Chemistry 115

Chapter 1

Qualitative & Quantitative Measurement

- We can make qualitative observations about reactions:
  - E.g. Changes in colour, texture or state, comparisons

- We can also make quantitative measurements
  - E.g. Quantified changes in temperature or mass

Quantitation

- In order to make quantitative measurements, we need to be clear about the numbers that we use:
  - How do we write numbers that are very big and very small?
  - How many of the numbers from the calculator screen do we include in our quiz answer?

Scientific Notation

- Large and small numbers: Scientific Notation
- Example: Avogadro's Number:
  - 602 214 129 000 000 000 000 000 /mol
  - Write as a **number between 1 and 10**, multiplied by 10 to the appropriate power:
  - $6.02214129 \times 10^{23}$ /mol
- For numbers less than 1, use negative powers of 10

Significant Figures

- How many numbers: Significant Figures

- Your result or answer can only be as precise as the least precise number in the calculation

  - $0.6754 \times 6.02 \times 10^{23}$ / mol = 406590800000000000000000? NO
  - $0.6754 \times 6.02 \times 10^{23}$ / mol = 4.065908×10^{23}? NO
  - $0.6754 \times 6.02 \times 10^{23}$ / mol = 4.07×10^{23}? NO
  - $0.6754 \times 6.02 \times 10^{23}$ / mol = 4.07×10^{23} / mol? YES

Significant Figures

- How many numbers: Significant Figures
- The rules
  - All numbers that are not zero are always significant
  - Zeros at the beginning of a number are never significant – they just show where the decimal point is
  - Zeros at the end of a number before a decimal point are never significant
  - Zeros at the end of a number after a decimal point are always significant
  - Zeros between two other/significant digits are always significant

Significant Figures

- How many numbers: Significant Figures

- For complex calculations with many steps:
  - Retain more than the number of significant numbers needed throughout the calculation (e.g., if 3 s.f. are given, use 4 or 5 at each step)

  - First, write the final answer with more significant figures than required

  - Then, write the final answer with the correct number of significant figures at the end

Accuracy vs Precision

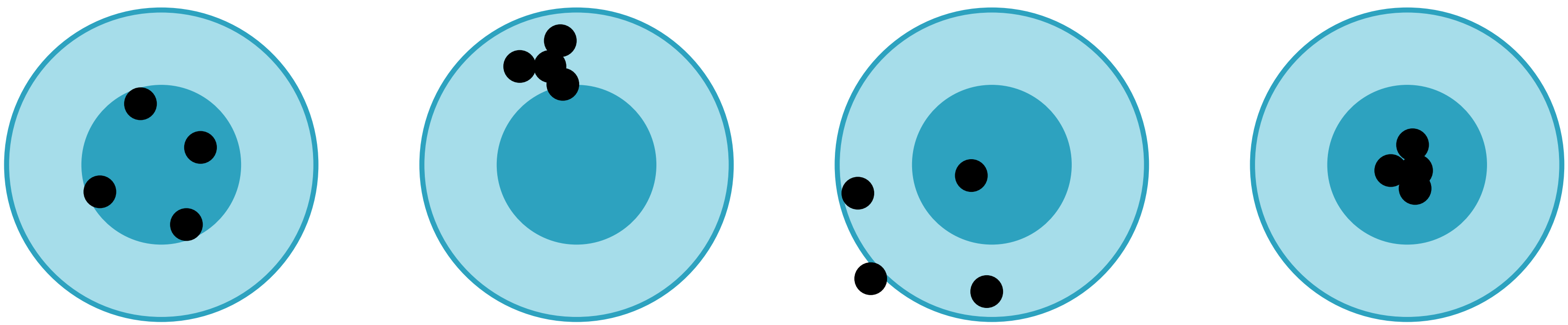
- In order to make quantitative measurements, we need to be clear about the accuracy and units of the measurements we make:

  - What is the difference between accuracy and precision?

  - What units should we use, and why?

Accuracy vs Precision

- Accuracy and Precision



Precision	*x*	*Precision*	☺	*Precision*	x	*Precision*	☺
Accuracy	☺	*Accuracy*	*x*	*Accuracy*	x	*Accuracy*	☺

Errors – Absolute

- Absolute experimental error – a measure of accuracy

$$Error = (measured\ value) - (accepted\ value)$$

- **Example:** I put a thermometer in the kettle, and measure the temperature of boiling water as 100.9°C
  - What is the error in this measurement? 0.9°C

Errors – Percentage

- Percentage error – a measure of accuracy

$$\% \, Error = \left(\frac{error}{accepted\ value}\right) \times 100$$

- (Same) **Example:** I put a thermometer in the kettle, and measure the temperature of boiling water as 100.9°C
  - What is the percentage error in this measurement? 0.9%

Standard Deviation

- What: A measure of the spread (precision) of measurements

$$Standard\ Deviation(\sigma) = \sqrt{\frac{\sum_{1-n}(x-\bar{x})^2}{n-1}}$$

- Small value of standard deviation = precise
- Large value of standard deviation = imprecise

- WRITE THIS DOWN –EQUATION IS **NOT** IN THE TEXTBOOK!

Standard Deviation

- Example: Again, measuring the temperature of water, to brew English Breakfast Tea, which requires a temperature of 100°C

- I measure the temperature six times (because tea is important)

Images from Microsoft

Standard Deviation

- Measured values:
  100.9℃, 99.8℃, 101.1℃, 99.5℃, 99.4℃, 100.1℃

- **Strategy:**
  - Calculate average ($\bar{x}$)
  - Calculate $(x - \bar{x})$ for each value
  - Calculate $(x - \bar{x})^2$ for each value, and then the sum $\sum(x - \bar{x})^2$
  - Calculate standard deviation (σ)

- Easiest way to do this is to make a table

Standard Deviation

- Measured values:
  100.9°C, 99.8°C, 101.1°C, 99.5°C, 99.4°C, 100.1°C

- Calculate $\bar{x}$

  - $\bar{x}$ = (100.9°C+99.8°C+101.1°C+99.5°C+99.4°C+100.1°C)/6
    = 100.133°C

Standard Deviation

- Measured values:
  100.9°C, 99.8°C, 101.1°C, 99.5°C, 99.4°C, 100.1°C

- $\bar{x} = 100.1$°C (rounded to 4 s.f.)

x / °C	$(x - \bar{x})$ / °C	$(x - \bar{x})^2$ / °C
100.9		
99.8		
101.1		
99.5		
99.4		
100.1		

Standard Deviation

- Measured values:
  100.9℃, 99.8℃, 101.1℃, 99.5℃, 99.4℃, 100.1℃

- $\bar{x} = 100.1$℃

x / ℃	$(x-\bar{x})$ / ℃	$(x-\bar{x})^2$ / ℃
100.9	0.8000	
99.8	-0.3000	
101.1	1.000	
99.5	-0.6000	
99.4	-0.7000	
100.1	0.000	

Standard Deviation

- Measured values:
  100.9°C, 99.8°C, 101.1°C, 99.5°C, 99.4°C, 100.1°C

- $\bar{x} = 100.1$°C

x / °C	$(x - \bar{x})$ / °C	$(x - \bar{x})^2$ / °C
100.9	0.8000	0.6400
99.8	−0.3000	0.09000
101.1	1.000	1.000
99.5	−0.6000	0.3600
99.4	−0.7000	0.4900
100.1	0.000	0.000

Standard Deviation

- Measured values:
  100.9℃, 99.8℃, 101.1℃, 99.5℃, 99.4℃, 100.1℃

- $\bar{x} = 100.1$℃

x / ℃	$(x - \bar{x})$ / ℃	$(x - \bar{x})^2$ / ℃
100.9	0.8000	0.6400
99.8	−0.3000	0.09000
101.1	1.000	1.000
99.5	−0.6000	0.3600
99.4	−0.7000	0.4900
100.1	0.000	0.000

- $\sum_{1-n}(x - \bar{x})^2 = 2.580$ ℃^{2}

Standard Deviation

- Answer

$$Standard\ Deviation(\sigma) = \sqrt{\frac{\sum_{1-n}(x-\bar{x})^2}{n-1}}$$

$$\sigma = \sqrt{\frac{2.58\ {}^\circ C^2}{5}} = 0.71833{}^\circ C$$

$$\sigma = 0.718{}^\circ C$$

Units

- In order to be clear about the numbers that we use, we must ALWAYS state the units

- Consistent and explicit

- Always state the units of any number that you write down

- Always be certain that numbers describing the same property (mass, temperature etc) are in the same units before you do any calculation

Units: SI System

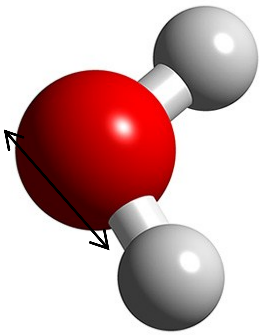
- Scientific measurements are usually made in SI units
  - Metric: powers of 10 (10^6, 10^3, 10^{-3}, 10^{-6} etc)
  - Based on 7 internationally agreed base units: m, kg, s, A, K, cd and mol

- Other units, or multiples of units, may also be used in certain circumstances
  - E.g. The Watt ($1\ W = 1\ J/s$) to measure power, and the Joule ($1\ J = 1\ kg \cdot m^2/s^2$) or the electron Volt ($1\ eV = 1.60218 \times 10^{-19}\ J$) for energy

Units: Length

- Length and distance are measured in meters (m)
- Different prefixes are used to describe distances at different scales
  - 1000m = 1km (kilometer)
  - 0.001m = 1mm (millimeter)
  - 0.000000001m = 1nm (nanometer)

O–H distance
$= 9.58 \times 10^{-11}$m
= 0.0958 nm
= 95.8 pm
= 0.958 Å



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Units: Mass

- Mass (measured in grams) is a measure of how much of something you have
- Don’t use ‘mass’ and ‘weight’ interchangeably.
  - Your mass will not change if you go into orbit, but your weight (the amount of force that gravity exerts on you) will change a lot!

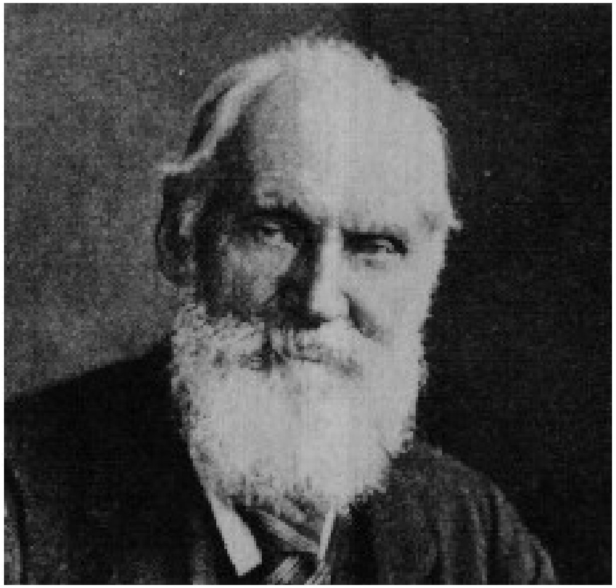
Images from Microsoft

Units: Temperature

- The SI unit for temperature is Kelvin (K)
- Other common temperature scales are Celsius/Centigrade and Farenheit

Anders Celsius
(1701–1744)

Lord Kelvin
(1824–1907)



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Units: Temperature

- Temperatures in Kelvin, Celsius and Farenheit:

	Kelvin	Celsius	Farenheit
Water boils	373K	100°C	212°F
	100K	100°C	180°F
Ice melts	273K	0°C	32°F

Units: Temperature

- Absolute zero, $0K$:
  - Thermal energy is zero
  - Minimum entropy
  - Weird quantum effects (ask the physics professors…)

Units: Temperature

- Temperature should usually be stated in Kelvin
- To convert from Celsius to Kelvin:

$$T\ (K) = T\ (°C) + 273.15$$

- Note that $|1°C| = |1K|$

$$Body\ Temperature\ = 37°C + 273 = 310K$$

$$Liquid\ Nitrogen\ = 77K\ - 273 = -196°C$$

Units: Temperature

- Temperature should usually be stated in Kelvin

- To convert from Farenheit to Kelvin:

$$T\ (K) = \frac{5K(T\ (°F) - 32°F)}{9°F} + 273.15K$$

- Note that $|1°\mathrm{F}| < |1\mathrm{K}|$

Unit Conversion – Dimensional Analysis

- In general, conversions from one unit to another should be carried out using the dimensional analysis method:

$$\text{Number in original unit}\ \underbrace{\left[\frac{\text{new unit}}{\text{original unit}}\right]}_{\text{Conversion factor}} = \text{new number in new unit}$$

Quantity to express in new units

Quantity now expressed in new units

- **Example:** Convert 14 feet to metres, using the conversion factor 1 foot = 0.33 m

Properties: Physical and Chemical

- Physical properties can be measured without changing composition
  - Examples: State (solid, liquid or gas), temperature…

- Chemical properties can only be measured by performing a reaction, and therefore changing composition
  - Examples: Burning, rusting…

Properties: Intensive and Extensive

- Intensive properties are not dependent on how much of something you have
  - Examples: Density, temperature…
- Extensive properties depend on the quantity of the thing
  - Examples: Mass, volume…

Physical Properties

- What are some other physical properties of matter?
  - Colour
  - Odour
  - Melting / boiling temperature
  - Tensile strength
- Please do not use ‘odour’ to identify chemicals in the lab. You will end up with a headache or worse!

Physical Properties: States

- The three simplest states of matter:
- SOLID: Rigid shape, fixed volume, shape may reflect the atomic/molecular structure
- LIQUID: No fixed shape, does not expand to fill the container
- GAS: Expands to fill the container

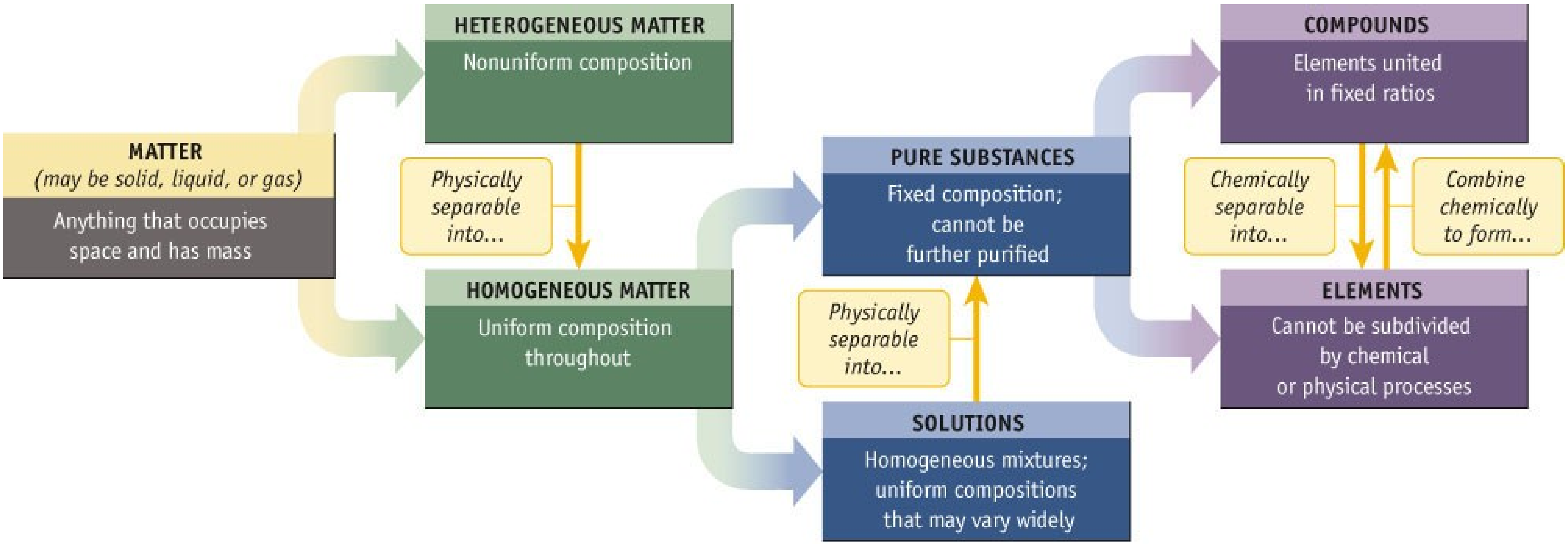
Physical Properties: States

- The three simplest states of matter in kinetic terms:
- SOLID: Atoms/molecules vibrate, but do not ‘translate’ through the material
- LIQUID: Atoms/molecules vibrate, rotate and translate
- GAS: Atoms/molecules vibrate, rotate and translate, with greater kinetic energy than in a liquid

States of Matter

- Examples
- Solids – molecular structure affects properties (carbon)
- Liquids – solutions and suspensions (tea versus latte)
- Mixtures – homogeneity and heterogeneity

Classifying Matter



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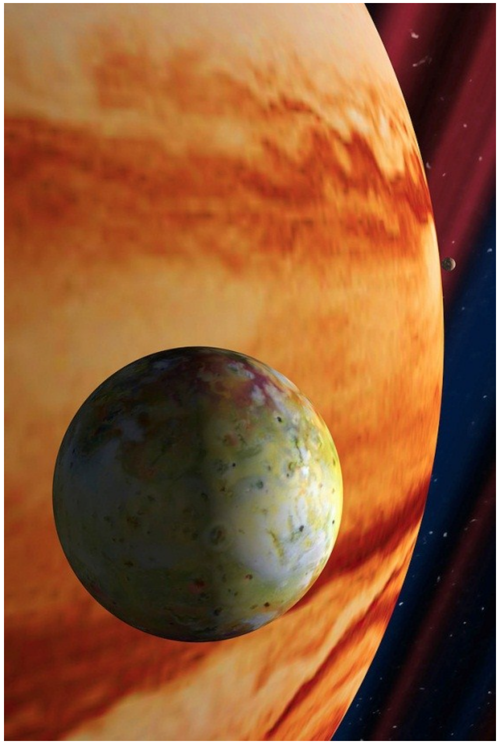
States of Matter

- Solid, liquid or gas?
- Heterogeneous or homogeneous?



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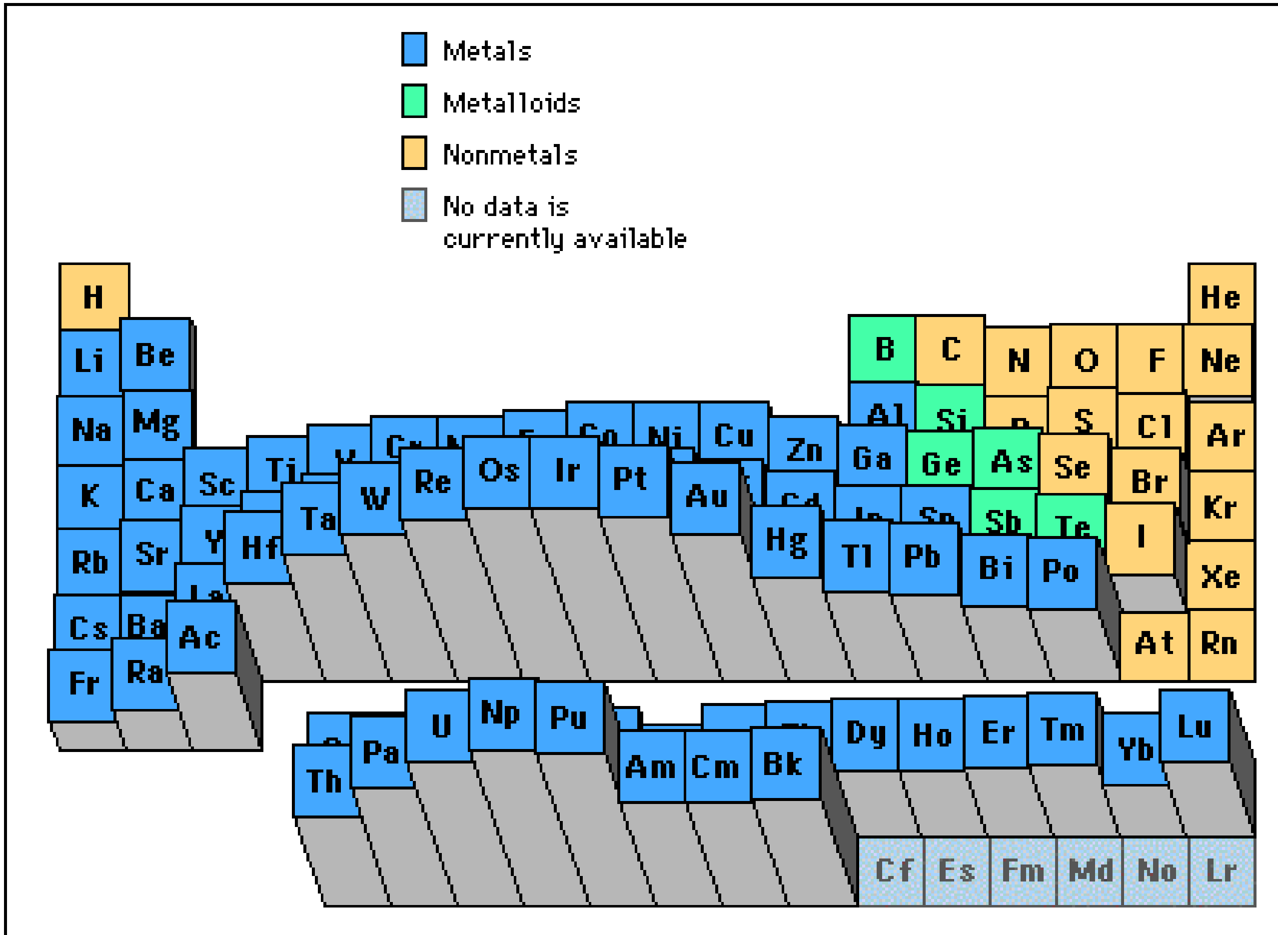
Properties: Density

- Density is a particularly useful property, defined as:

$$Density = \frac{mass\ (g)}{volume\ (cm^3)}$$

Properties: Density

- Relative densities of the elements:



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PROBLEM: Density

- A piece of copper has a mass of 57.54g. It is 9.36cm long, 7.23cm wide, and 0.95mm thick.

- What is its density (g/cm^3)?

ANSWER: Density

- A piece of copper has a mass of 57.54g. It is 9.36cm long, 7.23cm wide, and 0.95mm thick.

1. Put dimension values in common units

$$0.95\ mm = 0.95\ \times \left(\frac{1\ cm}{10\ mm}\right) = 0.095\ cm$$

2. Calculate volume in cubic centimetres

$$Volume = 9.36\ \times 7.23\ \times 0.095 = 6.4\ cm^3$$

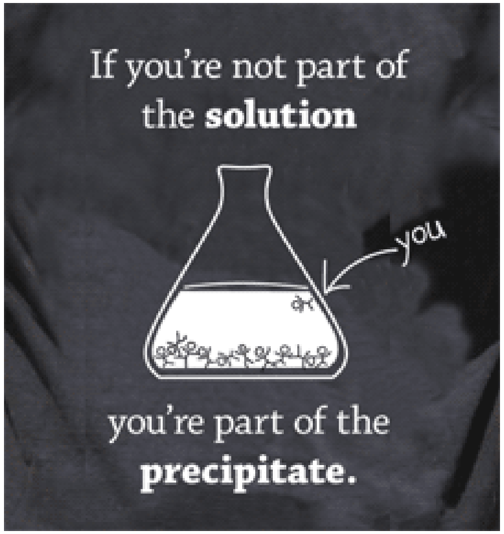
3. Calculate density in grams per cubic centimetre

$$Density = \frac{57.54\ g}{6.4\ cm^3} = 9.0\ g/cm^3$$

Properties: Physical Changes

- Examples of observable physical changes:
  - Melting
  - Boiling
  - Dissolving

- Others?



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Properties: Chemical Changes

- Definition: Changing one or more types of molecule into different types of molecule

- This process is known as a reaction:

$$Vinegar + Baking\ Soda \rightarrow Sodium\ acetate + Water + Carbon\ dioxide$$

$$CH_3COOH + NaHCO_3 \rightarrow NaCH_3COO + H_2O + CO_2$$

- We'll come back to the subject of chemical reactions in a lot more detail later in the course