Chemistry Toolkit

Unit 0

Chemistry Toolkit: Introductory Material

Thinking like a scientist:

Niel deGrasse Tyson (1958-):

1. **Question authority.** No idea is true just because someone says so. Not even me.
2. **Question yourself.** Think for yourself. Believing something doesn't make it so.
3. **Test ideas by the evidence gained from observation and experiment.** If a favorite idea fails a well-designed test, it's wrong. Get over it.
4. **Follow the evidence wherever it leads.** If you have no evidence, reserve judgement.
5. **Remember: you could be wrong.** Even the best scientists have been wrong about some things. Newton, Einstein, and every other great scientist in history – they all made mistakes. Of course they did. They were human.

The scientific method:

Sir Francis Bacon (1561-1626): made use of empirical data

1. Baconian method begins with careful, systematic observations to produce quality facts
2. Use induction to generalize axioms – avoid over generalization beyond the facts
3. Gather additional data, especially negative and exceptional instances
4. Repeat the process stepwise to increase the complexity of the knowledge base

Robert Boyle (1627-1691): described careful, systematic experimentation building on the Baconian method into what has evolved today into the scientific method (model for scientific inquiry)

Scientific method (model for scientific inquiry)

- Purpose: a statement of the problem or area of interest
- Hypothesis: tentative explanation to be tested
- Experimentation: controlled observations that test a hypothesis
  - Independent variable: chosen to be changed systematically in the experiment
  - Dependent variable: measured response to changes in the independent variable
  - Control: baseline or standard for comparison of changes in the dependent variable
  - Observation
    - Qualitative: sensory or descriptive data or information
    - Quantitative: numerical data or information (Antoine Lavoisier)
- Analysis: determination of the results of an experiment (including weaknesses in data collection or instrumentation) and any decisions regarding the next actions to take
- Results: the product of the analysis including any mathematical determinations – results should not include deductions
- Conclusion: induction and deduction based on experimental evidence
  - empirical evidence supports but does not prove
- Reporting (or peer review): allows other scientists to evaluate the findings for themselves and it also allows other researchers to confirm the results

Comparison of hypotheses, theories, and laws:

- Hypothesis: a tentative explanation (based on some observation or measurements)
- Theory: an explanation of a natural phenomenon (from many observations over time)
- Scientific Law: a relationship in nature that has been supported by many experiments (and agreed by peer review to be a useful model)

Laboratory model based on the scientific method:

1. **Title**
2. **Discussion**: information based on a search of previously known information
3. **Research Question**: what is to be investigated
4. **Materials**: supplies, equipment, and chemicals list
5. **Method**: numbered steps that describe the experiment
6. **Data Collection and Processing**:
   measurements, lists, and/or tables
   calculations or analysis
   graphs
7. **Conclusions and Evaluation**: interpret results, suggest improvements
8. **Applications**: What did you learn? How can ideas from the investigation be applied in the real world?

Hallmarks of a well-designed laboratory investigation (or *how to improve an investigation*):

1. Test *only one* variable at a time
2. Test *two samples* at the same time
   One of the samples is the *control* – no changes are made (baseline for comparison)
   The second sample is subjected to the testing of the independent variable
3. Repeat the experiment
4. Keep careful and complete records
5. Report results so other scientists can test and verify the results (scientific integrity)

Numbering and measuring like a scientist:

Kitty Ferguson (1941-): "Science is the art of measurement."

Numbers: what they tell a scientist

Example:
When counting: \$5 = \$5.00
In science: 5 g ≠ 5.00 g
5 g indicates a balance that measures to whole grams
5.00 g indicates a balance that measures to 1/100th of a gram

Measured numbers have three parts:

1. Magnitude: the size of the number

Scientific notation: used in science to express very large or very small numbers
Positive exponents are large numbers:
3.00×10^{8} m/s = 300 000 000 m/s
Negative exponents are small:
6.67×10^{-11} N·m²/kg² = 0.000 000 000 0667 N·m²/kg²
Scientific notation is *always* base 10
Mathematic operators and scientific notation
Multiplication: add exponents
Example: (1.50×10^{14} /s)(2.00×10^{-6} m) = 3.00×10^{8} m/s
Division: subtract exponents
Example: (3.00×10^{8} m/s)/(1.50×10^{14} /s) = 2.00×10^{-6} m

2. Unit: define the quantity or *what* is being measured (volume, mass, length, time, etc.)

Units will have an associated variable, symbol, and name to go with the quantity
Example: the quantity mass uses the variable m, the symbol g, and the name grams

Table D shows the symbol, name, and measured quantity for several metric units
Examples

Quantity	Variable	Symbol	Name
temperature	T	°C *or* K	degrees Celsius *or* kelvin
length	l	m	meter
mass	m	g	gram
pressure	P	kPa	kilopascal

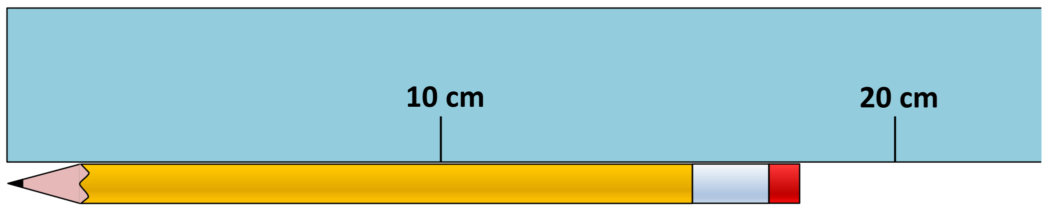
3. Precision: defines the reproducibility or random error in a measurement

One way scientists tell other scientists the precision of their instrumentation is by the number of *significant figures* they report
Precision is determined by how many marks or gradations an instrument shows and where the first digit that contains an error or an estimate

Significant figures: digits in a number that carry information that does not have so much error that the measurement has no meaning

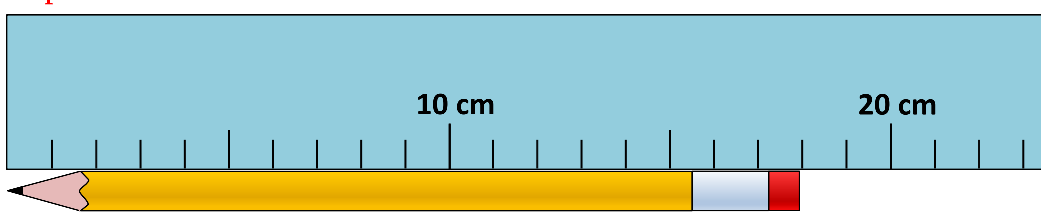
- The number of digits that carry valid information
- Applies to *measurements* and *calculations*, not to counting numbers or defined values
- Rounding rules eliminate non-significant digits

Example 1:



Is the pencil less than 15 cm, 15 cm, or more than 15 cm? Closer to 16, 17, 18, or 19 cm? If the ones digit (either 17 or 18) is already guessing, it makes no sense to say 17.6 cm.

Example 2:



More precision on the ruler allows us to see that the pencil is *just* under 18.0 cm. We might decide 17.8 or 17.9 cm, but no one would say 17.7 or 18.0 cm. Now, a third significant digit (±1.4 in the last digit) is valid but not a fourth ((±14 in the last digit).

Finding the number of significant figures:

1. All nonzero digits are significant
2. Zeros between nonzero digits are significant
3. Zeros to the right of a nonzero digit *and* to the left of a decimal are significant
4. Zeros after *both* a decimal *and* a nonzero digit are significant

Examples:	#Sig. Figs	Explain
1234	4	Rule #1
1230	3	Rule #1
1030	3	1st zero counts by Rule #2, last zero no by Rule #1
1030.	4	1st zero counts by Rule #2, last zero counts by Rule #3
103.0	4	1st zero counts by Rule #2, last zero counts by Rule #4
0.0080	2	1st three zeros no by Rule #1, last zero counts by Rule #4
2.30×10^3	3	*All* zeros in scientific notation always count by Rule #4

Sample Regents questions:

A student measures 10.25 mL of a 10.4 M NaOH solution. The concentration of the base is expressed to how many significant figures?

Answer: 3 SF

A sample of $Mg_{(s)}$ has a volume of 0.063 cm^3. To how many significant figures is the volume of the $Mg_{(s)}$ expressed?

Answer: 2 SF

Significant figures in mathematical operations (+, −, ×, and ÷)

- Direct measure (e.g. filling a graduated cylinder) determines the number of sig. figs.
- Calculations (e.g. V = l × w × h) almost always lead to unrealistic precision
- Rounding rules eliminate non-significant digits

The problem with indirect calculations:

Calculate the volume of a box with L=6.5 mm, W=4.7 mm, and H = 3.9 mm

The calculator gives 119.145 mm^3

This number implies that someone measured 6 sig figs

But, the measurements of the box only had 2 sig figs

Our final answer must reflect the original measurements, so we *must* report the volume as 120 mm^3

Significant Figure Rules for Addition and Subtraction

The precision of the answer must end at the same column as the term with the fewest decimal places that entered the calculation

650 g	10s place precision (second digit before the decimal)
65.60 g	100^{th} place precision (second digit after the decimal)
+ 95.7 g	10^{th} place precision (first digit after the decimal)
811.30 g	The worst precision in the numbers added is in the 10s column
810 g	The final answer rounded properly to the 10s column (2 sig figs)

Significant Figure Rules for Multiplication and Division

The number of significant digits in the answer should be the same as that of the quantity with the fewest significant digits of the quantities being multiplied or divided

650 cm x 32.87 cm	650 cm has 2 sig figs and 32.87 cm has 4 sig figs
= 21 365.5 cm^2	The answer from a calculator
= 21 000 cm^2	The final answer rounded properly to 2 sig figs
651 g ÷ 132.87 cm^3	651 g has 3 sig figs and 132.87 cm has 5 sig figs
= 4.89952585 g cm^{-3}	The answer from a calculator (suggests 9 sig figs)
= 4.90 g cm^{-3}	The final answer rounded properly to 3 sig figs

Sample Regents Style Question:

A sample of an element has a mass of 34.261 grams and a volume of 3.8 cubic centimeters. To which number of significant figures should the calculated density of the sample be expressed?

(1) 5 (2) 2 (3) 3 (4) 4

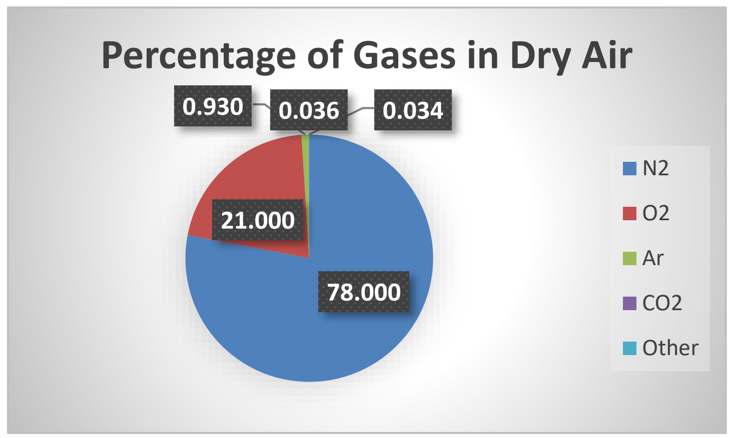
Analyzing data like a scientist:

Many times, scientists must draw conclusions based on a vast amount of data

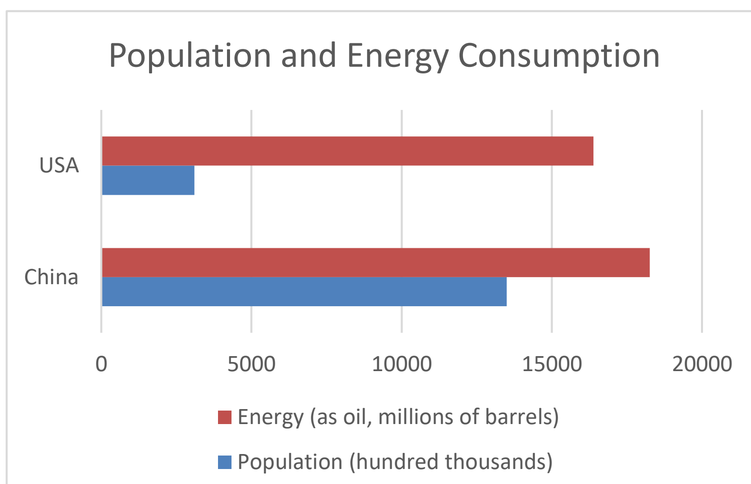
- Charts and Tables allow scientists to organize data to help us understand large amounts of data
- Charts, Tables, and Graphs allow scientists to summarize data
- Charts, Tables, and Graphs allow for interpolation and extrapolation of data
- Graphs help scientists find relationships and also help communicate large amounts of data

Graphing

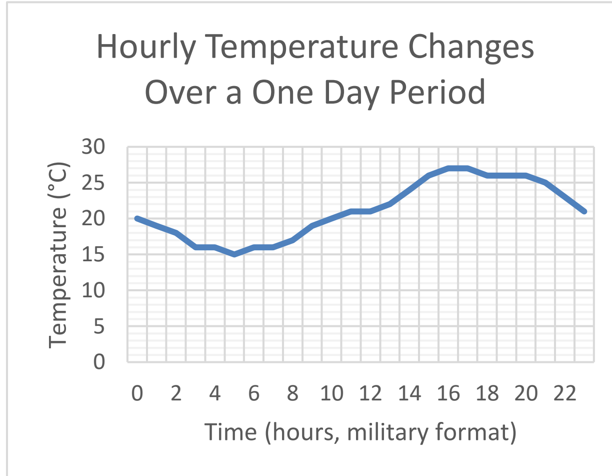
Circle or Pie Charts: best used to represent parts of a fixed whole (percent values) and these graphs do not show changes over time



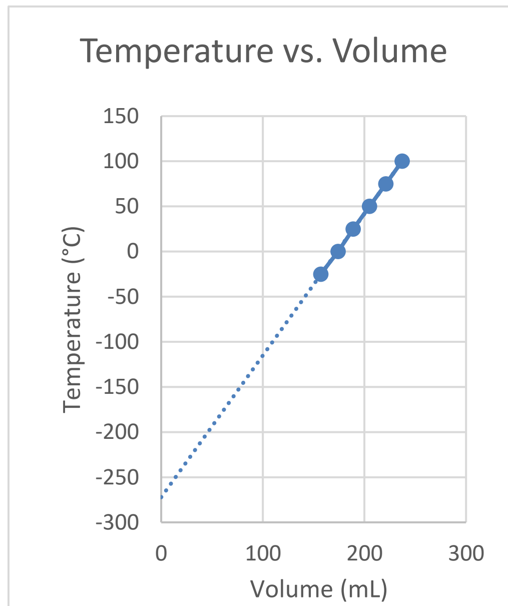
Bar Graphs or Histograms: best used to represent how quantities vary across categories (comparing US population and resource usage to China's population and resource usage) and are best used when the changes are fairly large



Line plots: best used to track small changes over short or long time periods or to compare several items that change over the same time period or to compare variables when order is important



Sc**atter plots or X–Y plots**: best used to determine the relationships between variables (look for correlations) – look for positive vs. negative, direct vs. inverse, or exponential relationships



When plotting (especially) line or scatter plots:

Always put the independent (manipulated) variable on the x-axis

Always put the dependent (responding) variable on the y-axis

Titles on graphs are usually Y (dependent variable) versus X (independent variable)

Slope equation for a straight line graph:

$$\text{slope} = \frac{\text{rise}}{\text{run}} = \frac{\Delta y}{\Delta x} = \frac{y2 - y1}{x2 - x1}$$

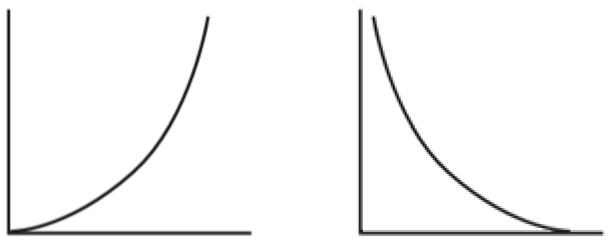
Interpolation: determine a value between two measured points

Extrapolation: determine a value beyond the measured points

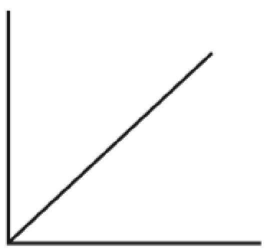
Types of Graphical Relationships

Constant

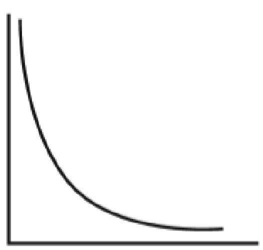
Exponential



Direct



Inverse



Question: What is different about exponential decay and inverse relationships?

Answer: Exponential decay can drop to zero while inverse relationships asymptote.

Percent error:

$$\% \text{ error} = \frac{\text{measured value} - \text{accepted value}}{\text{accepted value}} \times 100$$

- See Table T
- Always include the "× 100" in a set-up

Example:

A student measures the aspirin in a tablet and finds 272 mg when the actual amount of aspirin in a tablet is known to be 275 mg. Find the percent error.

$$\% \text{ error} = \frac{\text{measured value} - \text{accepted value}}{\text{accepted value}} \times 100 = \frac{272 \text{ mg} - 275 \text{ mg}}{275 \text{ mg}} \times 100 = -1.09\%$$

The negative sign in the answer indicates the measured value was too low.

Density:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad \text{or} \quad d = \frac{m}{V}$$

- See Table T for the formula and Table S for densities of the elements
- Units

Quantity	Variable	Symbol	Name
mass	m	g *or* kg	gram *or* kilogram
volume	V	mL *or* cm^3	cubic centimeter *or* milliliter
density	d	g/cm^3	gram per cubic centimeter

Example:

What is the volume of a solid lump of copper having a mass of 26.5 grams?

$d = 19.3\ g/cm^3$ (Table S)

$V = ?$

$m = 26.5\ g$

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

$$V = \frac{m}{d} = \frac{26.5\ g}{19.3\ g/cm^3} = 1.37\ cm^3$$

Dimensional analysis (or conversion factors):

$$\text{starting unit} = \frac{\text{ending unit}}{\text{starting unit}}$$

- Always start with the quantity you know
- Arrange units as a fraction with starting units in the denominator so they will cancel
- Use multiple fractions or conversion factors if necessary

Example:

Someone turns 62 years old on their birthday. How many seconds have they lived?

$$62 \text{ years} \times \frac{365.25 \text{ days}}{1 \text{ year}}$$

(it would have been okay to use 365 days)

$$62 \text{ years} \times \frac{365.25 \text{ days}}{1 \text{ year}} \times \frac{24 \text{ hours}}{1 \text{ day}}$$

$$62 \text{ years} \times \frac{365.25 \text{ days}}{1 \text{ year}} \times \frac{24 \text{ hours}}{1 \text{ day}} \times \frac{60 \text{ minutes}}{1 \text{ hour}}$$

$$62 \text{ years} \times \frac{365.25 \text{ days}}{1 \text{ year}} \times \frac{24 \text{ hours}}{1 \text{ day}} \times \frac{60 \text{ minutes}}{1 \text{ hour}} \times \frac{60 \text{ seconds}}{1 \text{ minute}} = 1\,956\,571\,200 \text{ seconds}$$

using significant figures, 2.0×10^9 seconds

using 365 days gives 1 955 232 000 seconds or 2.0×10^9 seconds

Metric Conversions using dimensional analysis (or conversion factors)

Table C
Selected Prefixes

(bigger)

Factor	Prefix	Symbol
10^{3}	kilo	k
10^{0}	NONE	g, L, m, s
10^{-1}	deci-	d
10^{-2}	centi-	c
10^{-3}	milli-	m
10^{-6}	micro-	µ
10^{-9}	nano-	n
10^{-12}	pico-	p

(smaller)

It takes a lot of smaller units to make up a bigger unit (100 mL = 1 dL)

- Always start with the quantity you know
- Arrange units as a fraction with starting units in the denominator so they will cancel
- Get the numbers and units from Table C
- Use multiple fractions or conversion factors if necessary

Example:

How many grams are represented by 12 kg?

$$12 \text{ kg} \times \frac{(10^{3} \text{ g})}{(1 \text{ kg})} = 12 \times 10^{3} \text{ g} = \underline{1.2 \times 10^{4} \text{ g}}$$

How many milligrams are represented by 167 nanograms?

$$167 \text{ ng} \times \frac{(10^{-9} \text{ g})}{(1 \text{ ng})} \times \frac{(1 \text{ mg})}{(10^{-3} \text{ g})} = 167 \times 10^{-9} \div 10^{-3} \text{ g} = 167 \times 10^{-6} \text{ mg} = \underline{1.67 \times 10^{-4} \text{ mg}}$$

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Unit 0

Periodic Table:

- See the Periodic Table of the Elements (between Tables R and S)

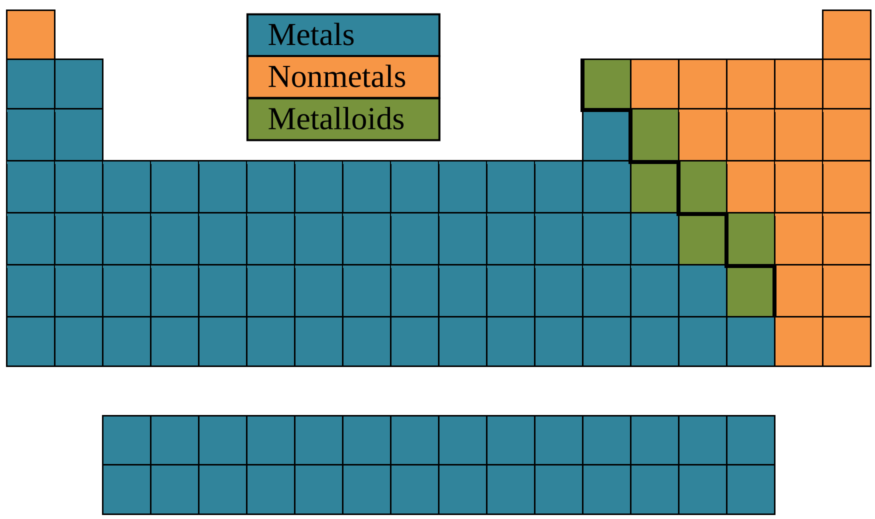
Groups and Periods

Vertical columns are called groups or families
Example: list the elements in Group 17
fluorine, chlorine, bromine, astatine, and Uus (tennessine)

Horizontal rows are called periods or series
Example: list the elements in Period 2
lithium, beryllium, boron, carbon, nitrogen, oxygen, fluorine, and neon

Metals, Nonmetals, and Metalloids

Metals comprise about 80% of the Periodic Table
Nonmetals represent another 15% of the Periodic Table
Metalloids (with properties between metals and nonmetals) are the remainder



Example: Which list includes a metal, a nonmetal, and a metalloid?
(1) Fe, Co, Ni (2) B, C, N (3) Al, P, Si (4) K, Ca, Sc

Blocks