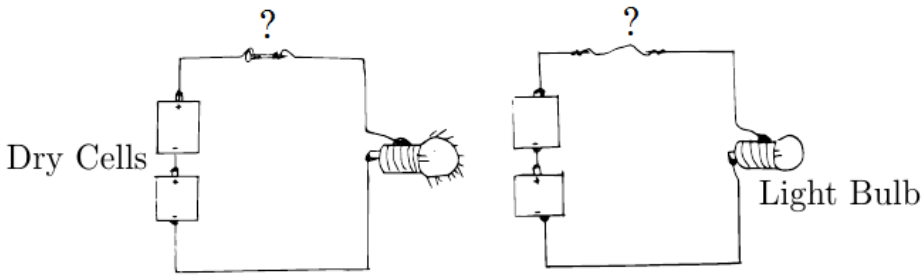
2. What happens when a base is mixed with a base?

3. What happens when an acid is mixed with a base?

Theory: Mixing acids with acids and bases with bases may cause the colour of the solution to turn darker or lighter depending on the solutions used. Mixing an acid with a base should reveal a colourless solution and produce carbon dioxide gas. You may need to vary the amounts of acid and base to get a colourless solution depending on their concentrations.

Physics

Complete the Circuit



Materials: Dry cell, speaker wire, bulb/ammeter, cardboard, various objects, e.g. rubber band, nail, paper, aluminum foil, toothpick, pen, scissors, bottle cap, coin, balloon, chalk
Setup: Connect a dry cell and bulb in series using speaker wire and attach to a sheet of cardboard. Leave two wires free and pin to the cardboard to act as a switch.
Problem: Which objects will light a bulb?

Object	**Hypothesis (Light or No Light)**	**Experimental Result**
Copper wire		
Pen		
Aluminum foil		
Paper		
Nail		
Toothpick		
Bottle cap		
Balloon		
Chalk		
Scissors (blade)		
Scissors (handle)		

Hypothesis: Predict which materials will cause the bulb to light when placed across the switch. Record predictions in the table.
Procedure: Test each object by placing it across the free wires to close the circuit.
Observations: Record the result for each item in the table.
Questions:

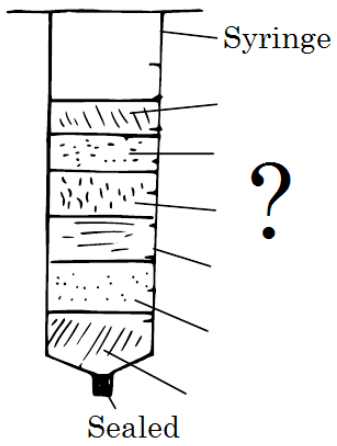
1. Which materials caused the bulb to light?

2. These objects are made from what kind or materials?

3. What other objects in the room can you find to test? Will they light the bulb?

Theory: *Conductors* are materials which easily allow electrons to flow through them. *Insulators* are materials which do not easily allow the the flow of electrons. Examples of good conductors are most metals, water and the human body. Examples of good insulators are rubber, wood and plastic.

Density Tower



Materials: Syringes, bottles, water, cooking oil, kerosene, spirit, honey, glycerine, tape, scissors
Setup: Prepare a test tube rack by cutting a bottle and filling it with dirt. Remove the plungers from the syringes and seal them with tape, super glue, or by melting to opening closed.

Problem: Which liquids are more dense than others?

Liquid	**Hypothesis (Position, 1 = bottom)**	**Experimental Result**
Water		
Cooking oil		
Kerosene		
Spirit		
Honey		
Glycerine		

Hypothesis: Predict the order in which the liquids will settle from the bottom of the syringe. Assign 1 to the bottom liquid, 2 to the one above it, and so on.
Procedure: Pour a small amount of each liquid into a syringe, observing after each addition.
Observations: After adding all liquids, record the order in which they rest, starting with 1 at the bottom.
Questions:

1. Which liquid finished at the bottom?
2. Which liquid finished at the top?
3. Which liquid has the greatest density?
4. Which liquid has the lowest density?
5. What happens if you place a small object (e.g. paper clip, eraser, paper) in the tower?

Theory: *Density* is a property of different materials and liquids. It is a ratio of its mass to its volume. Dense liquids sink to the bottom, while less dense liquids rise to the top. A small object placed in the tower will settle in the liquid which is nearest its own density.

Sinkers and Floaters

Materials: Basin of water, various objects, e.g. nail, paper clip, paper, aluminum foil, soda cap, matchbox, pen cap, toothpick, balloons, flour
Problem: Which objects sink or float when placed in water?

Object	**Hypothesis (Sink or Float)**	**Experimental Result**
Nail		
Paper clip		
Pen cap		
Soda cap (dropped)		
Soda cap (placed carefully)		
Toothpick		
Paper		
Aluminum foil		
Matchbox		
Balloon (empty)		
Balloon (filled with flour)		
Balloon (filled with water)		
Balloon (filled with air)		

Hypothesis: Predict whether each object will sink or float when placed in the basin of water. Record in the table.
Procedure: Place each object in the water. First place them very carefully, then drop them in.
Observations: Record the results in the table.
Questions:

1. What factors affect whether an object sinks or floats?

2. How do large objects such as boats float?

Theory: *Flotation* depends on several things. A bottle cap placed carefully on the surface of the water will float, but when pushed under, will sink. A sheet of aluminum foil will float while a sheet of the same size which is folded several times will sink. A balloon filled with flour sinks, one filled with water just floats, and one filled with air floats above the surface.

If an object's *total density* is greater than that of water, it sinks, but if less than water, it floats. Air has a density less than water, so when air is trapped in objects such as bottle caps or balloons, they float because their total density is less than water. When air is removed (folded aluminum foil) or replaced by water (bottle cap), the total density of the object is just the density of the material. A matchbox pushed under water rises back to the surface because its density is less than that of water.

Boats are able to float despite being built from dense materials because of the large volume of water they displace and the large amount of air inside the boat. A boat with a larger surface area displaces a larger volume of water and thus can carry a larger load before sinking.

Follow up this activity with the *Raft Rally* science competition.

Mixing Colours

Materials: Various food colours, syringes, bottle, scissors, tape, paper
Setup: Prepare a test tube rack by cutting a bottle and filling it with dirt. Remove the plungers from the syringes and seal them with tape, super glue, or by melting to opening closed.

Problem: What happens when we mix different colours?

Colours to Mix	**Hypothesis (What colour?)**	**Experimental Result**
Red and green		
Yellow and blue		
Red and yellow		
All colours		

Hypothesis: Predict which colour will result when the two colours given are mixed together. Record it in the table.
Procedure: Use syringes to remove small amounts of each colour and place on a sheet of paper. Be sure to lay down plenty of paper so that the colours do not bleed through onto the table!
Observations: Record the resulting colour mixture in the table.
Questions:

1. How can you make orange from other colours?
2. What colour do you get by mixing all of the colours together?
3. What are some uses of coloured dyes?

Theory: Red, green and blue are *primary colours* of light. Other colours are made by different combinations of these primary colours. Coloured dyes are used for many applications, including clothes, paper and printing pictures.

NECTA Practical Past Papers

Biology Past Papers

2013 - BIOLOGY 2A (ACTUAL PRACTICAL A)

1. You have been provided with solution **B**.

   (a) Identify the food substances present in solution **B** by using the reagents provided. Tabulate your work as shown in the following Table:

Food Tested	Procedure	Observation	Inference

   (b) For each food substance identified in 1(a);

      (i) Name two common sources.

      (ii) State their role in the body of human being.

   (c) The digestion of one of the identified food substance in 1(a) starts in the mouth.

      (i) Name this food substance.

      (ii) Identify the enzyme responsible for its digestion in the mouth.

   (d) The digestive system of human being has several parts.

      (i) Name the part of digestive system in which most of digestion and absorption of food takes place.

      (ii) Explain how the named part in (d) (i) is adapted for absorption of digested food substances.

2. You have been provided with specimens $\mathbf{S}_1$, $\mathbf{S}_2$, $\mathbf{S}_3$ and $\mathbf{S}_4$.

   (a) Use the hand lens to observe these specimens then:

      (i) Identify specimen $\mathbf{S}_1$, $\mathbf{S}_2$, $\mathbf{S}_3$ and $\mathbf{S}_4$ by their common names.

      (ii) Classify specimen $\mathbf{S}_1$, $\mathbf{S}_2$ and $\mathbf{S}_3$ to Class level.

   (b) Study specimen $\mathbf{S}_1$ carefully then answer the following questions:

      (i) Draw a neat, large and well labeled diagram of specimen $\mathbf{S}_3$.

      (ii) State the habitat of specimen $\mathbf{S}_3$.

      (iii) In what ways is specimen $\mathbf{S}_3$ important to a farmer?

   (c) State two advantages of specimen $\mathbf{S}_1$.

   (d) State four advantages of specimen $\mathbf{S}_4$.

   (e) Give reason why specimen $\mathbf{S}_4$ was formally placed in the Kingdom Plantae?

2012 - BIOLOGY 2A (ACTUAL PRACTICAL A)

1. You have been provided with specimens **F** and **G**.
   (a) Study specimens **F** and **G** carefully, then:
      (i) Identify specimens **F** and **G** using their common names.
      (ii) Compare specimens **F** and **G**, then state their observable differences.
      (iii) Briefly explain the types of germination which occurs in specimens **F** and **G**.
   (b) Using a scalpel, remove the outer coat from specimen **F**. Split the two parts with the inner sides facing upwards. Then:
      (i) Draw a well labelled diagram to show the structures of one part of the split specimen **F** as would be seen from above.
      (ii) For each structure labelled in specimen **F**, state the role they play in seed germination.
   (c) Using a scalpel, prepare a longitudinal section of specimen **G**.
      (i) Draw a well labelled diagram of the cut surface of specimen **G**.
      (ii) Identify the part used by specimen **G** to absorb water during seed germination.
2. You have been provided with specimens **H**, **I**, **J** and**K**.
   (a) Study carefully specimens **H** and **I** then:
      (i) Identify specimens **H** and **I** using their common names.
      (ii) Suggest the mode of locomotion of specimens **H** and **I**. Give reason to support your answer.
      (iii) State the features used to place specimen **H** in the Kingdom Animalia.
   (b) Use the hand lens to observe specimens **J** and **K** then:
      (i) Identify specimens **J** and **K** by their common names.
      (ii) Name the habitats for each of specimens **J** and **K**.
      (iii) Briefly explain the features which enable specimen **H** to survive in its habitat.
      (iv) Classify specimens **J** and **K** to the phylum level.
      (v) Write down one advantage and one disadvantage for each specimen **J** and **K**.

2011 - BIOLOGY 2A (ACTUAL PRACTICAL A)

1. The solution prepared contained various food substances.

   (a) Use the chemicals and reagents provided to identify the food substances present in solution $\mathbf{S}_1$. Tabulate your work as shown in the following Table:

FOOD TESTED	PROCEDURE	OBSERVATION	INFERENCE

   (b) State the function in the human body of each food identified in 1(a) above.

   (c) Name two enzymes necessary for digestion of food substance(s) identified in (a) above.

   (d) To each type of food identified above, name at least one source in which the food has been extracted.

2. Study specimen **A**, **B** and **C** then:

   (a) Write common names of specimen **A**, **B** and **C**.

   (b) Classify specimen **A** and **B** to the phylum level.

   (c) State the habitat and one economic importance of specimen **A**.

   (d) Outline four economic importance of specimen **B**.

   (e) Use the scalpel provided to cut specimen **C** longitudinally into two equal halves. Then, draw a neat, well labelled diagram of a specimen.

   (f) Name the division of specimen **C**.

   (g) State the observable features you can use to place the specimen into its respective phylum/division.

2010 - BIOLOGY 2A (ALTERNATIVE A PRACTICAL)

1. You have been provided with solution $\mathbf{T}_1$.

   (a) Carry out food tests to identify the substances present in solution $\mathbf{T}_1$. Record your work in a table as shown below.

Test for	Procedure	Observation	Inference

   (b) What are the functions of the food substances identified in $\mathbf{T}_1$ in the human body?

   (c) (i) State the favourable/suitable pH condition at which the enzymes which digest the food substances present in $\mathbf{T}_1$ work best.

   (ii) Which of the food substances present in $\mathbf{T}_1$ is not stored in the human body?

   (iii) What happens when the levels of this substance mentioned in (c) (ii) above, rises in the body?

2. You are provided with specimens **A**, **B**, **C**, **D** and **E**. Observe them carefully and answer the questions that follow:

   (a) (i) Write down the common names of specimens **A**, **B**, **C**, **D** and **E**.

   (ii) To which kingdom do specimens **C** and **D** belong?

   (iii) Name one (1) common epidemic disease transmitted by specimen **A**.

   (b) (i) Draw a large well labelled diagram of specimen **C**.

   (ii) State the economic importance of specimen **C**.

   (c) (i) What are the distinguishing characteristics of the Phylum/Division to which specimen **E** belongs?

   (ii) Where can specimen **E** be found?

2009 - BIOLOGY 2A (ALTERNATIVE A PRACTICAL)

1. (a) You are provided with solution $\mathbf{S}_1$. Carry out experiments to identify the food substances present in it. Record your procedure, observation and inferences as shown in the table below.

Test for	Procedure	Observation	Inference

   (b) (i) Name the food substances you have identified.
   (ii) State **two (2)** sources of each food substance named in 1(b) (i) above.
   (iii) Mention **one (1)** role of each food substance you have identified.

   (c) In which parts of the digestive system are the above mentioned food substances digested? In each case mention the enzyme and the products.

2. (a) Using a hand lens examine specimen A_1.
   (i) Identify specimen A_1 by its common name.
   (ii) Name the phylum and class to which specimen A_1 belongs.
   (iii) Give an example of another organism which belongs to the same phylum as specimen A_1.

   (b) Draw a well labelled diagram of specimen A_1.

   (c) How is specimen A_1 adapted to its mode of nutrition?

   (d) What is the economic importance of specimen A_1?

   (e) Where can specimen A_1 be found?

2008 - BIOLOGY 2A (ALTERNATIVE A PRACTICAL)

1. You have been provided with specimens S_1, S_2, S_3 and S_4. Observe the specimens carefully and answer the following questions:

   (a) (i) What characteristics are common among specimens S_1, S_2, S_3 and S_4? **(3 marks)**

   (ii) Name the kingdom and phylum/division to which specimens S_1, S_2, S_3 and S_4 belong. **(4 marks)**

   (iii) Why are S_3 and S_4 placed in different classes? **(2 marks)**

   (b) (i) What distinctive features place specimen S_2 in its respective kingdom? **(2 marks)**

   (ii) Why are specimens S_3 and S_4 classified under the same phylum? **(4 marks)**

   (c) (i) Suggest how the specimen labelled S_1 is adapted to its mode of life. **(4 marks)**

   (ii) Give reasons why specimen S_1 can not grow taller? **(2 marks)**

   (d) Describe the advantages and disadvantages of the organisms which belong to the class into which S_3 is found. **(4 marks)**

2. You have been provided with a variegated leaf and iodine solution. Carefully follow the instructions given below and answer the questions that follow.

   (i) Heat some water to boiling point in a beaker and then turn off the source of heat.
   (ii) Use forceps to dip the leaf in the hot water for about 30 seconds.
   (iii) Remove the leaf from the beaker.
   (iv) Push the leaf into the bottom of the test-tube and cover it with alcohol (ethanol).
   (v) Place the tube in hot water until the alcohol boils together with the leaf.
   (vi) Remove the leaf from the test-tube containing ethanol and dip it into hot water.
   (vii) Spread the decolourized leaf on a white tile and drop iodine solution on to it.

   (a) What was the aim of the experiment?

   (b) Why was the leaf dipped in hot water for 30 seconds?

   (c) (i) Give reason, why the leaf was boiled in ethanol?

   (ii) Why was the leaf dipped once again in hot water?

   (d) Give the interpretation of the results observed when a few drops of iodine solution were poured onto the decolourized leaf.

   (25 marks)

2007 - BIOLOGY 2A (ALTERNATIVE A PRACTICAL)

1. You are provided with solution **S**.

   (a) Carry out experiments to identify the food substance present in solution **S**.

   (i) Record your experimental work as shown in Table 1 below. **(16 marks)**

   Table 1

Test for	Procedure	Observation	Inference

   (ii) Solution **S** contains ——. **(3 marks)**

   (b) Suggest one storage organ in a plant from which solution **S** might have been prepared. **(1 mark)**

   (c) For each food substance identified in (a)(ii) above, name its end product(s) of digestion. **(4 marks)**

   (d) Which of the identified food substance is mostly needed by small children? **(1 mark)**

2. You are provided with a beaker, tea bag and hot water. Carry out the following experiment:

   Pour about 100 cm^3 of hot water into the beaker.
   Put the tea bag into the beaker containing hot water.
   Observe carefully the experiment for a few minutes.

   (a) (i) What happened to the tea bag when it was put in hot water? **(3 marks)**
   (ii) Explain why the changes you observed occurred? **(4 marks)**

   (b) (i) What do you think was the aim of the experiment? **(3 marks)**
   (ii) Draw a conclusion from the experiment. **(3 marks)**

   (c) (i) Name the physiological process investigated in this experiment. **(3 marks)**
   (ii) Define the process named in (c)(i) above. **(4 marks)**
   (iii) What is the importance of this process in nature? **(5 marks)**

3. Study the specimens **J**, **K**, **L**, **M** and **N** provided.

   (a) Identify specimens **J**, **K**, **L**, **M** and **N** by their common names. **(5 marks)**

   (b) Name the kingdoms for each of specimens **J**, **K**, **L**, **M** and **N**. **(5 marks)**

   (c) Suggest the possible habitats for specimens **J** and **K**. **(4 marks)**

   (d) Draw and label specimen **N**. **(7 marks)**

   (e) List **four (4)** observable differences between specimens **J** and **K**. **(4 marks)**

Chemistry Past Papers

2013 - CHEMISTRY 2A ACTUAL PRACTICAL A

1. You are provided with the following solutions:

   JJ: Containing 3.0 g of acetic acid in 0.50 dm^3 of solution;

   KK: Containing 1.5 g of impure potassium hydroxide in 250 dm^3 of solution;

   Phenolphthalein indicator.

   Questions:

   (a) Is the use of methyl orange indicator in this experiment as suitable as phenolphthalein? Give a reason for your answer.

   (b) Titrate the acid (in a burette) against the base (in a conical flask) using two drops of your indicator and obtain three titre values.

   (c) (i) ____ cm^3 of **JJ** required ____ cm^3 of **KK** for complete reaction.

   (ii) Write a balanced chemical equation for the reaction between **JJ** and **KK**.

   (d) Showing your procedures clearly, calculate the percentage purity of potassium hydroxide.

2. Your are provided with the following:

   $\mathbf{L}_1$: 0.50 M sodium thiosulphate;

   $\mathbf{L}_2$: 0.10 M hydrochloric acid;

   Distilled water;

   Stop watch;

   Plain paper.

   Theory
   When a solution of sodium thiosulphate is mixed with hydrochloric acid, they react quantitatively and gradually the solution becomes opaque.

   Procedure

   (i) Write a clear letter X on a white piece of paper.

   (ii) Place a 100 cm^3 beaker on top of letter X, such that the letter X is visible when viewed from above.

   (iii) Using a measuring cylinder, measure 25 cm^3 of $\mathbf{L}_1$ and pour into the 100 cm^3 beaker in (ii) above.

   (iv) Measure 25 cm^3 of $\mathbf{L}_2$ and pour it into the beaker containing solution $\mathbf{L}_1$ in (iii) above and immediately start the stop watch/clock.

   (v) Shake the reaction mixture only once and record the time taken for the letter X to disappear completely.

   (vi) Repeat steps (ii) to (v) by varying the volume of $\mathbf{L}_1$ and distilled water as indicated in Table 1.

Table 1

Volume of L_1 in cm^3	**Volume of water in cm^3**	**Volume of L_2 in cm^3**	**Time (t)/s**	**Rate of reaction $\frac{1}{t}(s^-1)$**
25	0	25		
20	5	25		
15	10	25		
10	15	25		
5	20	25		

Questions:

(a) What is the aim of this experiment?

(b) Complete Table 1.

(c) Write the electronic configuration of the product which causes the solution to cloud letter X.

(d) With state symbols, write the ionic equation for the reaction between $\mathbf{L}_1$ and $\mathbf{L}_2$.

(e) Plot a graph of volume of $\mathbf{L}_1$ against rate of reaction.

(f) What can you conclude from the graph?

3. Sample **U** contains one cation and one anion. Using systematic qualitative analysis procedures, record carefully your experiments, observations, inferences and finally identify the anion and cation present in sample **U**. Record your work in a tabular form as Table 2 shows.

S/n	**Experiment**	**Observation**	**Inference**

Conclusion

(i) The cation in sample **U** is ____.

(ii) The anion in sample **U** is ____.

2012 - CHEMISTRY 2A ACTUAL PRACTICAL A

1. You are provided with the following solution:

   TZ: Containing 3.5 g of impure sulphuric acid in 500 cm^3 of solution;
   LO: Containing 4 g of sodium hydroxide in 1000 cm^3 of solution;
   Phenolphthalein and Methyl indicators.

   Questions:

   (a) (i) What is the suitable indicator for the titration of the given solutions? Give a reason for your answer.
   (ii) Write a balanced chemical equation for the reaction between **TZ** and **LO**.
   (iii) Why is it important to swirl or shake the contents of the flask during the addition of the acid?

   (b) Titrate the acid (in a burette) against the base (in a conical flask) using two drops of your indicator and obtain three titre values.

   (c) (i) ____ cm^3 of acid required ____ cm^3 of base for complete reaction.
   (ii) Showing your procedures clearly, calculate the percentage purity of **TZ**.

   (20 marks)

2. Your are provided with the following materials:

   ZO: A solution of 0.13 M $Na_2S_2O_3$ (sodium thiosulphate);
   UU: A solution of 2 M HCl;
   Thermometer;
   Heat source/burner;
   Stopwatch.

   Procedure:

   (i) Place 500 cm^3 beaker, which is half-filled with water, on the heat source as a water bath.
   (ii) Measure 10 cm^3 of **ZO** and 10 cm^3 of **UU** into two separate test tubes.
   (iii) Put the two test tubes containing **ZO** and **UU** solutions into a water bath.
   (iv) When the solutions attain a temperature of 60^oC, remove the test tubes from the water bath and pour both solutions into 100 cm^3 empty beaker and immediately start the stop watch.
   (v) Place the beaker with the contents on top of a piece of paper marked **X**.
   (vi) Note the time taken for the mark **X** to disappear.
   (vii) Repeat step (i) to (vi) at temperature 70^oC, 80^oC and 90^oC.
   (viii) Record your results as in Table 1.

Table 1

Experiment	**Temperature**	**Time (s)**
1	60^oC	
2	70^oC	
3	80^oC	
4	90^oC	

Questions:

(a) Write a balanced chemical equation for reaction between **UU** and **ZO**.

(b) What is the product which causes the solution to cloud the letter **X**?

(c) Plot a graph of temperature against time (s).

(d) What conclusion can you draw from you graph?

(15 marks)

3. Substance **V** is a simple salt which contains one cation and one anion. Carry our the experiments described below. Record carefully your observations and make appropriate inferences and hence identify the anion and cation present in sample **V**.

S/n	**Experiment**	**Observation**	**Inference**
1	Observe the appearance of sample **V**.		
2	Put a little amount of sample **V** in a test tube then add water and shake.		
3	Heat a little amount of **V** in a dry test tube.		
4	To a little sample **V** in a test tube add dilute Hydrochloric acid. Add more of the acid until the test tube is half full. Divide the resulting solution into three portions and add the following:		
	a) To the one portion add NaOH solution drop wise then excess.		
	b) To the second portion add ammonia solution drop wise then in excess.		
	c) To the third portion add ammonium oxalate solution.		
5	Perform flame test.		

Conclusion

(i) The cation in sample **V** is ____.

(ii) The anion in sample **V** is ____.

(iii) The chemical formula of **V** is ____.

(iv) The name of compound **V** is ____.

(15 marks)

2011 - CHEMISTRY 2A ACTUAL PRACTICAL A

1. You are provided with the following:
   AA A solution of 0.2 M nitric acid (HNO_3);
   BB A solution of 4.2 g $NaxCO_3$ per 0.5 dm^3 of solution;
   MO is methyl orange indicator.

   Procedure

   Put solution **AA** into the burette. Pipette 20 cm^3 or 25 cm^3 of solution **BB** in a titration flask. Add two drops of methyl orange indicator into the titration flask. Titrate solution **BB** against **AA** until the end point is reached. Record the burette reading. Repeatthe procedure to obtain three more readings and record your results in a tabular form.

   Questions:

   (a) (i) Calculate the average titre volume.
   (ii) Summary: ____ cm^3 of solution **BB** required ____ cm^3 of solution **AA** for complete reaction.

   (b) If the mole ratio for the reaction is 1:1 find:
   (i) Concentration of $NaxCO_3$ in mol/dm^3 and g/dm^3.
   (ii) Molecular mass of $NaxCO_3$.
   (iii) Atomic mass of x and replace it in the formula $NaxCO_3$.

   (c) Write a balanced chemical equation for the reaction in this experiment.

   (d) What is the significance of the indicator in this experiment?

   (e) Why is there a colour change when enough acid has been added to the base?

   (20 marks)

2. Your are provided with the following materials:
   TT: A solution of 0.13 M $Na_2S_2O_3$ (sodium thiosulphate); **HH**: A solution of 2 M HCl; Distilled water; Stopwatch.

   Procedure:

   (i) Using 10 cm^3 measuring cylinder, measure 20 cm^3 of solution **TT** and put into 100 cm^3 beaker.

   (ii) Use different measuring cylinder to measure 10 cm^3 of **HH** and pour it into the beaker containing solution **TT**, immediately start the stop watch. Swirl the beaker twice.

   (iii) Place the beaker with the contents on top of a piece of paper marked **X**.

   (iv) Look down vertically through the mouth of the beaker so as to see the cross at the bottom of the beaker. Stop the clock when the letter **X** is invisible.

   (v) Record the time taken for the letter **X** to disappear completely.

   (vi) Repeat the experiment as shown in Table 1.

   (vii) Record your results in tabular form as shown in Table 1.

   Table 1: Table of results

Exp. No.	Vol. of **HH** (cm^3)	Vol. of **TT** (cm^3)	Vol. of Distilled water (cm^3)	Time (sec)	$\frac{1}{t}$ (s^-1)
1	10	20	0		
2	10	15	5		
3	10	10	10		
4	10	5	15		

Questions:

(a) Complete filling the table of results (Table 1).

(b) Write a balanced equation for reaction between **TT** and **HH**.

(c) What is the reaction product which causes the solution to cloud the letter **X**?

(d) How was the factor of concentration varied in this experiment?

(e) Plot a graph of 1/t against the volume of the thiosulphate.

(f) Use the graph to explain how variation of concentration affects the rate of chemical reaction.

(15 marks)

3. Sample **S** is a simple salt containing one cation and one anion. Carry our the experiments described below. Record your observations and inferences as shown in Table 2.
Table 2: Experimental results

S/n	**Experiment**	**Observation**	**Inference**
(a)	Observe the appearance of sample **S**.		
(b)	Place a spoonful of sample **S** in a test tube, add water and shake to dissolve.		
(c)	Put a spatulaful of sample **S** in a test tube and heat.		
(d)	Add three drops of sodium hydroxide solution to the solid sample in a test tube.		
(e)	Put a spatulaful of sample **S** in a dry test tube and add concentrated sulphuric acid. Warm the mixture and test for any gas evolved.		
(f)	Put a spatulaful of sample **S** in a dry test tube and add concentrated sulphuric acid and manganese dioxide. Warm the mixture and test for any gas evolved.		
(g)	To a portion of the solution from (f) add aqueous silver nitrate followed by aqueous ammonia.		

Conclusion:

(a) The cation present in **S** is ____ and the anion is ____.

(b) The name of sample **S** is ____.

(c) Write a balanced chemical equation for the reactions taking place in experiments (c) and (d).

(15 marks)

2010 - CHEMISTRY 2A ALTERNATIVE A PRACTICAL

1. You are provided with the following:
   Solution **D** containing 6.90 g of T_2CO_3 per 0.50 dm^3 of solution
   Solution **N** containing 1.55 g of hydrochloric acid per 200 cm^3 of solution
   Methyl orange indicator solution.

   Procedure:
   Put solution N in the burette. Pipette 20 cm^3 or (25 cm^3) of D into a titration flask. Add a few drops of methyl orange indicator. Titrate solution N from the burette against solution D in the titration flask to the end point. Note the burette reading. Repeat the procedure to obtain three more values and record the results as shown in the following table.

   (a) Table of results

   (i) Burette readings

Titration	**Pilot**	**1**	**2**	**3**
Final reading (cm^3)				
Initial reading (cm^3)				
Volume used (cm^3)				

   (ii) The volume of the pipette used was ____ cm^3.

   (iii) The volume of the burette used was ____ cm^3.

   (iv) ____ cm^3 of solution D required ____ cm^3 of solution N for complete reaction.

   (v) The colour change at the end point was from ____ to ____.

   (b) Write a balanced equation for the above neutralization reaction.

   (c) Calculate the following:

   (i) molarity of acid solution N.

   (ii) molarity of the base solution D.

   (iii) molar weight of T_2CO_3.

   (iv) atomic mass of element T.

   (d) Identify element T in T_2CO_3

(25 marks)

2. Sample **B** is a simple salt containing **one** cation and **one** anion. Carry out the experiments described in the following table carefully and record all your observations and appropriate inferences. Identify the cation and anion present in sample **B**.

	Experiment	Observation	Inference
(a)	Appearance of sample B.		
(b)	Put a spatulaful of sample B in a test-tube. Add water until half test-tubeful. Stir and divide the solution into five portions in different test tubes and do the following:		
	(i) add fresh zinc metal granules to the first portion. Heat for a while. Decant the result. Pour the solid material onto a filter paper and observe. Let it dry, then observe again.		
	(ii) add NaOH solution until excess to the second portion then heat and observe again.		
	(iii) add ammonia solution dropwise to the third portion until excess.		
	(iv) add $AgNO_3$ to the fourth portion followed by dil. HNO_3.		
	(v) add $AgNO_3$ to the fifth portion followed by ammonia solution.		

Conclusion

(a) The cation present in sample B is ____ and the anion is ____.

(b) What has been happening in the experiments (b)(i) and (b)(ii)? Use reaction equations where possible.

(25 marks)

3. Substance Z contains **one** basic radical and and **one** acidic radical. Using systematic qualitative analysis procedures carry out experiments on sample Z and make appropriate observations and inferences to identify the radicals.

Experiment	Observation	Inference

Conclusion

The Basic radical in sample Z is ____ and the acidic radical is ____.

(25 marks)

2009 - CHEMISTRY 2A ALTERNATIVE A PRACTICAL

1. You are provided with the following solutions:
   Solution WW containing 4.38 g of pure hydrochloric acid per dm^3 of solution.
   Solution ZZ containing 14.30 g of hydrated sodium carbonate [Na_2CO_3 .x H_2O] per dm^3.
   Methyl orange indicator.

   Procedure:
   Put solution WW in the burette. Pipette 20 cm^3 or (25 cm^3) of solution ZZ into a titration flask. Add about three to four drops of methyl orange indicator into the titration flask. Titrate solution WW against solution ZZ until the end point is reached. Note the burette reading. Repeat the procedure to obtain three more readings. Record your results as shown in the following Table.

   (a) Table of results

   (i) Burette readings

Titration	**Pilot**	**1**	**2**	**3**
Final reading (cm^3)				
Initial reading (cm^3)				
Volume used (cm^3)				

   (ii) The volume of pipette used was ____ cm^3.

   (iii) The volume solution WW needed for complete neutralization was ____.

   (iv) The colour change at the end point was from ____ to ____.

   (b) Write a balanced chemical equation for the reaction between solution ZZ and WW.

   (c) Calculate the molarity of

   (i) solution WW

   (ii) solution ZZ.

   (d) Calculate the value of x in the formula (Na_2CO_3 .x H_2O).

(25 marks)

2. Sample M is a simple salt containing **one** cation and **one** anion. Carry out carefully the experiments described in the following table. Record all your observations and appropriate inferences to identify the ions present in M.

S/N	Experiment	Observation	Inference
(a)	Appearance of sample M		
(b)	Place a spatulaful of sample M in a test-tube and heat while rotating the tube. Test for any gas(es) evolved.		
(c)	Place a spatulaful of sample M in a test tube and add dilute hydrochloric acid. Test for any gas(es) evolved. Add more of the acid until the test tube is half full. Divide the solution into three portions and then do the following; (i) add sodium hydroxide solution dropwise and then in excess to the first portion.		
	(ii) add a few drops of potassium iodide solution to the second portion.		
	(iii) add ammonium hydroxide solution dropwise till excess to the third portion.		

Conclusion
The cation in sample M is ____ and the anion is ____.

(25 marks)

3. Substance V is a simple salt containing **one** cation and **one** anion. Using systematic qualitative analysis procedures carry out tests on sample V and make appropriate observations and inferences to identify the cation and anion in V. Record your experiments, observations and inferences as shown in the following table:

Experiment	Observation	Inference

Conclusion
The cation in sample V is ____ and the anion is ____.

(25 marks)

2008 - CHEMISTRY 2A ALTERNATIVE A PRACTICAL

1. You are provided with the following:

   Solution M containing 9.0 g of H_2X per dm^3 of the solution.

   Solution N containing 4.91 g of sodium hydroxide per dm^3 of the solution.

   Solution P is phenolphthalein indicator.

   Procedure

   Put solution M into the burette. Pipette 25 cm^3 (or 20 cm^3) of solution N into the titration flask. Put two to three drops of P into the titration flask. Titrate solution M from the burette against solution N in the titration flask until a colour change is observed. Note the burette reading. Repeat the procedure to obtain three more readings. Record your results as shown in Table 1.

   Results

   Table 1: Burette readings

Titration	**Pilot**	**1**	**2**	**3**
Final reading (cm^3)				
Initial reading (cm^3)				
Volume used (cm^3)				

   (a) Give the volume of the pipette used.

   (b) Give the volume of solution M needed for complete neutralization of solution N.

   (c) Tell the colour change of the indicator at the end point of the titration.

   (d) Write the balanced chemical equation for the reaction between solution M and N.

   (e) Calculate the

      (i) molarity of solution M

      (ii) molar mass of H_2X

      (iii) mass of X in H_2X.

(25 marks)

2. Sample D is a simple salt containing one cation and one anion. Carry out carefully the experiments described below recording all your observations and appropriate inferences as shown in Table 2 to identify the cation and anion present in D.

Table 2

	Experiment	Observation	Inference
(a)	Observe the appearnce of salt D.		
(b)	Put a little solid sample D in a clean and dry test tube and heat.		
(c)	Put a spatulaful of sample D in a test tube, add distilled water, stir and divide the obtained solution into four portions in different test tubes. To the		
	(i) first portion of the solution of sample D in a state tube add aqueous ammonia slowly till excess.		
	(ii) second portion of the solution of sample D in a test tube add aqueous ammonia slowly till excess.		
	(iii) third portion of the solution of sample D in a test tube add potassium hexacyanoferrate (II).		
	(iv) fourth portion of the solution of sample D in a test tube add dilute HCl followed by $BaCl_2$ solution.		

Conclusion:

The cation in sample D is ____ and the anion is ____.
The molecular formula of salt D is ____.

(25 marks)

3. Sample Y is a simple salt containing one anion and one cation. Using systematic qualitative analysis procedures carry out tests on sample Y and make appropriate observations and inferences to identify the cation and anion present in sample Y.

Experiment	Observation	Inference

Conclusion:

The cation in sample Y is ____ and the anion is ____.

(25 marks)

2007 - CHEMISTRY 2A ALTERNATIVE A PRACTICAL

1. You are provided with the following:

   Solution **G** containing 0.05 M sulphuric acid.
   Solution **H** containing 2 g of **X**OH in 500 cm^3 of the solution.
   Solution **F**, methyl orange indicator.

   Determine the atomic mass of **X** in **X**OH.

   Procedure:

   Put solution **G** in the burette. Pipette 20 cm^3 or (25 cm^3) of solution **H** into the conical flask. Add two or three drops of methyl orange indicator. Titrate solution **H** against solution **G** from the burette until a colour change is observed. Note the burette reading. Repeat the procedure to obtain three more readings.

   (a) Record your results in a table as shown below.

   (i) Burette readings.

Titration	**Pilot**	**1**	**2**	**3**
Final reading (cm^3)				
Initial reading (cm^3)				
Volume used (cm^3)				

   (ii) The volume of pipette used was ____ cm^3.

   (iii) Calculate the mean titre volume.

   (iv) The volume of solution **H** needed for complete neutralization of ____ cm^3 of solution **G** was ____ cm^3.

   (b) Write a balanced chemical equation for the reaction between solution **G** and **H**.

   (c) Calculate the

   (i) molarity of **H**.

   (ii) concentration of **H** in g/dm^3.

   (iii) molar mass of **X**OH.

   (iv) atomic mass of **X** in compound **X**OH.

(25 marks)

2. Sample **B** is a simple salt containing one cation and one anion. Carry out carefully the experiments described below and record all your observations and appropriate inferences. Identify the cation and anion present in sample **B**.

Experiment	Observation	Inference
(a) Appearance of sample **B**.		
(b) To half a spatulaful of sample **B** in a test tube, add concentrated sulphuric acid and warm.		
(c) To a spatulaful of sample **B** in a test tube, add 10 cm^3 of distilled water and stir to obtain a stock solution them divide into three portions.		
(d) To the first portion of the stock solution, add sodium hydroxide till excess.		
(e) To the second portion of the stock solution, add barium chloride solution.		
(f) To the third portion of the stock solution, add freshly prepared acidified ferrous sulphate solution followed by concentrated sulphuric acid added slowly along the walls of the test tube.		
(g) Perform a flame test on sample **B**.		

Conclusion
The cation in sample **B** is ____ and the anion is ____.

(25 marks)

3. Sample **N** is a simple salt containing one cation and one anion. Using systematic qualitative analysis procedures, carry out tests on the sample and make appropriate observations and inferences to identify the cation and anion in sample **N**.

Experiment	Observation	Inference

Conclusion
The cation present in **N** is ____ and the anion is ____.

(25 marks)

Physics Past Papers

2013 - PHYSICS 2A ACTUAL PRACTICAL A

1. You are provided with a metre rule, a knife edge, two strings of length 100 cm each and two weights W_1 and W_2 of masses 50 g and 100 g respectively. Proceed as follows:

   (a) Balance a metre rule on a knife edge, put a mark and write G at the balancing point using a piece of chalk or a pencil. Measure and record the length l, width w and thickness t of a metre rule using a vernier caliper.

   (b) Place the metre rule on a knife edge so that the knife edge is at 60 cm of your metre rule (see Figure 1 (a)). Suspend weight W_2 of 100 g on the right hand side of the knife edge. Adjust W_2 until the metre rule balances horizontally. Read and record lengths 'b' and 'c' as seen in Figure 1 (a).

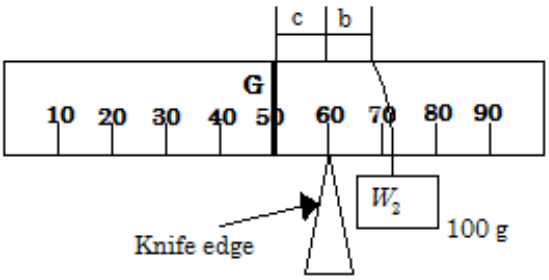


Figure 1 (a)

   (i) Suspend weight W_1 of 50 g on the left hand side of the knife edge at the position 47 cm and adjust weight W_2 until the metre rule balances horizontally as seen in Figure 1 (b). Read and record the lengths 'a' and 'b'.

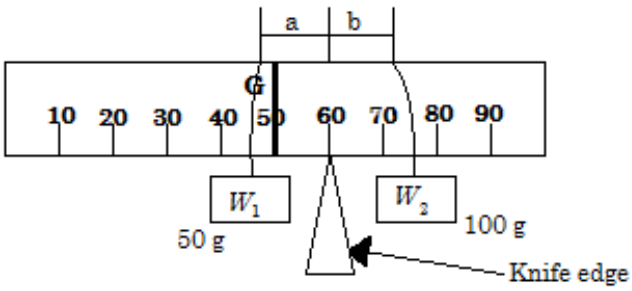


Figure 1 (b)

   (ii) Repeat the procedures in (b) (i) by adjusting the position of W_1 to the left at the interval of 3 cm to obtain other four (4) readings.

   (c) Tabulate your results as shown in Table 1.

   Table 1

a (cm)	b (cm)

   (d) Plot a graph of "b" against "a".

   (e) What is the nature of the graph?

   (f) Calculate the slope S of the graph.

   (g) (i) Read the b-intercept, given that $b = Sa + \dfrac{W}{W_2} \times c$

   (ii) What does $\left(\dfrac{W}{W_2}\right) c$ represent in your graph?

(iii) Calculate the value of W using the relation $W_2 = \dfrac{Wc}{9.5\text{cm}}$. What does W represent?

(h) (i) Find the value of the ratio $P = \dfrac{l \times w \times t}{m}$.

Note: The mass m of a meter rule can be obtained by calculations.

(ii) What is the physical meaning of the value of P?

(i) State a possible source of error in this experiment.

(j) How can you minimize error in 1 (i)?

(k) State the aim of this experiment.

(25 marks)

2. You are provided with a Plane mirror, a Ruler, Protract, Drawing board, Optical pins, Office pins and Plain papers. Proceed as follows:

(a) On the plain paper provided, draw a line 13 cm from the top of the paper and call it M_1M_2. Pin your paper on the board provided and place the reflecting surface of the mirror along the line M_1M_2 as seen in Figure 2.

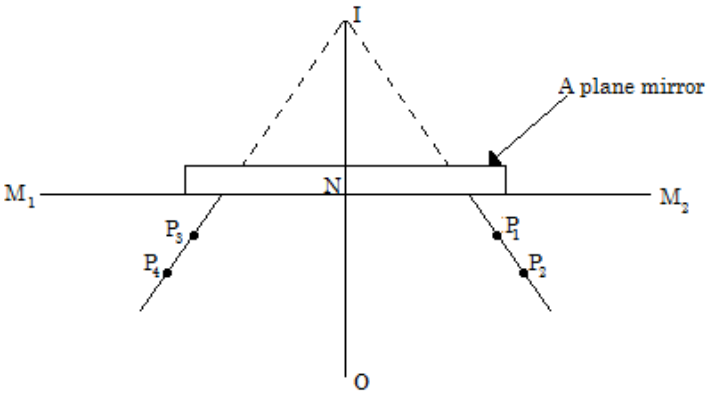


Figure 2

(b) Insert pin O as an object at 4.0 cm in front of the mirror. Place pins P_1 and P_2 so as to appear in one straight line with the image of object O seen in the plane mirror.

(c) Remove pins P_1 and P_2, using other pins, place pins P_3 and P_4 so as to appear in a straight line with the image of object O in the other side (see Figure 2).

(d) Remove the mirror and pins. Draw lines joining P_1 and P_2 on one side and the other joining P_3 and P_4 on the other side of object O, extend both lines to meet at I on the other side of line M_1M_2.

(e) Join OI, a line cutting the reflecting surface at N.

(f) Repeat this procedure for the distance of an object being 6, 8, 10 and 12 cm.

(g) On all the diagrams drawn:

(i) Measure the distance ON and NI.

(ii) Comment on the distances obtained in 2 (g) (i).

(iii) What is the nature of image? Give reasons for your answer.

(iv) State four characteristics of the image you obtained.

(v) What is the aim of this experiment?

(vi) Mention and state the law governing this experiment.

(vii) Explain a source of error in this experiment.

(viii) How can you minimize the error in (vii) above?

Note: The papers used for drawing should be attached and collected together with answer booklets.

(25 marks)

2012 - PHYSICS 2A ACTUAL PRACTICAL A

1. You are provided with a measuring cylinder, eureka can, nylon thread, standard masses and water. Proceed as follows:

   (a) Pour water into eureka can until it is just beginning to overflow.

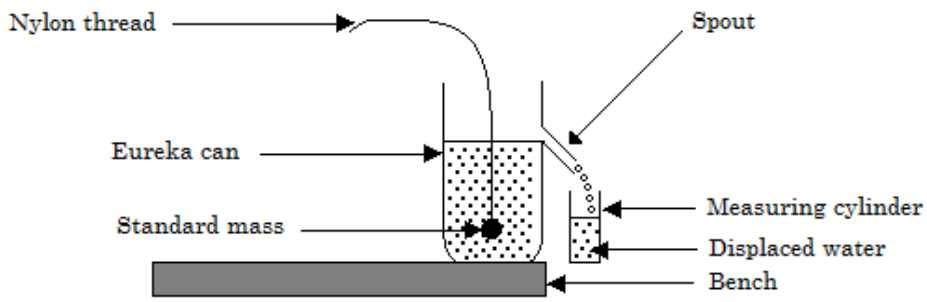


Figure 1

   (b) Hold a suitable measuring cylinder under the spout and immerse a standard mass of 50 g into eureka can as shown in Figure 1. Water will pass through the spout and will be collected by the measuring cylinder. Wait for it to drop until it starts to cease and take long interval to drop. Record the reading of the water collected.

   (c) Repeat the procedures in 1 (b) for standard masses of 100 g, 150 g, 200 g and 250 g.

   (d) Tabulate your results showing the quantities as follows:

Mass (g)	Volume (cm^3)	Mass ÷ Volume (g/cm^3)
50		
100		
150		
200		
250		

   (e) Plot a graph of mass against volume.

   (f) State the nature of the graph.

   (g) From the graph:

      (i) Calculate the slope.

      (ii) What does the slope of the graph show?

      (iii) What is the relationship between mass and volume?

      (iv) Establish the formula governing the experiment.

   (h) Identify with reasons the best to the least satisfactory method of finding the constant value of mass divide by volume.

   (i) State two possible errors in this experiment.

   (j) How can you minimize errors in 1 (i)?

2. You are provided with two plane mirrors, an optical pin, a sheet of plane drawing paper, mirror holder or office pins, a protractor, a ruler and a drawing table. Proceed as follows:

   (a) Draw two lines at right angles.

   (b) Place the two plane mirrors along the top two lines using the mirror holders or office pins as shown in Figure 2.

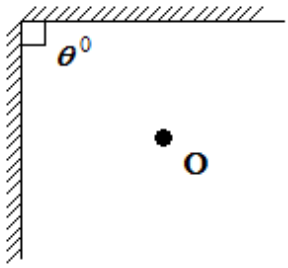


Figure 2

   (c) Put an optical pin at O when $\theta = 90°$. Look onto one of the mirrors and count the number of images, n, you see.

   (d) Repeat the procedures in 2 (c) for $\theta = 72°$, $\theta = 60°$, $\theta = 45°$ and $\theta = 30°$.

   (e) Tabulate your results for the values of θ, n and $\frac{360°}{\theta}$.

   (f) Plot a graph of number of images, n, against $\frac{360°}{\theta}$.

   (g) From the graph:

      (i) Determine the slope.

      (ii) Find the number of images when $\frac{360°}{\theta} = 9$

      (iii) Find the value of the y-intercept.

      (iv) Derive the equation relating the number of images and $\frac{360°}{\theta}$

   (h) From your experiment:

      (i) What happens to the number of images as the value angle θ is reduced?

      (ii) What happens to the number of images when $\theta = 0°$?

   (i) State a possible source of error and how you can minimize it.

   (j) What is the aim of this experiment?

2011 - PHYSICS 2A ACTUAL PRACTICAL A

1. The aim of this experiment is to determine the mass of a given dry cell size "AA". Proceed as follows:

   (a) Locate and note the centre of gravity C of the metre rule by balancing it on the knife edge.

   (b) Suspend the 50 g mass at length 'a' cm on one side of the metre rule and the 20 g mass together with the dry cell at length 'b' cm on the other side of the metre rule. Fix the 50 g mass at length 30 cm from the fulcrum and adjust the position of the 20 g mass together with the dry cell until the metre rule balances horizontally. Read and record the values of a and b as a_0 and b_0 respectively.

   (c) Draw the diagram for this experiment.

   (d) By fixing $a = 5$ cm from fulcrum C, find its corresponding length b.

   (e) Repeat the procedure in (d) above for $a = 10$ cm, 15 cm, 20 cm and 25 cm. Tabulate your results.

   (f) Draw a graph of 'a' against 'b' and calculate its slope G.

   (g) Calculate X from the equation $50 = \frac{b_0}{a_0}(20 + X)$.

   (h) Comment on the value of $\frac{b_0}{a_0}$.

   (i) Sate the principle governing this experiment.

(25 marks)

2. You are provided with an ammeter, A, resistance box, R, dry cell, D, a key, K and connecting wires. Proceed as follows:

   (a) Connect the circuit in series.

   (b) Put $R = 1\ \Omega$ and quickly read the value of current I on the ammeter.

   (c) Repeat procedure (b) above for $R = 2\ \Omega$, $3\ \Omega$, $4\ \Omega$ and $5\ \Omega$. Record your results in a tabular form.

   (d) Draw the circuit diagram for this experiment.

   (e) Plot the graph of R against $\frac{1}{I}$.

   (f) Determine the slope of the graph.

   (g) If the graph obeys the equation $R = \frac{E}{I} - r$, then

      (i) suggest how E and r may be evaluated from your graph.

      (ii) compute E.

      (iii) compute r.

   (h) State one source of error and suggest one way of minimizing it.

   (i) Suggest the aim of this experiment.

(25 marks)

2010 - PHYSICS 2A ALTERNATIVE A PRACTICAL

1. The aim of this experiment is to find the mass of the unknown load labeled "W" and the spring constant K. Proceed as follows:

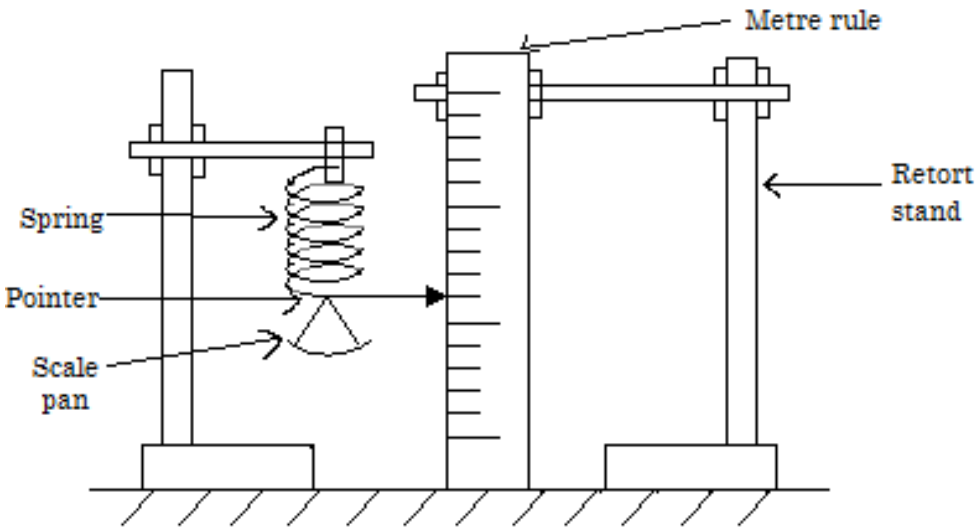


Figure 1

Set up the apparatus as shown in Figure 1. Put a mass of 50 g on the scale pan and record the equilibrium position X_0 of the pointer. Put on the scale pan the unknown weight marked W. Without removing W and the 50 g mass in the scale pan, add a load L of 50 g and record the new position of the pointer X. Calculate the extension $E = (X - X_0)$. Repeat this process for L = 100 g, 150 g, 200 g and 250 g.

(a) Record you conclusions as shown in Table 1.

Equilibrium position X_0.....

Table 1

Load (g)	X (cm)	$E = X - X_0$ (cm)
50		
100		
150		
200		
250		

(b) Plot the graph of load L against absolute value of extension E. The scale of the vertical axis should be chosen to range from 200 g to 300 g.

(c) From the graph, determine the unknown weight marked W, given that L = KE + W where K is a constant.

(d) What does the gradient of the graph represent?

(e) State the sources of errors and precautions that should be taken in the experiment.

(25 marks)

2. The aim of this experiment is to determine the refractive index of water. Proceed as follows:

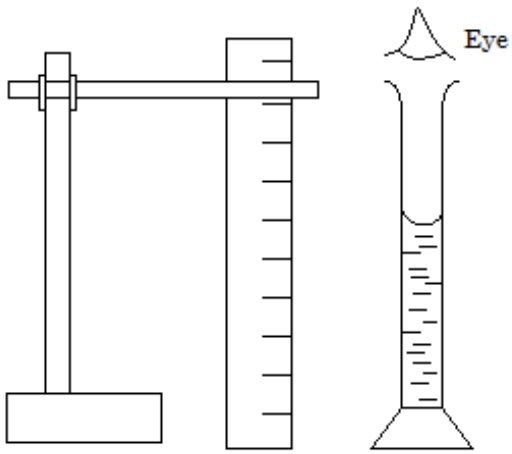


Figure 2

(a) Arrange your apparatus as in Figure 2. Put about 150 cm^3 of clear water in the measuring cylinder. Drop an office pin at the bottom so that it rests touching the wall of the cylinder.

(b) Look in the cylinder from Figure 2. Use another office pin as a search pin, move it up and down outside the cylinder, and locate the image position by no parallax method. Locate the image position of the ruler. Measure and record the depth (H_1) of the image. Measure and record the depth (H_2) of water. Repeat the experiment with 175 cm^3, 200 cm^3, 225 cm^3 and 250 cm^3 of water in the measuring cylinder.

(c) (i) Record in Table 2 your values of H_1 and H_2 corresponding to the volumes of water in the measuring cylinder.

Table 2

Volume of water V (cm)	H_1	H_2
150		
175		
200		
225		
250		

(ii) Plot the graph of H_2 versus H_1.

(iii) Determine the slope of the graph.

(iv) What is the physical meaning of the slope?

(v) State sources of error in this experiment.

(25 marks)

3. The aim of this experiment is to determine the resistivity of an electrical conductor P.

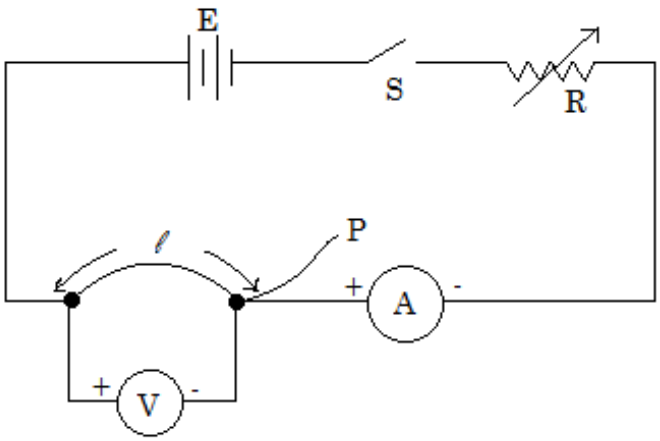


Figure 3

With P having a length $l = 50$ cm, connect up the circuit as shown in Figure 3. Close one key S and adjust the rheostat R so that the current in P is 0.20 A. Record the current I and the potential difference V between its ends.

Repeat the procedure with current I =0.30 A, 0.40 A, 0.50 A and 0.60 A.

(a) Record your results in Table 3.

Table 3

Current I (A)	P.d. (volts)
0.20	
0.30	
0.40	
0.50	
0.60	

(b) Plot a graph of V against I and calculate the slope G.

(c) Deduce the resistivity of the conductor P given that; $\rho = \dfrac{G\pi d^2}{4l}$.

Where ρ = resistivity
d = diameter of P (measured using the micrometer screw gauge provided).

(25 marks)

2009 - PHYSICS 2A ALTERNATIVE A PRACTICAL

1. In this experiment you are required to find the relationship between the length of a simple pendulum and its period. Proceed as follows:

   (a) Suspend a simple pendulum of length L = 100 cm. Displace the pendulum through a small angle so that it swings parallel to the edge of the bench or table, determine the time for 20 oscillations. Continue reducing the length of the pendulum by 10 cm each time and obtain a total of six readings.

   (b) Record your readings in a table as shown below.

Length of pendulum L (cm)	Log_{10}L	Time for 20 oscillations	Period T	Log_{10}T

   Assuming that $T \propto L^a$, we have $T = kL^a$ and taking logarithms to base ten on both sides we get $\log_{10}T = a\log_{10}L + \log_{10}k$.

   (i) Plot a graph of $\log_{10}T$ (vertical axis) against $\log_{10}L$ (horizontal axis) hence determine the values of a and k each correct to one decimal place.

   (ii) From your answer in (i) above write down the values of a and k each in the form of $\frac{b}{c}$ where b and c are integers (i.e. whole numbers).

   (iii) From the assumption and your answer in (ii) deduce the form of the equation governing the motion of the simple pendulum.

(25 marks)

2. The aim of this experiment is to determine the refractive index η of a given glass block.

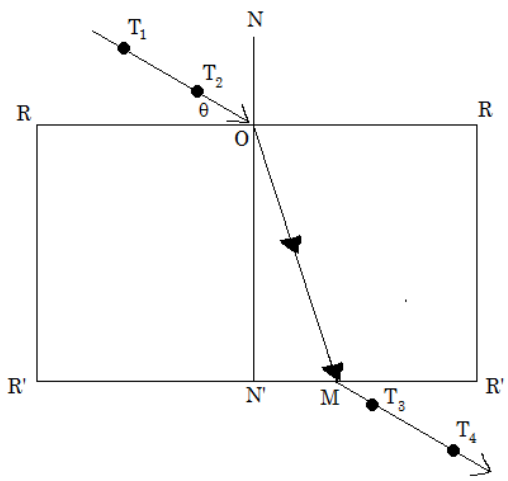


Figure 1

Place the rectangular glass block on the white paper on a drawing board. Using a pencil trace the outline of the block. Remove the glass block and draw a normal NOM near the left end of the block (Figure 1).

Using a protractor and a pencil measure $\theta = 20°$, draw a line making the angle 20° with the surface RR of the block. Erect two pins T_1 and T_2 on this line and at a suitable distance from one another. Return the block and erect the pins T_3 and T_4 at positions such that they lie in a straight line with pins T_1 and T_2 as seen through the block. Now remove the block and draw a complete path of the ray (Figure 1).

Measure the length MN′ and ON′: Repeat the procedure for values of $\theta = 30°, 40°$ and 60° respectively. In each case make a drawing on a fresh part of the drawing paper.

(a) Record the values of θ, MN′, ON′, $\frac{\text{MN}'}{\text{ON}'}$ and $\cos\theta$ in a tabular form.

(b) Plot a graph of $\frac{\text{MN}'}{\text{ON}'}$ against $\cos\theta$.

(c) Find the slope G of the graph.

(d) Calculate the value of the refractive index η; given that $G = \frac{1}{\eta}$.

(e) State two sources of errors. **(25 marks)**

3. The aim of this experiment is to verify Ohm's Law.

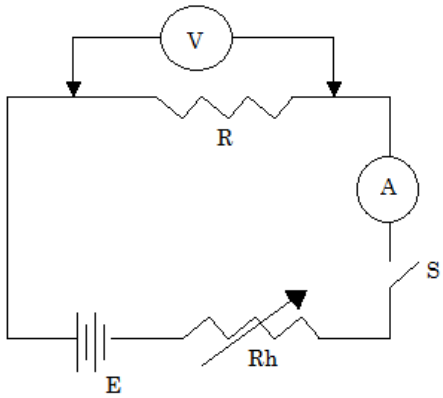


Figure 2

(a) Set up the apparatus as shown on Figure 2, close switch S. Adjust the Rheostat Rh by sliding slowly from one end, read and record the value V of the voltmeter and current I of the ammeter.

(b) Repeat the experiment by changing the Rheostat slider to obtain about five pair of readings.

NB: Adjust the Rheostat until when the pointer is exactly on the division of the metre scale.

Table of results

V (V)							
I (A)							

(c) Plot a graph of V (vertical axis) against I (horizontal axis).

(d) (i) Find the slope of the graph.

(ii) What is the relation between V and I?

(iii) Find the resistance R. **(25 marks)**

2008 - PHYSICS 2A ALTERNATIVE A PRACTICAL

1. The aim of this experiment is to investigate whether string A obeys Hooke's law.

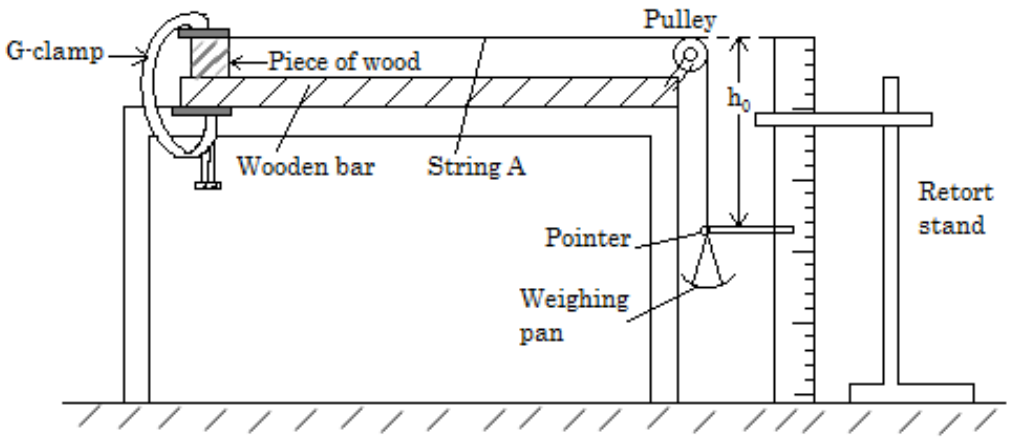


Figure 1

Proceed as follows:

Clamp string A at one end, attach a weighing pan at the other end and a pointer to give a reading on a scale as shown in figure 1 above.

Measure the height, h_0 when the pan is empty.

Place 50 g mass on the pan and record the new height h indicated by the pointer.

Add another 50 g mass each time up to 300 g, and record the corresponding values of h for added mass.

(a) Tabulate your results as shown in the table below.

$h_0 =$ _____ cm.

Mass, m (g)	Height, h (cm)	Extension $(h - h_0)$ (cm)	Stretching force, F (N)
50			
100			
150			
200			
250			
300			

(b) Plot a graph of force F (N) against extension (cm).

(c) From the graph find the

(i) slope, K of the graph.

(ii) extension caused by a mass of 180 g.

(d) Deduce whether string A obeys Hooke's law.

(e) State the law. **(25 marks)**

2. You are provided with a glass block, four sheets of drawing paper, four optical pins (or office pins) and a drawing board.

Proceed as follows:

Place the glass block flat on the drawing paper fixed to the drawing board and with a sharp pencil, draw its outline.

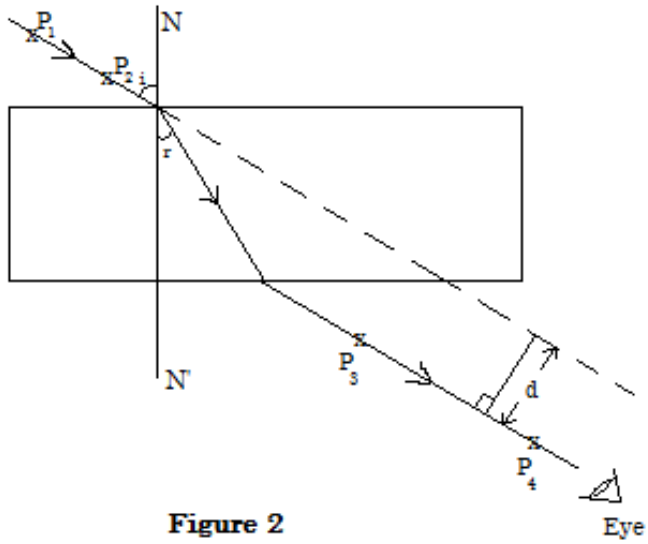


Figure 2

Remove the glass block and draw a normal NN' to the longer edge of the block (see fig. 2). Draw a line making an angle of incidence (i) of 30°. Stick two vertical pins P_1 and P_2 on this line. Replace the glass block. Stick two more pins P_3 and P_4 on the other side of the block so that they appear to be in the same straight line with the images of pins P_1 and P_2 as seen through the block.
Remove the block and draw the complete path of the ray entering and leaving the block. Measure the angle of refraction (r).

Produce the incident ray as shown in fig. 2 and measure the perpendicular distance (d) between the incident ray and the emergent ray.
Repeat this procedure for angles of incidence of 40°, 50°, 60° and 70°. In each case draw the block again on a fresh part of the paper.

(a) Record your results in a table as follows:

$i°$	$r°$	d (cm)	$d \cos r$ (cm)	$\sin(i-r)$
30				
40				
50				
60				
70				

(b) Plot a graph of $d \cos r$ (vertical axis) against $\sin(i-r)$ (horizontal axis).

(c) Find the gradient of the graph.

(d) Measure the width of the glass block.

(e) How is the gradient related to the with of the glass block?

NB: Hand in your diagrams together with your answer booklet. **(25 marks)**

3. The aim of this experiment is to determine the resistance of a wire W.
Proceed as follows:

(a) Connect in series the full length of wire W of unknown resistance, battery B (3 V), a switch K, a rheostat Rh of a few ohms and an ammeter A of $0-1$ A.
Connect the voltmeter V of $0-3$ V across W. Check that the +ve side of the ammeter A and the +ve side of the voltmeter V are both on the +ve side of the battery B.

(b) Switch on the current. Adjust the rheostat to obtain five widely different values of V and corresponding values of current I.

(c) Tabulate your results as follows:

Potential difference V (volts)	Current I (amperes)

(d) (i) Draw a circuit diagram.

(ii) Plot a graph of potential difference V against current I.

(iii) Find the slope of the graph.

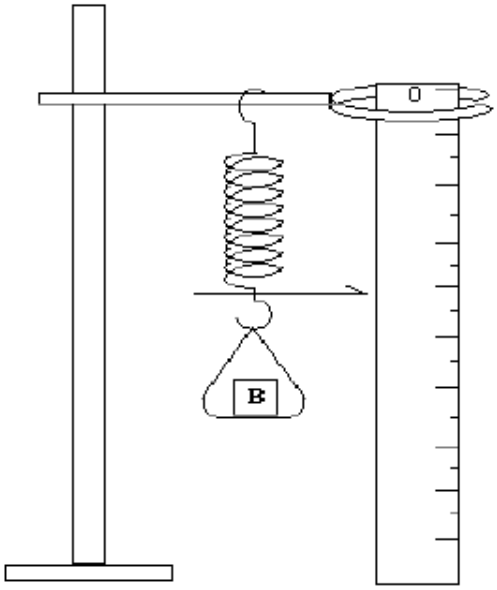
(iv) Determine the resistance of the wire W.

(v) Mention **two (2)** main precautions to be taken in this experiment.

(25 marks)

2007 - PHYSICS 2A ALTERNATIVE A PRACTICAL

1. The aim of this experiment is to determine the mass of a given object "B", and the constant of the spring provided.



(i) Set up the apparatus as shown in Fig. 1 with zero mark of the metre-rule at the top of the rule and record the scale reading by the pointer, S_0.

(ii) Place the object "B" and standard weight (mass) W equal to 20 g in the pan and record the new pointer reading S_1. Calculate the extension, $e = S_1 - S_0$ in cm.

(iii) Repeat the procedure in (ii) above with W = 40 g, 60 g, 80 g and 100 g.

(a) Record your results in tabular form as shown below:
Table of Results:

$S_0 =$			
Mass (kg)	Force, F (N)	Pointer reading S_1 (cm)	Extension $= S_1 - S_0$ (cm)
0			
0.02			
0.04			
0.06			
0.08			
0.10			

(b) Plot graph of Force F (vertical axis) against extension e (horizontal axis).

(c) Use your graph to evaluate

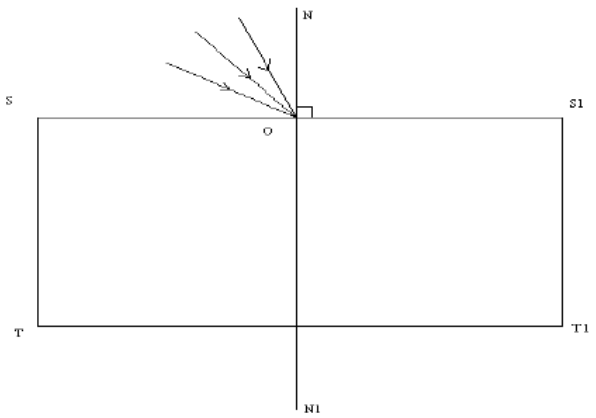
(i) mass of B

(ii) spring constant, K, given that force, extension, constant and weight of B are related as follows:
F = Ke - B

(25 marks)

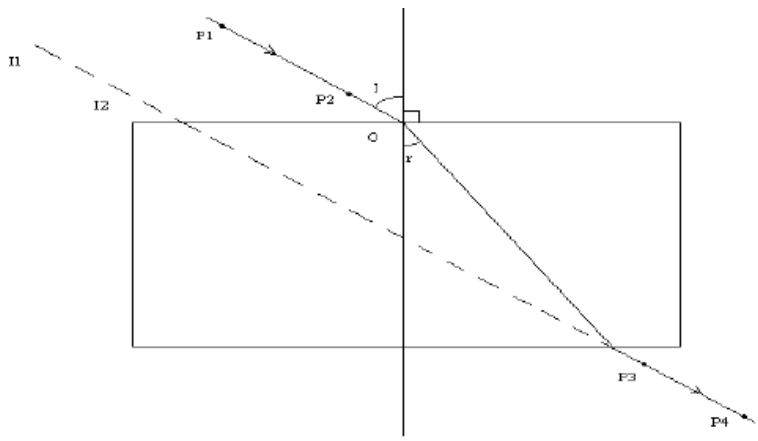
2. The aim of this experiment is to find the refractive index of a glass block. Proceed as following:

Place the given glass block in the middle of the drawing paper on the drawing board. Draw lines along the upper and lower edge of the glass block. Remove the glass block and extend the line you have drawn. Represent the ends of those line segments as SS^1 and TT^1. Draw the normal NN^1 to the parallel lines SS^1 and TT^1 as shown in Fig. 2(a).



Draw five evenly spaced lines from O to represent incident rays at different angles of incidence (10°, 20°, 30°, 40° and 50° from the normal). Replace the glass block carefully between SS^1 and TT^1. Stick two pins P_1 and P_2 as shown in Fig. 2(b) as far apart as possible along one of the lines drawn to represent an incident ray. Locate an emergent ray by looking through the block and stick pins P_3 and P_4 exactly in line with images I_1 and I_2 of pins P_1 and P_2. Draw the emergent ray and repeat the procedure for all the incident rays you have drawn. Finally draw in the corresponding refracted rays.

NOTE: The drawing paper should be handed in together with other answer sheets.



(a) Record the angles of incidence I and the measured corresponding angles of refraction "r" in a table. Your table of results should include the values of sin I and sin r.

(b) Plot the graph of sin I (vertical - axis) against sin r (horizontal - axis).

(c) Determine the slope of the graph.

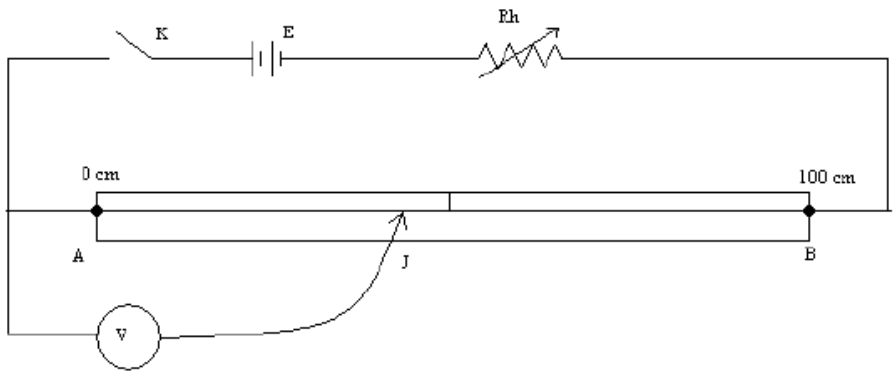
(d) What is the refractive index of the glass block used?

(e) Mention any sources of errors in this experiment.

(25 marks)

3. The aim of this experiment is to determine the potential fall along a uniform resistance wire carrying a steady current.

Proceed as follows:



Connect up the circuit as shown in Fig. 3. Adjust the rheostat so that when the sliding contact J is near B, and the key is closed the voltmeter V indicates an almost full scale deflection. Do not alter the rheostat again.
Close key K and make contact with J, so that AJ = 10 cm. Record the potential different V volts between A and J as registered on the voltmeter.
Repeat this procedure for AJ = 20 cm, 30 cm, 50 cm and 70 cm.

(a) Tabulate your results for the values of AJ and V.

(b) Plot a graph of V (vertical axis) against AJ (horizontal axis).

(c) Calculate the slope of the graph.

(d) What is your comment on the slope?

(e) State any precautions on the experiment.

(25 marks)

2006 - PHYSICS 2A ALTERNATIVE A PRACTICAL

1. In this experiment you are required to determine the mass of unknown object "X".

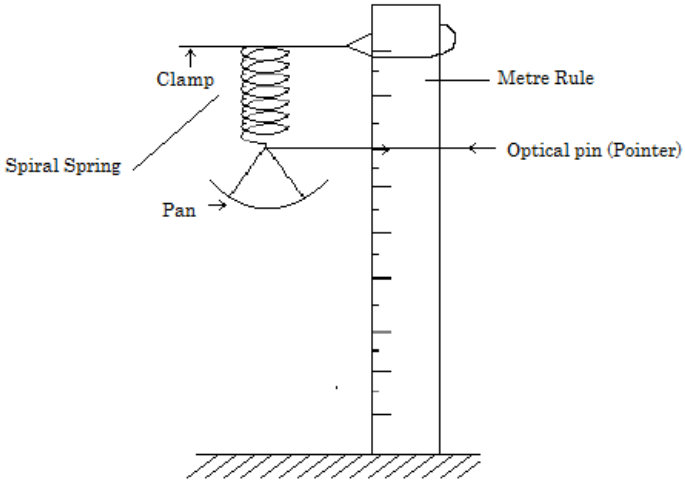


Fig. 1

Assemble the pieces of apparatus as shown in Figure 1, with zero mark scale of the rule at the lower most end.
Record the reading of the position of pointer on the scale of metre-rule when the pan is empty as S_0.
Put 20 g to the pan and record pointer reading S.
Find extension $e = S - S_0$ cm.
Repeat the procedure for mass of 40 g, 60 g, 80 g and 100 g. Put object X on the pan and record its pointer reading.

(a) Summarize your results in a table as follows:

Mass on pan (g)	20	40	60	80	100	X
Pointer reading (cm)						
Extension, $e = S - S_0$ (cm)						

(b) Plot graph of mas against extension (m Vs. e).
(c) Find slope, P, of your graph.
(d) Find mass X.
(e) Find Q, given that Q = P $\times$ e_x, where e_x is extension of X.
(f) Comment on Q and X.

2. Set up the experiment as shown in the diagram below using plane mirror, soft board, three pins and a white sheet of paper.

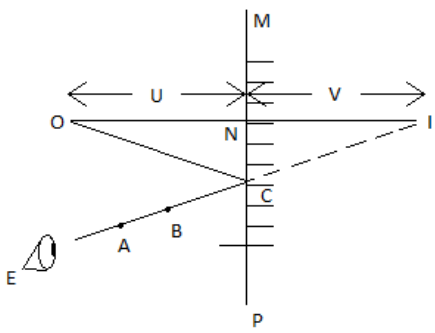


Fig. 2

Fix a white sheet of paper on the soft board. Draw a line across the width at about the middle of the white sheep (MP). Draw line ONI perpendicular to MP.
Fix optical pin O to make ON = U = 3 cm. By using plasticine or otherwise, fix plane mirror along portion of MP with O in front of the mirror. With convenient position of eye, E, look into the mirror and fix optical pins A and B to be in line with image, I, of pin O.
Measure and record NI = V. Repeat procedure for U = 6 cm, 9 cm and 12 cm.

(a) Tabulate your results as follows:

U (cm)	3	6	9	12
V (cm)				

(b) Plot graph of U against V.

(c) Calculate slope, m, of the graph to the nearest whole number.

(d) State relationship between U and V.

(e) Write equation connecting U and V using numerical value of m with symbols U and V.

(f) From your equation give position of the image when object is touching the face of the mirror.

3. You are required to determine the unknown resistance labeled X using a metre bridge circuit. Connect your circuit as shown below, where R is a resistance box, G is a galvanometer, J is a jockey and others are common circuit components.

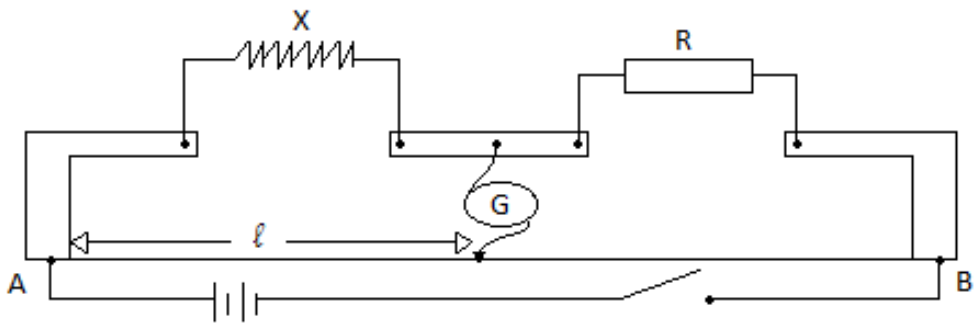


Fig. 3

Procedure:
With R = 1 Ω, obtain a balance point on a metre bridge wire AB using a jockey J. Note the length l in centimetres. Repeat the experiment with R equal to 2 Ω, 4 Ω, 7 Ω and 10 Ω. Tabulate your results for R, l and $^1/_l$.

(a) (i) Plot a graph of R (vertical axis) against $^1/_l$ (horizontal axis).
(ii) Determine the slope S of your graph.
(iii) Using your graph, find the value of R for which $^1/_l = 0.02$.

(b) Read and record the intercept R_0 on the vertical axis.

(c) Given that,

$$R = \frac{100X}{l} - X$$

Use the equation and your graph to determine the value of X.

(d) Comment on your results in (a)(iii), (b) and (c) above.

2005 - PHYSICS 2A ALTERNATIVE A PRACTICAL

1. The aim of this experiment is to determine the mass of unknown weight labelled **X** and the force constant of the spring **k**.

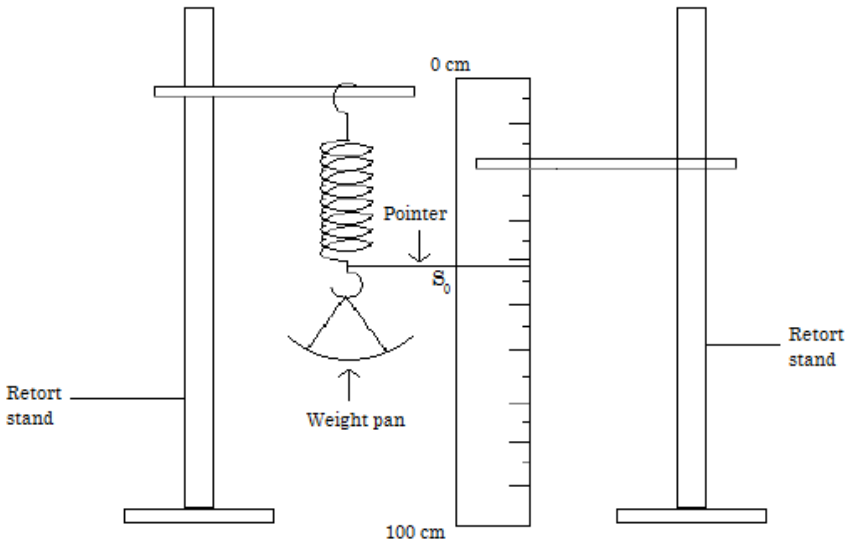


Figure 1

Set up the apparatus provided as shown in figure 1 above. Add 50 g mass on to the weight pan so that any "kinks" in the spring are removed. Leave this weight for the whole experiment but ignore it in all readings. Record the scale reading S_0. Add 50 g on to the weight pan and record the new scale reading S. Calculate the extension ($e = S - S_0$) caused by the weight. Repeat with different weights (W) to obtain at least five readings. Tabulate your results. Replace the weights (W) by the weight **X** provided and find the corresponding extension.

Record this extension as S_X ..... cm

(a) Plot a graph of load against extension.

(b) (i) Find the gradient (G) of your graph.

(ii) What is the physical meaning of the gradient?

(c) From the graph, what is the mass of the weight labelled **X**?

2. The aim of this experiment is to find the critical angle **C** of the given glass block.

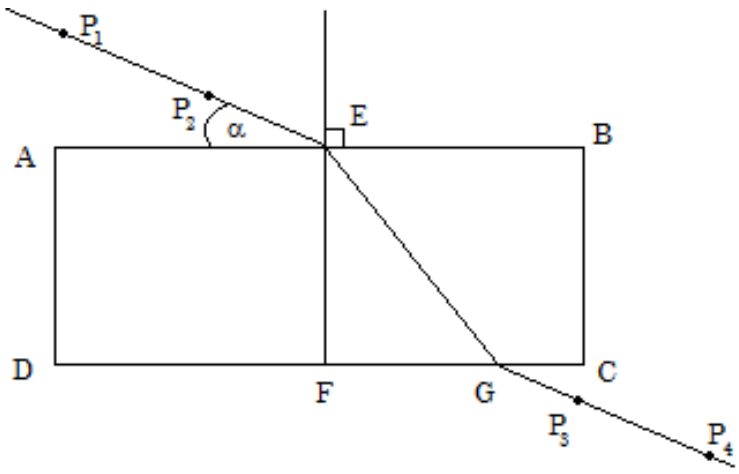


Figure 2

Proceed as follows:

Place a white sheet of paper on the drawing board. Place the glass block, with one of its largest surfaces top most on top of the white paper. Mark the outline of the glass block on the paper with a pencil. Then remove the glass block and draw a line which cuts its largest sides normally at E and F as shown in figure 2 above.

Using a protractor draw an angle $\alpha = 30°$ with the glass block. Replace the glass block in its original position and stick the first pin P_1 and second pin P_2 along the line of angle $\alpha = 30°$. Stick the third and fourth pins P_3 and P_4 respectively on the opposite side of the glass block such that P_3 and P_4 fall on a straight line with P_1 and P_2 when viewed through side CD of the glass block.

Remove the glass block and trace the straight path taken by the ray G P_3 P_4. Using a ruler, join G and E.

Measure the angle of refraction $r°$, then calculate the values of $\cos\alpha$ and $\sin r°$. Repeat the same procedure for values $\alpha = 40°$, $50°$, $60°$, $70°$ and $80°$. Record your results in tabular form for the values of α, $r°$, $\sin r°$ and $\cos\alpha$.

(a) Plot a graph of $\sin r°$ (vertical axis) against $\cos\alpha$ (horizontal axis).
(b) Find the slope of the graph.
(c) Calculate the value of C where slope $= \sin C$.
(d) State the possible sources of error and precautions you have taken during the experiment.

3. The aim of this experiment is to determine the **e.m.f. E** and internal resistance **r** of a cell.

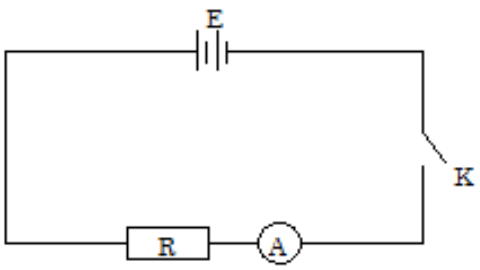


Figure 3

(a) Connect the circuit as shown in figure 3 above. Put $R = 1\Omega$ and quickly read the value of i on the ammeter.
(b) Repeat the procedure in 3 (a) above, for values of $R = 2\Omega$, 3Ω, 4Ω and 5Ω respectively.
(c) Tabulate your results and complete the following table.

Resistance R **(Ω)**	**Current** i **(A)**	$\frac{1}{i}$ **(A^{-1})**
1		
2		
3		
4		
5		

(d) Plot the graph of R against $\frac{1}{i}$.
(e) The graph uses the equation $R = \frac{E}{i} - r$.
  (i) Suggest how E and r may be evaluated from your graph.
  (ii) Evaluate E for one cell.
  (iii) Evaluate r for one cell.
(f) State one source of error and suggest one way of minimizing it.

2004 - PHYSICS 2A ALTERNATIVE A PRACTICAL

1. The aim of this experiment is to determine the mass of a given dry cell, size "AA".

   You are provided with a dry cell, a knife edge, two weights 50 g and 20 g, and a metre rule.

   Proceed as follows:

   (a) Locate and note the centre of gravity C of the metre rule by balancing on the knife edge.

   (b) Suspend the 50 g mass on one side of the metre rule, and 20 g together with the dry cell on the other side of the metre rule adjusting their position until the metre rule balances horizontally, as shown in Figure 1 below.

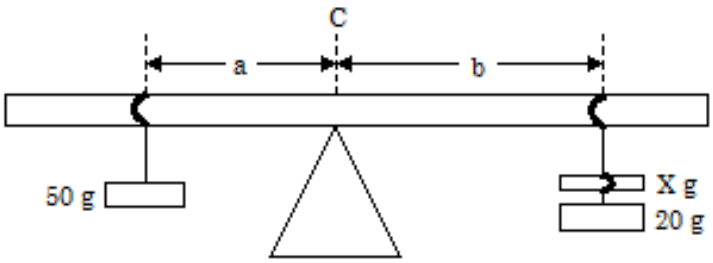


Fig. 1

   (c) By fixing a = 5 cm from C find its corresponding length, b, from C.

   (d) Repeat and tabulate your results using a = 10 cm, 15 cm, 20 cm and 25 cm.

   (e) Draw a graph of "a" against "b" and calculate its slope G.

   (f) Calculate X from the equation $G = \frac{20 + X}{50}$. **(25 marks)**

2. You are provided with a glass block, drawing board, optical pins and plane papers.
   Place a white piece of paper on the drawing board. Place the glass block with one of its largest surface top most on top of the white paper. Mark the outline of the glass block on the paper with a pencil. Remove the glass block and draw a normal as shown in Figure 2 below.

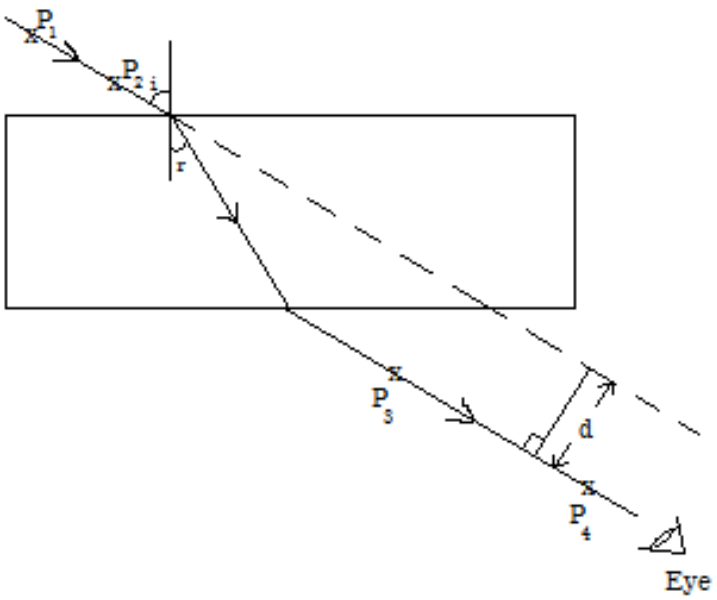


Fig. 2

(a) Draw a line making an angle of incidence, i of 30°. Erect two pins P_1 and P_2 on this line at a suitable distance apart. Replace the glass block and erect two more pins P_3 and P_4 at positions which appear to be in a straight line with the other two pins as seen through the glass block from the other side.

Remove the glass block and draw the complete path of the ray (see Fig. 2). Measure the angle of refraction, r.

(b) (i) Extend the direction of the incident ray as shown by the dotted line.

(ii) Measure the perpendicular distance 'd' between extended incident ray and the emergent ray.

(c) Repeat the procedure in (a) and (b) above for angles of incidence of 30°, 40°, 50°, 60° and 70°. (In each case make your drawings on a fresh part of the drawing paper).

(d) Tabulate your results as shown in Table 1 below.

i (deg)	r (deg)	d (cm)	$d \cos r$	$\sin(i - r)$
30				
40				
50				
60				
70				

(i) Plot a graph of $d \cos r$ against $\sin(i - r)$.

(ii) Find the gradient of the graph.

(iii) Measure the width of the glass block.

(iv) How is the gradient of the graph in 2 (a)(ii) and the width of the glass block in 2 (a)(iii) related?

NB: Hand in your diagrams (drawings) together with your answer booklet.

(25 marks)

3. Determine the resistivity ρ of the wire labelled W and the internal resistance of the battery provided.
Proceed as follows:

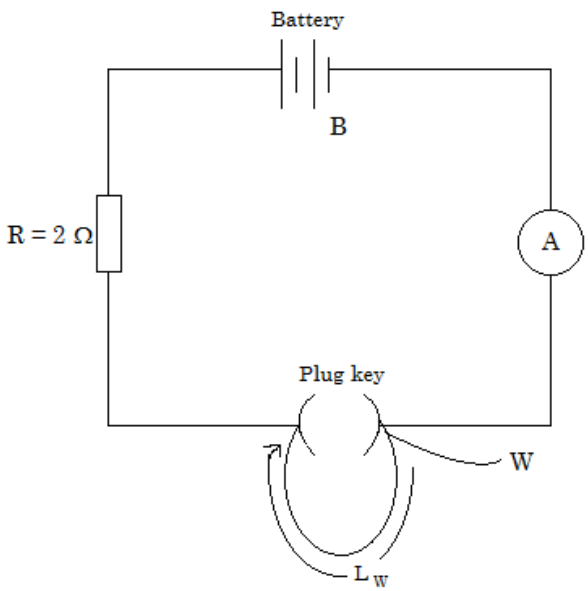


Fig. 3

Connect the circuit as shown in fig. 3 above. With the plug key open adjust the length of wire W to a value of 20 cm. Note the ammeter reading.

NB: The plug key should remain open throughout the experiment.

(a) Repeat the procedure above for L_W = 40 cm, 60 cm, 80 cm and 100 cm each time recording the ammeter reading.

(b) Tabulate your results as shown in Table 2 below.

Length L_W of wire (cm)	Current I (A)	$\frac{1}{I}\left(\mathrm{A}^{-1}\right)$

(c) (i) Plot a graph of $\frac{1}{I}$ (vertical) against L_W (horizontal).

(ii) Determine the slope G.

(iii) Determine the intercept Y on the vertical axis.

(d) Measure and record the diameter at four different places on the wire. Hence find the mean value of diameter d.

(e) Given that $G = \frac{4\rho}{\pi d^2 E}$ and $Y = \frac{R+r}{E}$

Where E is the emf of the battery, and $R = 2\Omega$, Find the

(i) Resistivity ρ of the wire.

(ii) Internal resistance r of the battery. **(25 marks)**

Appendix A

Sources of Laboratory Equipment

In order to gain a thorough understanding of science, students must be able to make a connection between classroom learning and the outside world. The following is a list of locally available materials which may be used to substitute conventional materials and apparatus for various activities. These materials have the following advantages:

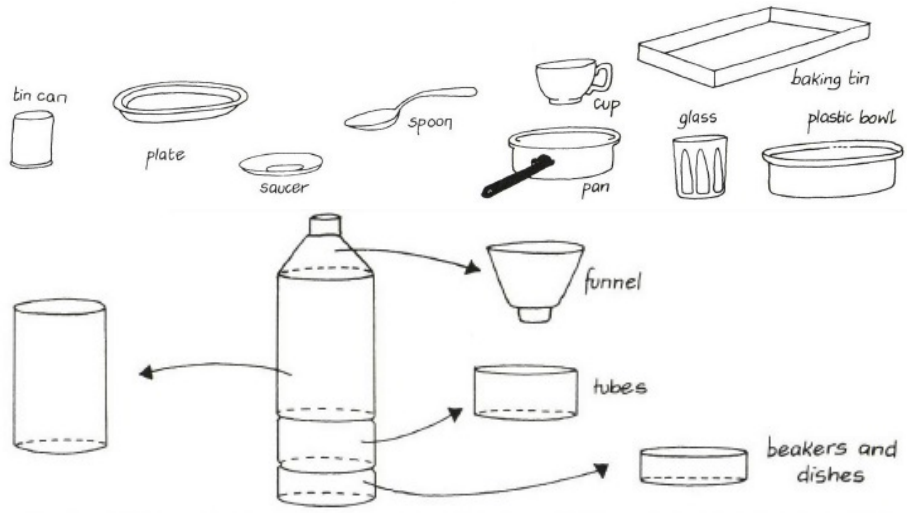
- They are readily available in the village or a nearby town;
- They are cheaper than conventional materials;
- They may safely substitute the conventional materials without fear of losing accuracy or understanding;
- They help students to draw a connection between science education and the world around them.

Imagination and innovativeness is encouraged on the part of the student and teacher to find other suitable local substitutions.

Below are common apparatus you might order from a laboratory supply company, and comments about which have good if not superior alternatives available in villages and towns. Given equal quality, it is generally better to use local materials, because these help connect classroom learning to students' lives. The apparatus listed in this section are the following:

1. Alligator Clips
2. Balance
3. Beakers
4. Blowpipe
5. Bulbs
6. Bunsen Burner
7. Burettes
8. Circuit Components
9. Crucible
10. Containers
11. Deflagrating Spoon
12. Delivery Tube
13. Drawing Board
14. Droppers
15. Electrodes
16. Electrode Holders
17. Eureka Can
18. Filter Paper
19. Flasks
20. Funnel
21. Glass blocks
22. Gloves
23. Goggles
24. Heat Sources
25. Indicators
26. Iron Filings
27. Masses
28. Measuring Cylinder
29. Metre Rule
30. Microscope
31. Mirrors
32. Mortar and Pestle
33. Nichrome Wire
34. Optical Pins
35. Pipettes
36. Pulleys
37. Resistors
38. Retort Stand
39. Scale Pans
40. Scalpels
41. Slides and Cover Slips
42. Spatula
43. Spring Balance
44. Springs
45. Stoppers
46. Stopwatches
47. Test Tubes
48. Test Tube Brush
49. Test Tube Holder / Tongs
50. Test Tube Racks
51. Tripod Stands
52. Volumetric "Glass"ware
53. Wash Bottle
54. Water Bath
55. Weights
56. White Tiles
57. Wire
58. Wire Gauze

How many experiments can be carried out with everyday items?

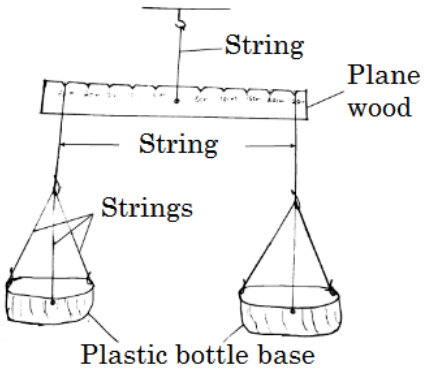


A.1 Alligator Clips

Use: Connecting electrical components
Materials: Clothespins, aluminum foil, glue
Procedure: Glue aluminum foil around the clamping tips of a clothespin.

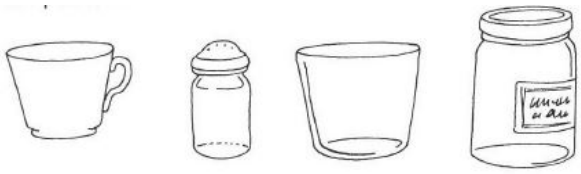
A.2 Balance

Use: Measuring mass
Materials: Ruler or wooden bar 30 cm × 2 cm, nails, razor/knife, string/wire, pen, 2 Scale Pans
Procedure: Find the balancing point of the ruler/wood block and mark it with a pen. Use a heated nail to make a hole through this point. Make notches at 5 cm intervals on either side of the center hole using a razor/knife to suspend scale pans. Use a string/wire tied through the center hole to suspend the balance.



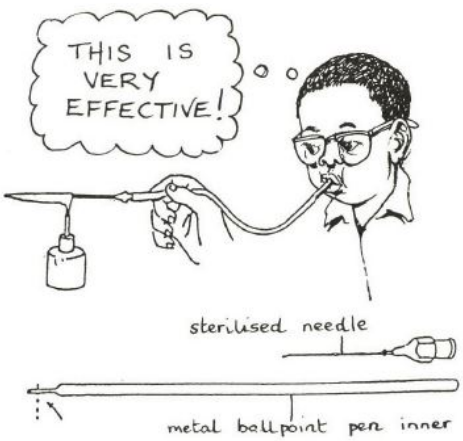
A.3 Beakers

Use: To hold liquids, to heat liquids
Materials: Water bottles, jam jars, metal cans, knife/razor
Procedure: Take empty plastic bottles of different sizes. Cut them in half. The base can be used as a beaker. Jam jars made of glass, cut off metal cans and aluminum pots may be used when heating.
Safety: Glass containers may shatter if heated too much. Use standard laboratory equipment if extreme heating is needed.



A.4 Blowpipe

Use: Increasing temperature of flames
Materials: Syringe needle, tube/straw/pen tube
Procedure: For sterilisation heat the needle in open fire for a longer time before using it. A drinking straw or a clean plastic tube can be used as a connection to the mouth.



A.5 Bulbs

Use: Electrical circuits, diodes
Materials: Broken phone chargers, flashlights, other electronic devices
Procedure: Look for LEDs from broken items at hardware stores, local technicians, or small shops.

A.6 Bunsen Burner

See Heat Sources (p. 213).

A.7 Burettes

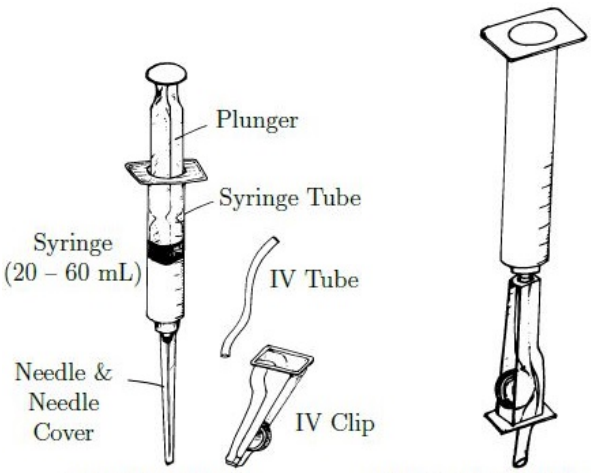
Use: Titration

A.7.1 Version 1

Materials: 10 mL syringes
Procedure: Use 10 mL disposable plastic syringes with 0.2 mL gradations. Students can estimate between the lines to at least 0.05 mL. If you must buy, buy plastic. Note that broken burettes can often be repaired – see Repairing Burettes (p. 19).

A.7.2 Version 2

Materials: Syringe, IV giving set, super glue, knife
Procedure: Cut off the part of the IV tube with the flow control slider. Remove the plunger from the syringe and use superglue to attach the tube to the nozzle of the syringe.



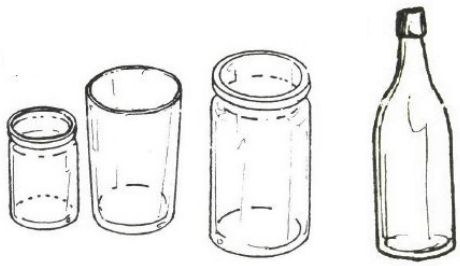
(a) Burette Materials (b) Completed Burette

A.8 Circuit Components

Use: Building simple circuits, Ohm's Law, amplifier, wave rectifiers
Materials: Broken radio, computer, stereo, other electrical devices
Procedure: Remove resistors, capacitors, transistors, diodes, motors, wires, transformers, inductors, rheostats, pulleys, gears, battery holders, switches, speakers and other components from the devices. Capacitors tend to state their capacitance in microFarads on their bodies.

A.9 Containers

Use: Measuring large volumes (100 mL – 2 L) of solution, titration, storage
Materials: Plastic water bottles, jars, tin cans
Procedure: Identify the volume of useful marks on the bottles and combine to measure accurate volumes.

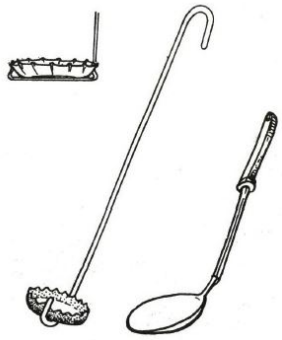


A.10 Crucible

Use: Heating substances at very high temperatures
Materials: 2 metal spoons, wire
Procedure: Place the material in one spoon and then wire 2 spoons together.

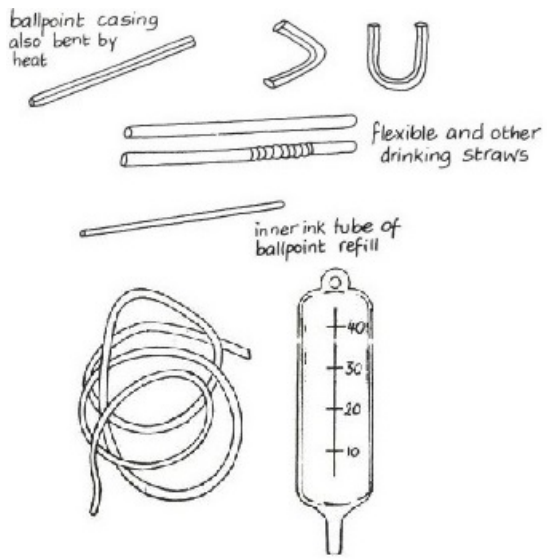
A.11 Deflagrating Spoon

Use: For heating chemicals to observe melting, decomposition, or other changes on heating
Materials: Metal spoons, galvanised wire, soda bottle cap
Procedure: Bend 30 cm of galvanised wire as shown. The wire should hold the bottle cap firmly.



A.12 Delivery Tube

Use: Movement and collection of gases, capillary tubes, hydraulic press
Materials: Straws, pen tubes, IV tubing (giving sets) from a pharmacy, bicycle tubing



A.13 Drawing Board

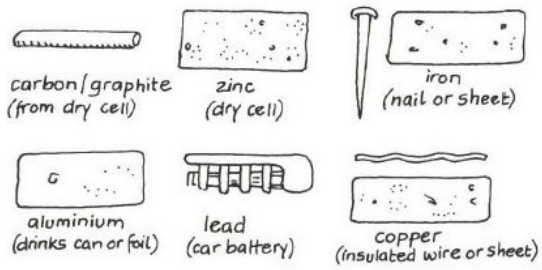
Use: Dissection, reflection, refraction of light
Materials: Thick cardboard

A.14 Droppers

Use: To transfer small amounts of liquid
Materials: 2 mL syringes, straws
Procedure: Take a syringe. Remove the needle to use as a dropper. Or insert a straw into a liquid and then plug the free end with a finger to remove a small amount and use as a dropper.

A.15 Electrodes

Use: Electrolysis

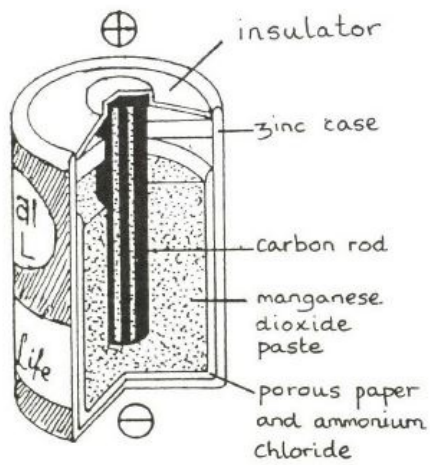


A.15.1 Graphite

Materials: Old dry cell batteries
Procedure: Gently smash an old battery (D size) with a rock and pull out the electrode with pliers. DO NOT do this with alkaline batteries (most AA size) as they contain caustic liquids.

A.15.2 Zinc

Materials: New dry cell batteries
Procedure: Carefully open up a NEW dry cell (D size) battery by peeling back the steel shell and slicing the plastic inside. You should find a cylindrical shell of zinc metal. Empty out the black powder inside (manganese dioxide mixed with zinc chloride and ammonium chloride; wash your hands after) and keep the graphite electrode for another day. The zinc shell should then be cut into strips, scraped clean, and boiled in water or washed with soap to remove any residual chemicals that might affect your experiment.



A.15.3 Iron

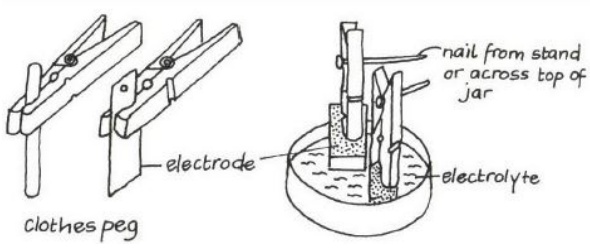
Materials: Ungalvanized nails from a hardware store

A.15.4 Copper

Materials: Thick wire stripped of its insulation, also from a hardware store. Note that copper earthing rods have only a thin surface layer of copper these days.

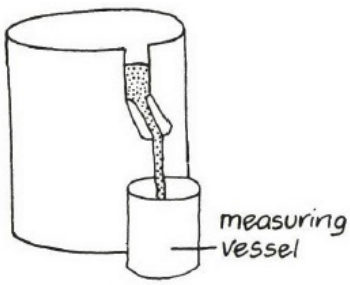
A.16 Electrode Holders

Use: Electrolysis
Materials: Clothes pins



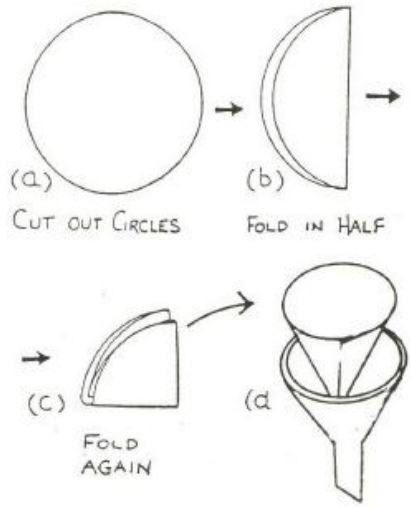
A.17 Eureka Can

Use: To measure volume of an irregular object, Archimedes' Principle, Law of Flotation
Materials: Plastic bottle, knife, Optional: super glue, straw, nail, candle
Procedure: Cut the top off of a 500 mL plastic bottle. Then cut a small strip at the top (1 cm wide by 3 cm long) and fold down to make a spout. Alternatively, heat a nail using a candle and poke a hole near the top of a cut off bottle. Super glue a straw so that it fits securely in the hole without leaking.



A.18 Filter Paper

Use: Filtration, separating mixtures, solutions
Materials: Cement bag paper, toilet paper, cloth

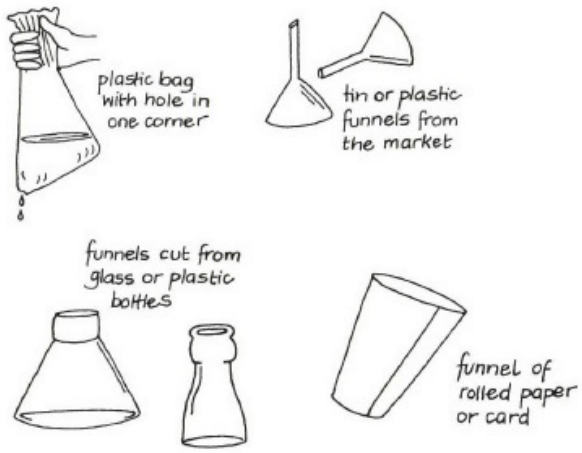


A.19 Flasks

Use: Titrations, mixing solutions
Materials: Clean used liquor bottles, small water bottles
Procedure: When using these flasks for titrations, students must practice swirling enough that the solution remains well mixed.
Safety: When heating glass liquor bottles, make sure the cap is off.

A.20 Funnel

Use: To guide liquid or powder into a small opening
Materials: Empty water bottles, knife
Procedure: Take an empty water bottle and remove the cap. Cut it in half. The upper part of the bottle can be used as a funnel.



A.21 Glass blocks

Use: Refraction of light
Materials: 8 mm - 15 mm slabs of glass
Procedure: Have a craftsman make rectangular pieces of glass with beveled edges, so students do not cut themselves. Glass blocks from a lab supply company are generally 15 mm thick. 8 mm and 10 mm glass is relatively common in towns. 12 mm and thicker glass exists though is even more difficult to find. Stack several pieces of thinner glass together and turn them on their edge.

A.22 Gloves

A.22.1 Latex gloves

Use: First aid, when one has open cuts on hands, handling specimens. They are worthless to the chemist because they make the hands less agile and give the user a false sense of security.
Safety: Concentrated acids and organic chemicals burn straight through latex.

A.22.2 Thick gloves

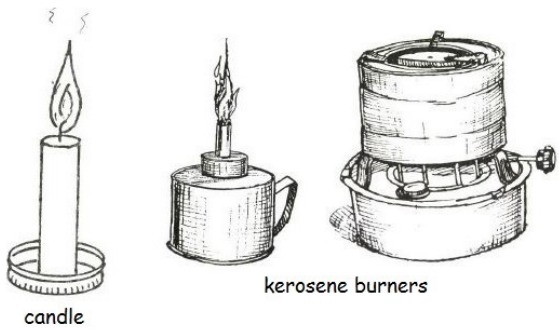
Use: For working with organic solvents. Remember that the most dangerous organic solvents (benzene, carbon tetrachloride) should never be used in a school, with or without gloves.
Materials: Thick rubber gloves from village industry supply companies and some hardware stores
Safety: In general, avoid using chemicals that would make you want to wear gloves.

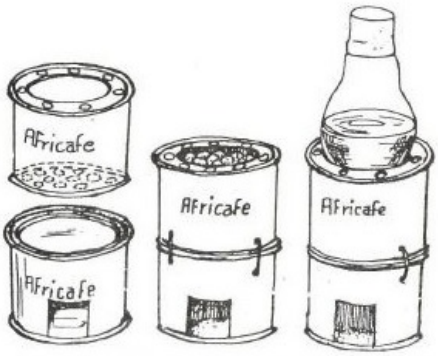
A.23 Goggles

Use: Handling concentrated acids
Materials: 1.5 L plastic water bottles, cardboard, sunglasses
Procedure: Cut a strip of plastic from a water bottle. Attach around your head with string or by using stiff cardboard as a frame. Goggles do not need to be impact resistant – they just need to stand between hazardous chemicals and your eyes.

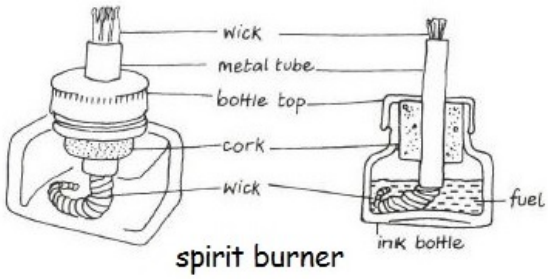
A.24 Heat Sources

Use: Heating substances
Materials: Candles, kerosene stoves, charcoal burners, Motopoa (alcohol infused heavy oil), butane lighters, spirit burners, metal can, bottle caps
Motopoa provides the best compromise heat source - it is the easiest to use and safest heat source with locally available burners.
Procedure: Cut a metal can in half or use a bottle cap and add a small amount of Motopoa.
Safety: Always have available fire-fighting equipment that you know how to use. Remember that to put out a Bunsen burner safely, you need to turn off the gas.





charcoal burner



spirit burner

A.24.1 Heating Solutions

The ideal heat source has a high heat rate (Joules transferred per second), little smoke, and cheap fuel, i.e. Motopoa. A charcoal stove satisfies all of these but takes time to light and requires relatively frequent re-fueling. Kerosene stoves have excellent heat rates but are smoky.

A.24.2 Heating Solids

The ideal heat source has a high temperature and no smoke, i.e. a Bunsen burner. For heating small objects for a short time (no more than 10-20 seconds), a butane lighter provides a very high temperature. Motopoa will provide a flame of satisfactory temperature for as long as necessary.

A.24.3 Flame Tests

The ideal heat source has a high temperature and produces a non-luminous flame, i.e. a Bunsen burner. Motopoa is next best – hot and non-luminous. Spirit burners produce a non-luminous flame at much greater cost, unless methylated spirits are used as fuel in which case the flame is much cooler. A butane lighter produces a very hot flame of sufficient size and time for flame tests although the non-luminous region is small. Kerosene stoves will work for some salts.

A.25 Indicators

Use: Determine presence of acid or base, determine pH
Materials: Rosella leaves, hot water, bottle
Procedure: Place some coloured leaves into a bottle of warm water to extract the colour. Use a straw to drop onto solutions or prepare indicator paper by dipping thing strips into the coloured solution. Rosella turns red for acids and greenish blue for bases.

A.26 Iron Filings

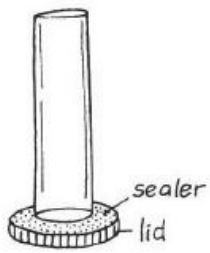
Use: To map magnetic fields
Materials: Steel wool / Iron wool used for cleaning pots
Procedure: Rub some steel wool between your thumb and fingers. The small pieces that fall are iron filings. Collect them in a matchbox or other container to use again.

A.27 Masses

See Weights (p. 219).

A.28 Measuring Cylinder

Use: Measuring volume
Materials: Plastic bottles of different sizes, syringes (10 mL - 50 mL), fluorescent light tubes, marker pen, ruler, bucket of water
Procedure: Using the syringe, transfer a known volume of water from the bucket to the empty bottle. Use the marker pen to mark the level of water on the bottle. Repeat for a range of volumes, using a ruler to complete the scale.



A.29 Metre Rule

Use: Measuring length
Materials: Slabs of wood, ceiling board, permanent pen
Procedure: Buy one, take it and a permanent pen to a carpenter, and leave with twenty. Measure each new one to the original rule to prevent compounding errors.

A.30 Microscope

See Low Tech Microscopy (p. 76).

A.31 Mirrors

A.31.1 Plane Mirrors

Use: Microscope, Laws of Reflection
Materials: piece of thin glass, kibatari, super glue, small wooden blocks
Optional: Small pieces of mirror glass are cheap or free at a glass cutter's shop
Procedure: Light the kibatari so that it creates a lot of smoke. Pass one side of the glass repeatedly over the kibatari until that side is totally black. The other side acts as a mirror. Super glue to small wooden blocks to stand upright.

A.31.2 Curved Mirrors

Use: Curved mirror practicals
Materials: Spoons
Procedure: Inside surface is a concave mirror; back surface is a convex mirror.

A.32 Mortar and Pestle

Use: To powder chemicals
Materials: 2 metal spoons, glass bottle
Procedure: Place chemicals between two nested metal spoons and grind down. Alternatively, crush chemicals on a sheet of paper by pressing on them with the bottom of a glass bottle.

A.33 Nichrome Wire

For flame tests in chemistry, you can use a steel wire thoroughly scraped clean with iron or steel wool. For physics experiments, see Wire (p. 220).

A.34 Optical Pins

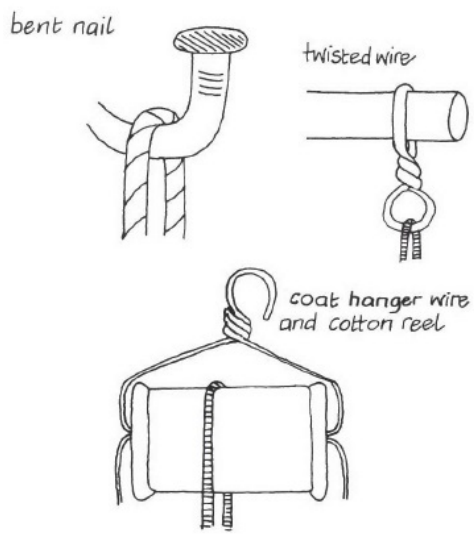
Use: Compass needles, making holes, dissection, mirror practicals
Materials: Office pins, sewing needles, needles from syringes

A.35 Pipettes

Use: Transferring small amounts of liquid
Materials: Disposable plastic syringes (1, 2, 5, 10, 20, 25, 30 and 50 mL sizes)
Procedure: Suck first 1 mL of air and then put the syringe into the solution to suck up the liquid. There should be a flat meniscus under the layer of air.
Safety: Avoid standard pipettes to eliminate danger of mouth pipetting.

A.36 Pulleys

Use: Simple machines
Materials: Bent nail, twisted wire, thread reel, water bottle, string, coat hanger
Procedure: Cut off the top of a water bottle just below the lip where the top screws on. Run string or stiff wire through the centre to hang from a table or chair.



A.37 Resistors

Use: Electrical components
Materials: Old radios, circuit boards, soldering iron
Procedure: Remove resistors from old radios and circuit boards by melting the solder with a soldering iron or a stiff wire heated by a charcoal stove. If you need to know the ohms, the resistors tell you. Each has four strips (five if there is a quality band) and should be read with the silver or gold strip for tolerance on the right. Each color corresponds to a number:

black = 0	yellow = 4	violet = 7
brown = 1	green = 5	gray = 8
red = 2	blue = 6	white = 9
orange = 3		

and additionally for the third stripe: gold = -1 and silver = -2.
The first two numbers should be taken as a two digit number, so green-violet would be 57, red-black 20, etc. The third number should be taken as the power of ten (a 10^n term), so red-orange-yellow would be $23 \times 10^4 = 230000$, red-brown-black would be $21 \times 10^0 = 21$ and blue-gray-silver would be $68 \times 10^{-2} = 0.68$. The unit is always ohms. The fourth and possibly fifth bands may be ignored.

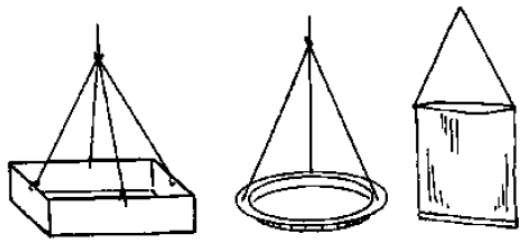
A.38 Retort Stand

Use: To hold springs, burettes, pendulums or other objects
Materials: Filled 1.5 L water bottle, straight bamboo stick, tape, marker
Procedure: Tape the bamboo stick across the top of the water bottle so that it reaches out 20 cm to one side. Attach a small clamp if required or hang the object directly from the bamboo stick.

Alternatively, place a 1 cm piece of reinforcing rod in a paint can full of wet cement and let it dry. Then attach a boss head and clamp.

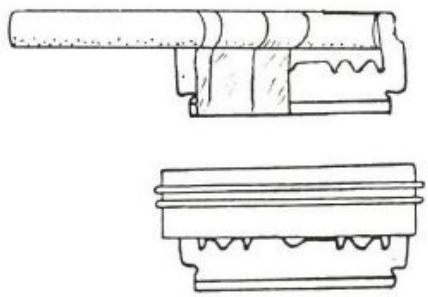
A.39 Scale Pans

Use: Beam balance
Materials: Plastic bottle, cardboard box, string
Procedure: Cut off the bottom of a plastic bottle or cardboard box. Poke 3 or more holes near the top and tie string through each hole. Join strings and tie at the top to hang from a single point.



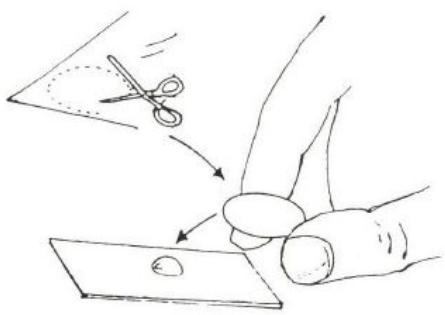
A.40 Scalpels

Use: Dissection
Materials: Razor blades, tongue depressors, super glue
Procedure: Add a handle by gluing a tongue depressor on either side of the razor blade. Hold together with a rubber band until dry.
Safety: Dull blades should be discarded. Because students need to apply more pressure when using them, there is a greater risk of slipping and thus of cuts. Sharp tools are much safer.



A.41 Slides and Cover Slips

Use: Microscopy
Materials: Small pieces of glass, stiff plastic
Procedure: Small piece of glass provides a slide for mounting the specimen. Cover slips can be made from thin (but stiff) transparent plastic from display packing or bottles. Cut into small squares or circles.

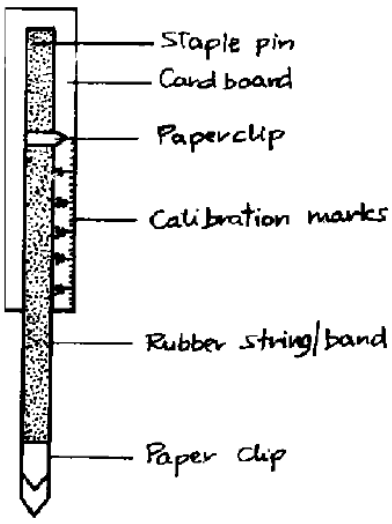


A.42 Spatula

Use: Transferring salts
Materials: Stainless steel spoons
Procedure: Use the handle end to remove salts from containers.
Safety: Clean all metal tools promptly after using with hydroxide, potassium manganate (VII), or manganese (IV) oxide. If the spoon corrodes, scrape with another spoon or steel wool.

A.43 Spring Balance

Use: To measure force applied on an object
Materials: Strip of cardboard, rubber band, 2 paper clips, staple pin, pen
Procedure: Cut a rubber band and fix one end to the top of a cardboard strip using a staple pin. (A stronger rubber band allows for a greater range of forces to measure.) Attach one paper clip near the top as a pointer. Attach the other paper clip as a hook at the bottom of the rubber band. Calibrate the spring balance using known masses. Write the equivalent force in Newtons on the cardboard. (A 1 g mass has a weight of 0.01 N, 100 g has a weight of 1 N, etc.)



A.44 Springs

Use: Hooke's Law, potential energy, work, spring balance
Materials: Springs from hardware stores, bike stores, junk merchants in markets, window blinds; stiff wire; rubber bands; strips of elastic
Procedure: Remove plastic covering if necessary and cut to a desired length (5 cm). Alternatively wind a stiff wire around a marker pen or use rubber bands or elastic from a local tailor.

A.45 Stoppers

Use: To cover the mouth of a bottle, hold a capillary tube
Materials: Rubber from old tires or sandals, cork, plastic bottle cap, pen tube, super glue
Procedure: Cut a circular piece of rubber. If the stopper is being used to hold a capillary tube, a hole can be melted in a plastic cap or rubber stopper. Alternatively, super glue a pen tube to a plastic bottle cap and connect to rubber tubing.

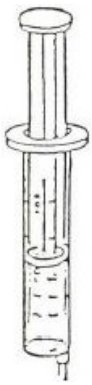
A.46 Stopwatches

Use: Simple pendulum, velocity, acceleration
Materials: Athletic and laboratory stopwatches from markets, digital wristwatches

A.47 Test Tubes

A.47.1 Plastic Test Tubes

Use: To heat materials without a direct flame, to combine solutions
Materials: 10 mL syringes, matches
Procedure: Remove the needle and plunger from 10 mL syringes. Heat the end of the shell with a match until it melts. Press the molten end against a flat surface (like the end of the plunger) to fuse it closed. If the tube leaks, fuse it again. Test tubes made this way may be heated in a water bath up to boiling, hot enough for most experiments.

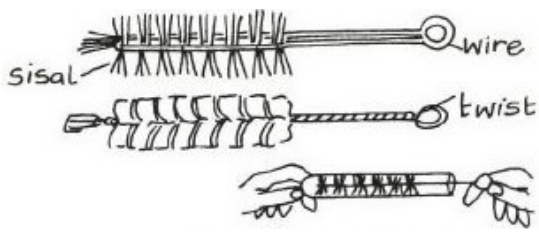


A.47.2 For Thermal Decomposition

See Deflagrating Spoon (p. 211).

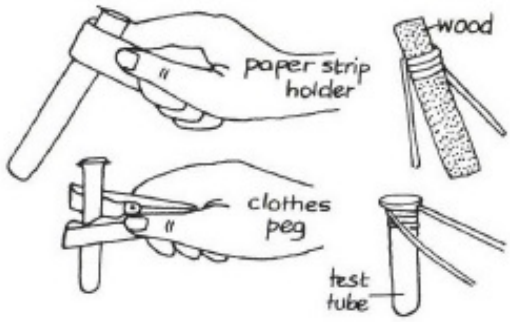
A.48 Test Tube Brush

Use: Cleaning test tubes
Materials: Sisal, wire
Procedure: Twist the wire around the sisal as shown or put a little sand in the test tube as an abrasive.



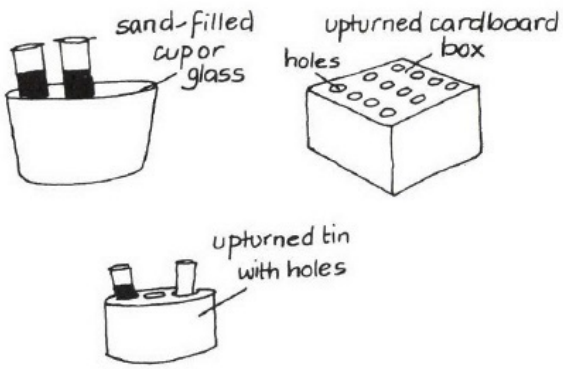
A.49 Test Tube Holder / Tongs

Use: To handle test tubes
Materials: Wooden clothespins, stiff wire, strip of paper or cloth
Procedure: Use clothespins or stiff wire for prolonged heating, or strips of paper or cloth for short-term heating.



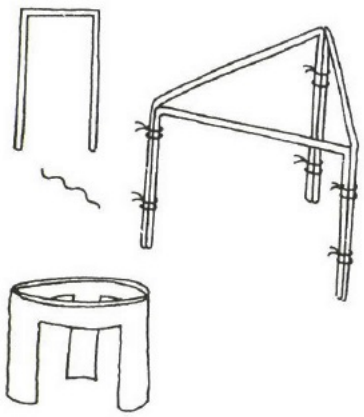
A.50 Test Tube Racks

Use: To hold test tubes vertically in place
Materials: Wire grid from local gardening store, styrofoam block, plastic bottle, sand, knife
Procedure: Fold a sheet of wire grid to make a table; punch holes in a piece of styrofoam; cut a plastic bottle in half and fill it with sand to increase stability. Or cut a plastic bottle along its vertical axis and rest the two cut edges on a flat surface. Cut holes into it for the test tubes.



A.51 Tripod Stands

Use: For supporting containers above heat sources, for elevating items
Materials: Stiff wire, metal rods, tin can
Procedure: Join bent pieces of thick wire together. Or cut the sides of a tin can to leave 3 legs.



A.52 Volumetric "Glass"ware

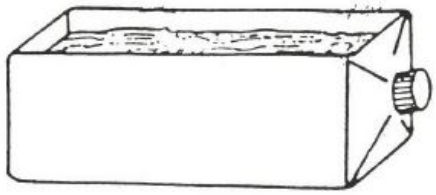
See Containers (p. 210).

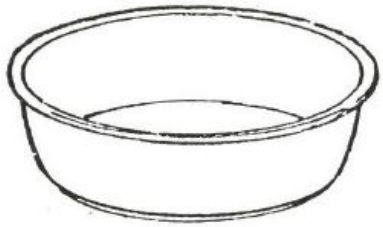
A.53 Wash Bottle

Use: Washing hands after experiments
Materials: Water bottle, detergent, needle
Procedure: Put a hole in the cap of a water bottle using a syringe needle.

A.54 Water Bath

Use: To heat substances without using a direct flame
Materials: Heat Sources, water, cooking pot
Procedure: Bring water to a boil in a small aluminum pot, then place the test tubes in the water to heat the substance inside the test tube. Prevent test tubes from falling over by clamping with clothespins or placing parallel wires across the container.





A.55 Weights

A.55.1 Crude Weights

Use: Concept of units, mass, weight
Materials: Batteries, coins, glass marbles from town, etc.
Procedure: Use objects of unknown mass to create new units and impart the concept of unit measure.

A.55.2 Adding Weight in Known Intervals

Use: Hooke's Law practical
Materials: Water bottles, syringe
Procedure: Consider "zero added mass" the displacement of the pan with an empty water bottle. Then add masses of water in g equal to their volumes in mL (e.g. 50 mL = 50 g).

A.55.3 Precise Weights

Materials: Plastic bags, sand, stones, 250 mL water bottles (all identical), tape, pen
Procedure: Use a beam balance and known masses at a market or nearby school to measure exact masses of bags of sand or stones. Use a marker pen to mark the masses on the bags.

If using water, use a beam balance from a nearby school to measure the exact mass of an empty water bottle. Add a volume of water in mL equal to the mass in g needed to reach a desired total mass. (The density of water is 1.0 g/mL.) This can be done precisely by using a plastic syringe. Label the bottle with tape and a pen.

A.56 White Tiles

Use: Titration
Materials: White paper
Procedure: If students are using syringes as burettes, they can also hold their flask up against a white wall.

A.57 Wire

A.57.1 Connecting Wires

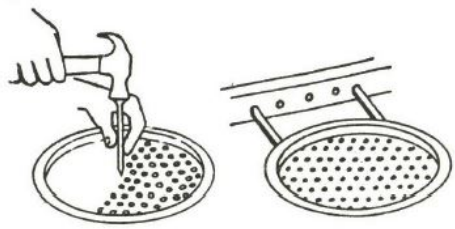
Use: Connecting circuit components, current electricity
Materials: Speaker wire, knife
Procedure: Speaker wire can be found at any hardware store or taken from old appliances - the pairs of colored wires brained together. Strip using a knife, scissors or a wire stripper.

A.57.2 Specific Gauge Wire

Use: Electrical components, motors, transformers, simple generators
Materials: Copper wire without plastic covering (transformer wire), knife/scissors, matches
Procedure: Scrape or burn off the insulating varnish at any points you wish to make electrical contact. These wires come in a variety of diameters (gauges). A useful chart for converting diameter to gauge may be found at http://www.dave-cushman.net/elect/wiregauge.html. If the wire is sold by weight, you can find the length if you know the diameter - the density of copper metal at room temperature is 8.94 g/cm^3. For example, with 0.375 mm wire, 250 g is about 63 metres.

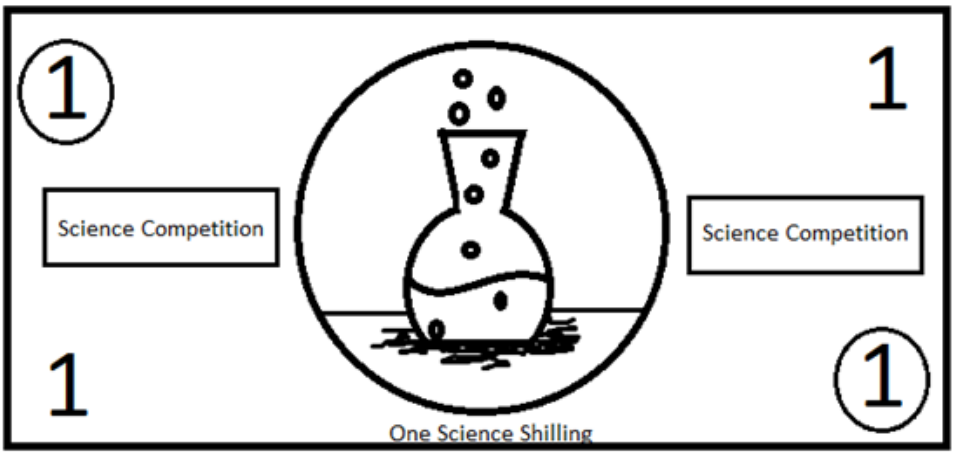
A.58 Wire Gauze

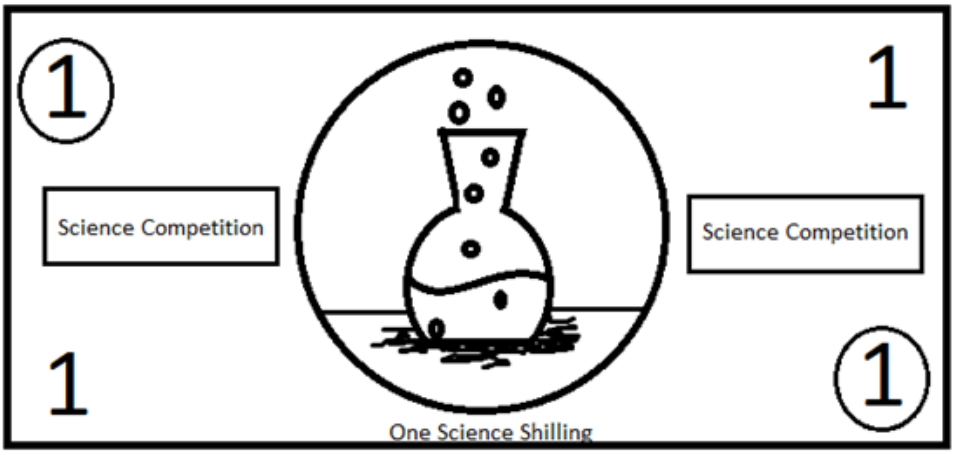
Use: Placing objects over heat
Materials: Tin can lid
Procedure: Poke holes in a tin can lid.

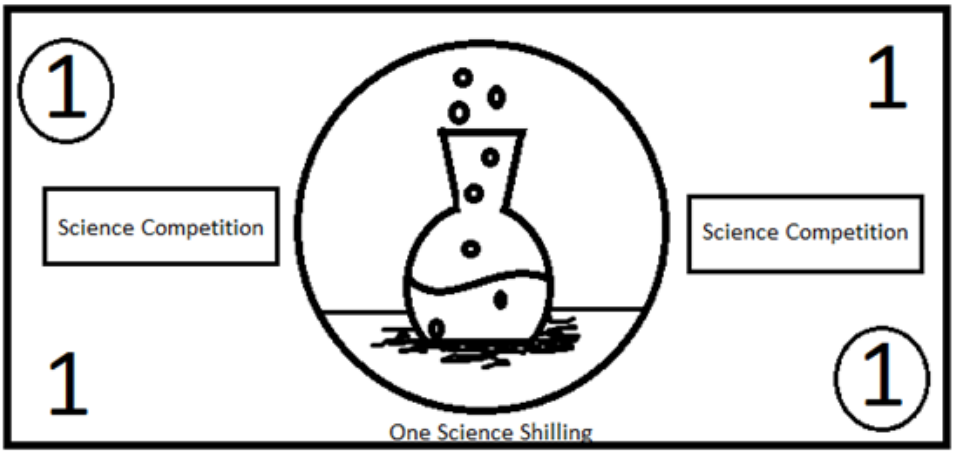


Appendix B

Science Shillings







Appendix C

Kiswahili Laboratory Glossary

While managing a lab filled with dangerous chemicals, breakable glassware, open flames, and inexperienced students, it is important to know how to communicate with students easily to keep the place safe and running smoothly. Below is a list of Kiswahili words and phrases that may be helpful in a lab setting.

English	Kiswahili
to absorb	-fyonza
to be absorbed	-fyonzwa
accident	ajali
to affect, to influence	-athiri
ashes	majivu
battery acid	maji makali
to boil Ex. I boil the water.	-chemsha Ninachemsha maji.
To be boiling Ex. The water is boiling.	-chemka Maji yanachemka.
to break Ex. I broke the glass.	-vunja Nimevunja kioo.
to be broken Ex. The glass broke.	-vunjika Kioo kimevunjika.
broken, bad, rotten	bovu
to burn Ex. I am burning paper.	-unguza Ninaunguza karatasi.
(to be) burnt Ex. The paper is burnt.	-ungua Karatasi imeungua.
calcium hydroxide solution (lime water)	maji chokaa
carefully	taratibu
to cause	-sababisha
to be caused by	-sababishwa
caution Ex. Heat with caution.	utaratibu Ex. Pasha na utaratibu.
to change Ex. I am changing the color of this.	-badilisha Ninabadilisha rangi ya hii.
to be changed Ex. The color has been changed.	-badilika Rangi imebadilika.
changes	mabadiliko
chemical	kemikali
container, glassware	chombo
to clean	-safisha
to collide	-gongana
color	rangi
to cool	-poa
danger	hatari
to decrease (transitive) Ex. Decrease the heat.	-punguza Punguza moto.
to decrease in size, to grow smaller (intransitive) Ex: The heat has decreased.	-pungua Moto umepungua.
to destroy, to damage, to contaminate Ex. I contaminated/damaged the chemical.	-haribu Nimeharibu kemikali.
to be destroyed, damaged, contaminated, or expired. Ex. The chemical has gone bad or expired.	-haribika Kemikali imeharibika.

distilled water (from a hardware or car repair shop)	maji baridi
to distribute	-gawa
to draw	-chora
to drink Ex. Do not drink!	-nywa Usinywe!
drop (as in "a drop of water")	tone
to dry	- kausha
effect	athari
to estimate	- kadiri
to evaporate (causative) Ex. I am evaporating water.	-vukiza Ninavukiza maji.
to explode	- lipuka
to fill Ex. I filled the container.	-jaza Nimejaza chombo.
to be full Ex. The container is full.	-jaa Chombo kimejaa.
to filter	-chuja
fire	moto
gas, air	hewa
glass	kioo
group	kikundi (vi-)
to grow	kukua
harm (harmful)	madhara (yenye madhara)
to haul off and slap someone	-ezeka makofi
to heat	-pasha
height, length	urefu
to hit, to knock	-gonga
to increase in size or number, to swell (intransitive) Ex. The amount of food present increased.	-ongezeka Chakula kimeongezeka.
to increase, to add (transitive) Ex. I added more food.	-ongeza Nimeongeza chakula.
instrument/apparatus/	kifaa (vi-)
tool	
to kick someone out Ex. If you break a rule, you will be kicked out of the laboratory.	-fukuza Kama unakiuka sheria, utafukuzwa maabara.
lab activity/experiment	practical
living	hai
match / lighter	kiberiti (vi-) / kiberiti cha gasi
to measure, to test	-pima
to be melted, to be dissolved (Not exactly "to be melted." It's just the intransitive form of to melt or dissolve.) Ex.The salt melted/dissolved.	-yeyuka Chumvi imeyeyuka.
mass	uzito
methylated spirits	spiriti
microscope	hadubini
to mix	-changanya
to be mixed	-changanyika
mixture	mchanganyiko
particle, atom, small piece	chembe ndogo ndogo
permission	ruhusa
poison	sumu
to pour	-mimina
to pour out	-mwaga

to press on, to push on (used to described squeezing of eye droppers)	-bonyeza
to put, to keep	-weka
rubber, rubber tubing	mpira
to rush, to hurry Ex. Do not rush!	-harakisha usiharakishe
scientist	mwanasayansi
serious, attentive	makini
to shake	-tikisa
to share out, to divide	-gawana
slowly	pole pole
to solidify, to freeze	-ganda
steam	mvuke
to stir	-koroga
to stop	-acha
stove	jiko (ma-)
strong, harsh, fierce, dangerous, concentrated (for an acid)	kali
to suck, to pull	-vuta
to throw away, to chuck	-tupa
to touch Ex. Do not touch!	-gusa Usiguse!
to turn off	-zima
to turn on, to light a flame	-washa
volume	ujazo
to wash dishes or glassware	-osha
to wash your hands	-nawa
to watch	-angalia
yeast	hamira

Index

Alligator clips, 210
Ammeters
    checking, 19

Beakers, 210
Beam balance
    construction of, 210
    use of, 57
Benedict's solution
    for food tests, 81
Blowpipe, 210
Box of fun, 143
Bridge Challenge, 151
Bulbs, 211
Bunsen burner, *see* Heat sources
Burettes
    repairing, 20
    sources of, 211

Chemical kinetics, 120
Chemicals
    dangerous, 38
    unknown, 22
Circuit components, 211
Citric acid
    for food tests, 82
    preparation of, 66
Classification, 86
Containers, 211
Copper, *see* Electrodes
Copper sulphate
    for food tests, 81
Cover slips, 77, 217
Critical angle, 136
Crucible, 211
Culture media, 73

Deflagrating spoon, 212
Delivery tube, 212
Dissection, 71
Drawing board, 212
Drop Zone, 150
Droppers, 212

Egg drop, 147
Electrodes, 212
    holders, 213
Equilibrium, *see* Chemical kinetics
Equipment
    salvaging, 19
    sources of, 209
Ethanol
    hazards and use of, 43
Eureka can, 213

Filter paper, 213
First aid, 35
Flasks, 213
Food tests, 81

Funnel, 213

Gallery walk, 143
Glass blocks, 214
Gloves, 214
Goggles, 214
Graphite, *see* Electrodes

Heat sources, 214
Hooke's law, 128
Hydrochloric acid
    for food tests, 82
    identification of, 24
Hydrogen peroxide
    identification of, 26

Indicators
    preparation of, 100
    properties of, 99
    sources of, 215
Industrial ecology, 54
Iodine solution
    for food tests, 82
    identification of, 26
Iron, *see* Electrodes
Iron filings, 215

Jenga Jengo, 148

Laboratory
    classroom management, 47
    cleanup and upkeep, 49
    improving, 17
    rules, 31
    safety guidelines, 31
    starting, 13
    technical needs, 14
Laboratory techniques, 56
    Biology, 70
    Chemistry, 57

Masses, *see* Weights
Measuring cylinder, 215
Metre bridge, 138
Metre rule, 216
Microscope
    low tech, 77
    use of, 75
Mirrors, 216
    curved, 216
    plane, 216
        practical of, 133
Molality, 62
Molarity
    calculating for bottled liquids, 63
    measure of, 61
Mortar and pestle, 216

NECTA Practicals, *see* Practicals
Normality, 61

Ohm's law, 140
Optical pins, 216
Organic waste
  disposal of, 51
  recycling, 53

Past Papers, 162
  Biology, 163
    2007, 169
    2008, 168
    2009, 167
    2010, 166
    2011, 165
    2012, 164
    2013, 163
  Chemistry, 171
    2007, 183
    2008, 181
    2009, 179
    2010, 177
    2011, 175
    2012, 173
    2013, 171
  Physics, 186
    2004, 206
    2005, 204
    2006, 202
    2007, 199
    2008, 196
    2009, 194
    2010, 191
    2011, 190
    2012, 188
    2013, 186
Phenolphthalein
  preparation of, 100
Photosynthesis, 92
Pipettes, 216
  mouth pipetting, 44
Potentiometers, 137
Practicals, 80
  Biology, 80
    classification, 86
    food tests, 81
    photosynthesis, 92
    respiration, 89
    transport, 90
  Chemistry, 95
    chemical kinetics and equilibrium, 120
    qualitative analysis, 107
    volumetric analysis, 96
  Physics, 124
    Hooke's law, 128
    mathematics of, 126
    metre bridge, 138
    Ohm's law, 140
    plane mirror, 133
    potentiometers, 137
    principle of moments, 131
    refraction, 134
    simple pendulum, 130
Principle of moments, 131
Pulleys, 216

Qualitative analysis, 107

Raft Rally, 149
Recycling, 53
Refraction, 134
Relative standardization, 67
Resistors, 217
Respiration, 89
Retort stand, 217

Scale pans, 217
Scalpels, 217
Science
  competitions, 144, 146
  conferences, 145
  events, 143
  fair, 144
Scientific procedure, 153
Silver nitrate
  recycling, 53
Simple pendulum, 130
Slides, 217
Sodium hydrogen carbonate
  preparation of, 66
Sodium hydroxide
  for food tests, 82
  hazards and use of, 42
  identification of, 26
  preparation of, 66
Solutions
  preparation of, 64
    without a balance, 66
Spatula, 218
Specimens
  preparation of, 71
  preservation of, 70
Spring balance, 218
Springs, 218
Stoppers, 218
Stopwatches, 218
Sudan III solution
  for food tests, 82
Sulphuric acid
  identification of, 24
  preparation of, 66
Syringes
  use of, 59

Teacher trainings, 145
Test tubes, 219
  brushes, 219
  holders, 219
  racks, 219

Titration, *see* Volumetric analysis
Transport, 90
Tripod stands, 220

Voltmeters
    checking, 19
Volumetric analysis, 96
Volumetric glassware, *see* Containers

Wash bottle, 220
Waste disposal, 51
Water bath, 220
Weights, 220
White tiles, 221
Wire
    connecting, 221
    nichrome, 216
Wire gauze, 221

Zinc, *see* Electrodes