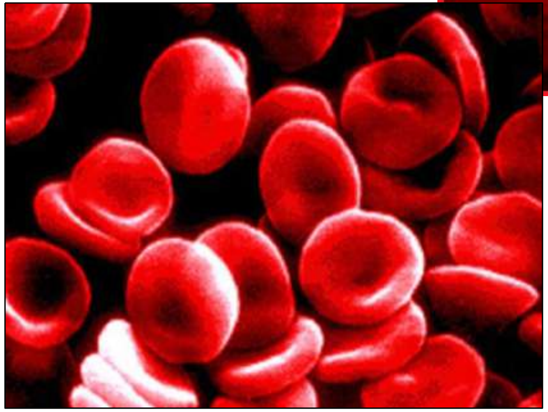
Chapter 1

AP Chemistry

What Is Chemistry

- Science of materials and the changes they undergo.
- Micro or Macro?
  - Macro - large scale
  - Micro - small scale



The Scientific Method

- The scientific method is a logical approach to solving scientific and everyday problems.
- The steps are as follows:
- Observation → Hypothesis

  Experiment → Conclusion

  Theory

Observation and Question

- An observation is a way of gathering information using one or more of your senses.
- There are two types of observations.

Hypothesis and Experiment

- A hypothesis is an educated guess to the question raised by the observation.
- An experiment is then run to test the validity of the hypothesis.
- The experiment is repeated many times to ensure correct results.

Conclusion

- A statement about if you have or have not disproved your hypothesis.
- If you have not disproved your hypothesis you may move on to theory.
- If you have disproved your hypothesis you must go back and form a new hypothesis and experiment.

Scientific Theory

- Once a scientific hypothesis passes the test of repeated experiments, it may become a theory.
- A theory explains *why* experiments give the result they do.
- A theory can never be proven because a new experiment may disprove it.

Scientific Law

- A law is a short statement of behavior or relation that always seems to be the same under the same conditions.
- A law describes natural phenomena *without* attempting to explain it.

Basic Units of SI

- These are the **basic units** of the SI system.
- You will need to memorize them, like now.

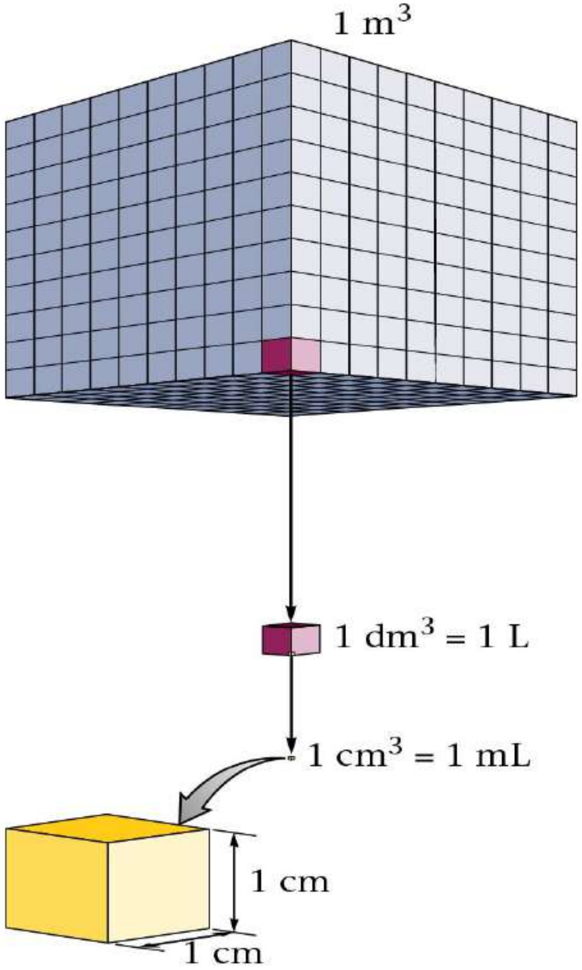
SI Base Units

Quantity	Base unit
Time	second (s)
Length	meter (m)
Mass	kilogram (kg)
Temperature	kelvin (K)
Amount of a substance	mole (mol)
Electric current	ampere (A)
Luminous intensity	candela (cd)

Prefix	Symbol	Meaning	Exponential Notation
Giga	G	1 billion	10^{9}
Mega	M	1 million	10^{6}
Kilo	K	1,000	10^{3}
Hecto	h	100	10^{2}
Deka	da	10	10^{1}
--	--	1	10^{0}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	u	1 millionth	10^{-6}
nano	n	1 billionth	10^{-9}
pico	p	1 trillionth	10^{-12}

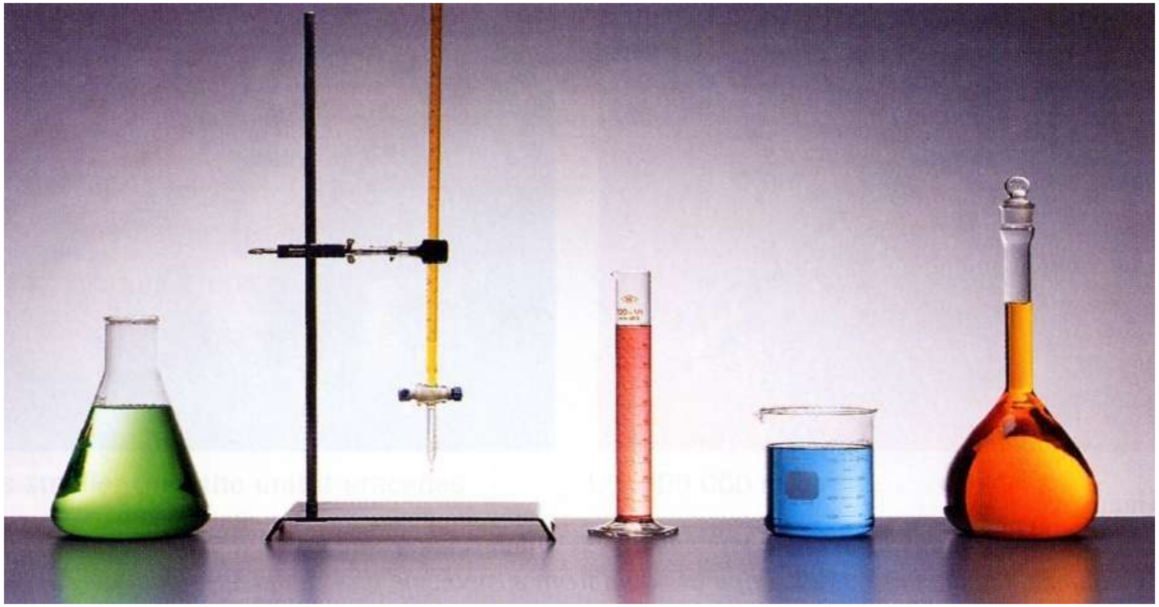
Volume

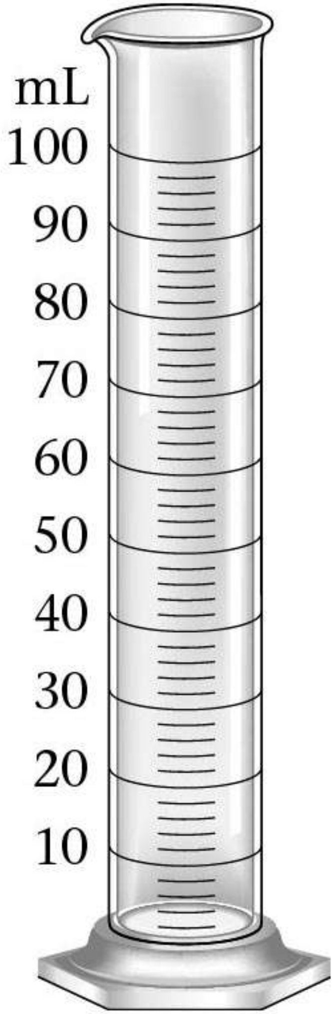
- **Volume** is how much 3D space an object takes up.
- SI unit is $\mathbf{m^3}$.
- One thousandth of which is a dm^3, aka a **liter (L)**.
- One thousandth of which is the cm^3 or **mL**.

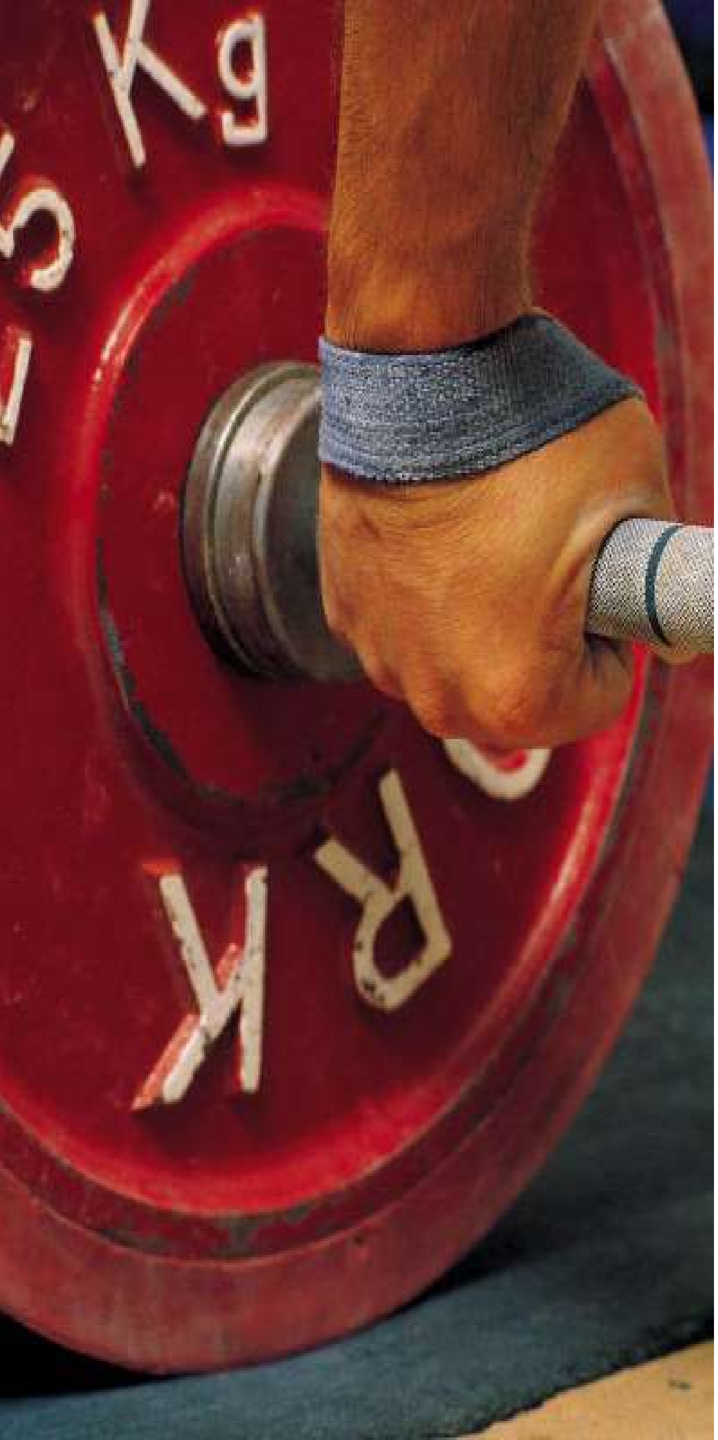


Measuring Volume

- We mostly measure V with a **graduated cylinder** but also with these, all of which are *marked* on the side.



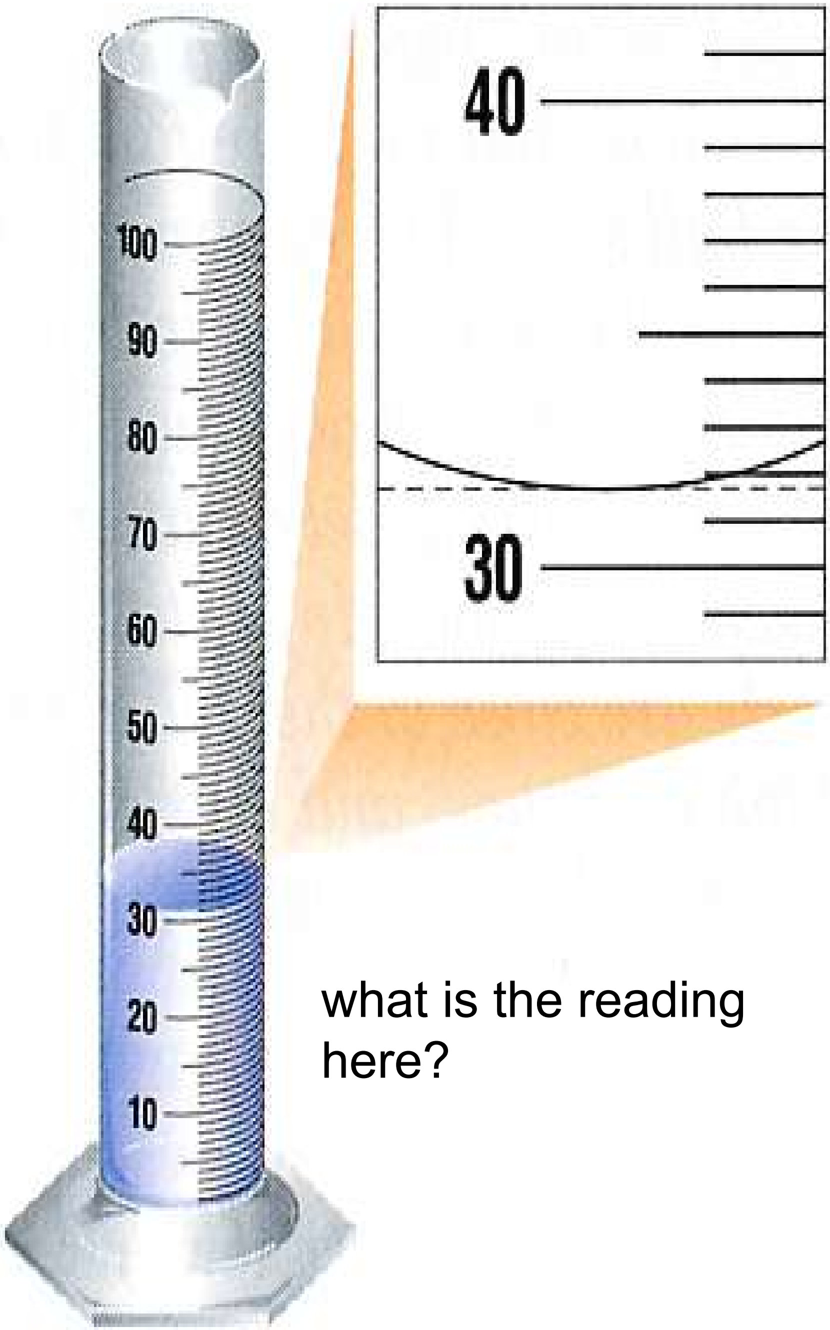




Mass vs. Weight

- **Mass** is the amount of “stuff” something has. (Resistance to change in motion)
- **Weight** is how much force that thing exerts because of gravity.
- *What if no gravity?*

5.4 uncertainty in measurement



what is the reading here?

- Many measurements are made of objects that make us **estimate**
- So, we'll always **argue about the last number or two.**
- The ones we *agree* on are called **certain**, the *argued* ones are **uncertain**.

Uncertainty

- Every measuring device has some degree of uncertainty
- *The certain numbers* + *the one uncertain #* are called **significant figures**

31.8

31.7

31.8

31.8

31.6

31.7

31.7

5.5 Significant Figures

- This balance displays 89.2863 grams and not just 89 grams **for a reason!!!**
- You must record the certain numbers and the one uncertain one as your data.

Rules for Counting Significant Figures

1. *Nonzero integers.* Nonzero integers *always* count as significant figures. For example, the number 1457 has four nonzero integers, all of which count as significant figures.

2. *Zeros.* There are three classes of zeros:

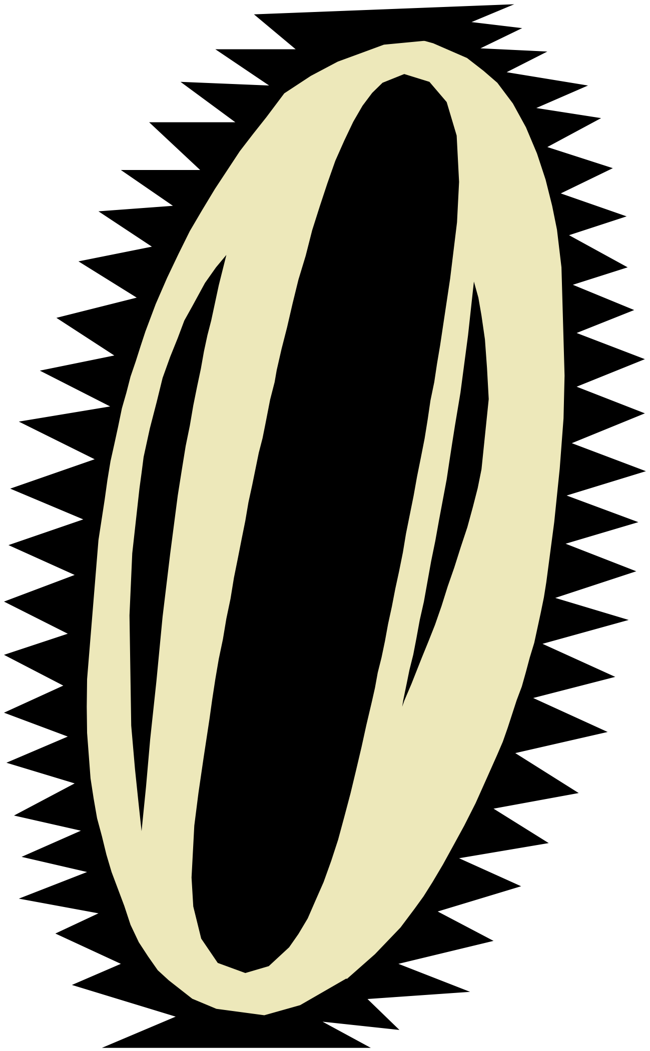
   a. *Leading zeros* are zeros that *precede* all of the nonzero digits. They *never* count as significant figures. For example, in the number 0.0025, the three zeros simply indicate the position of the decimal point. The number has only two significant figures, the 2 and the 5.

   b. *Captive zeros* are zeros that fall *between* nonzero digits. They *always* count as significant figures. For example, the number 1.008 has four significant figures.

   c. *Trailing zeros* are zeros at the *right end* of the number. They are significant only if the number is written with a decimal point. The number one hundred written as 100 has only one significant figure, but written as 100., it has three significant figures.

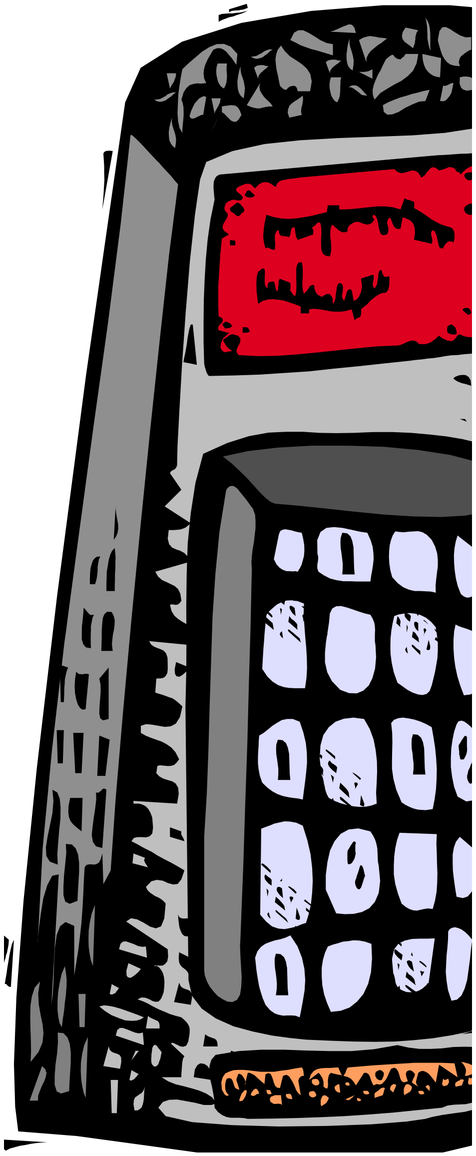
zero translation…

- Front? Never!
- Within? Always!
- End? Only if.



Sigfig Examples

- *The mass of an eyelash is 0.000304 g*
- **3**
- *The length of the skidmark was 1.270 x 10^2 m*
- **4**
- *A 125-g sample of chocolate chip cookie contains 10 g of chocolate*
- **3, 1**
- *The volume of soda remaining in a can after a spill is 0.09020 L*
- **4**
- *A dose of antibiotic is 4.0 x 10^{-1} cm^3*
- **2**



One More Thing…

3. *Exact numbers.* Often calculations involve numbers that were not obtained using measuring devices but were determined by counting: 10 experiments, 3 apples, 8 molecules. Such numbers are called *exact numbers*. They can be assumed to have an unlimited number of significant figures. Exact numbers can also arise from definitions. For example, 1 inch is defined as *exactly* 2.54 centimeters. Thus, in the statement 1 in. = 2.54 cm, neither 2.54 nor 1 limits the number of significant figures when it is used in a calculation.

Multiplying and Dividing

- *Answers will have as many sigfigs as the working number w/ the* ***FEWEST***
- **Examples**
- 2.34 • 3.2 = 7.488?
  - Smallest number of s/d is **2** so **7.5**
- 35.0 / 6.734 = 5.1975051975?
  - Smallest number of s/d is **3** so **5.20**

Addition and Subtraction

- ***First add them up! Don't worry about sigfigs until the end!***
- 3.75 + 4.1 = 7.85
  - ***You can only go to where all numbers have something to contribute****, so can only go to* **7.9**
- 3.987 + 4.60 = 8.587
  - But can only go to 0.01, so **8.59**

Precision vs. Accuracy

- Accuracy: Agreement with true value
- Precision: Agreement between measurements
- Random Error: Equal probability of being high or low. (Estimating)
- Systematic Error: Same direction each time. (Poor technique or inaccurate device)

5.6 Dimensional Analysis

- Converting one measurement into another.
- Uses equivalency statements (fancy term for two things that mean the same thing)
- Examples of equivalency statements:
  - 1ft. = 12in.5280ft. = 1mi.
  - 10^6m = 1Mm1000m = 1km
  - 1m = 100cm1m = 1000mm
  - 1m = 10^6μm1m = 10^9nm
  - 1cm = 10mmand on and on and on

Steps of Dimensional Analysis

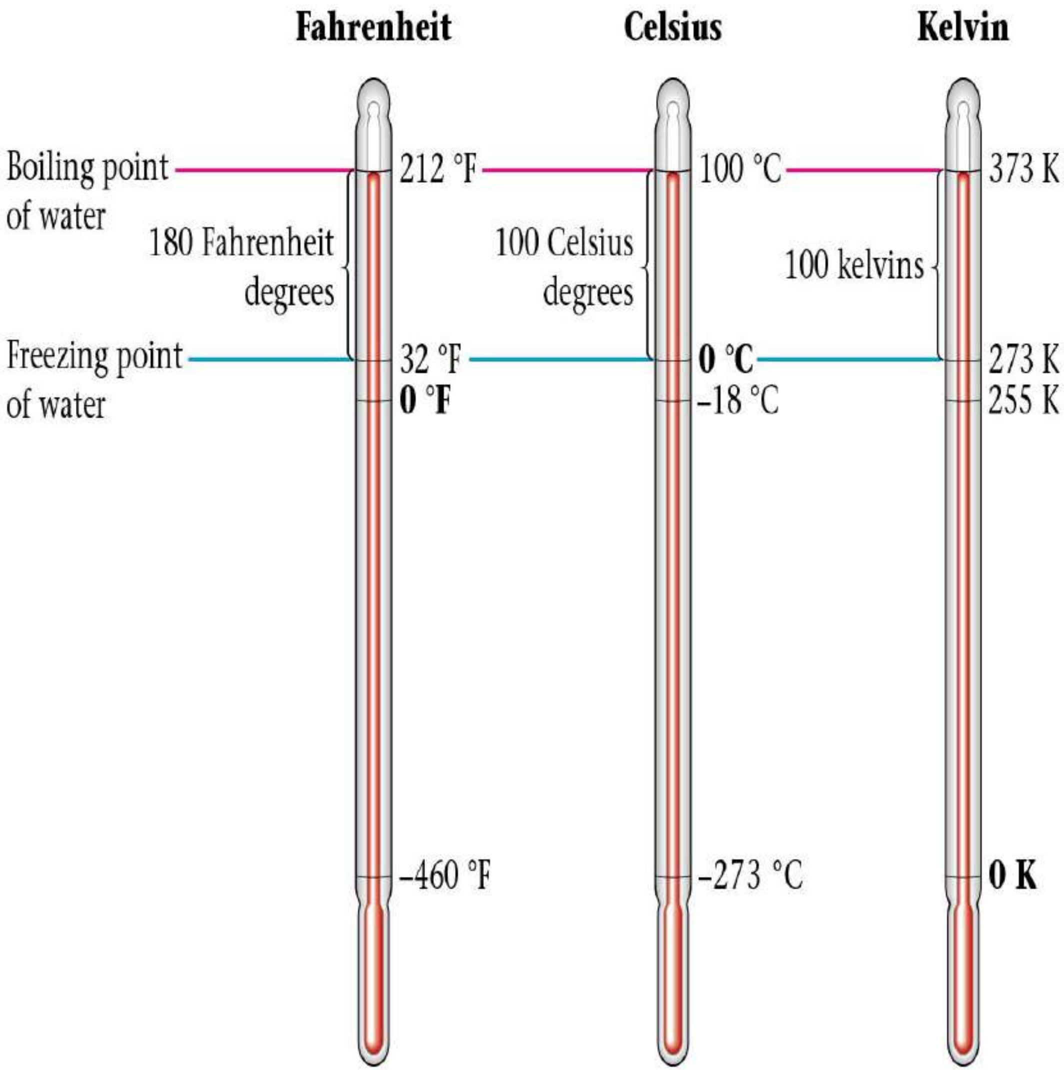
- 1. Write what you know (given)
- 2. Write where you are going (wanted)
- 3. Write a large H.
- 4. Determine conversion factor(s).
- 5. Fill in units (unit of given is on the bottom of the first H)
- 6. Fill in numbers of conversion factor.
- 7. Multiply the given by everything on the top and divide by everything on the bottom.

Examples

- 6.5dz = ___donuts
- 10.ft = ___in
- 10,000ft = ___mi
- 5m = ___cm
- 1050mg = ___g
- 17.4mi = ____ft
- 17.4mi = ____in
- 45 min = ____sec

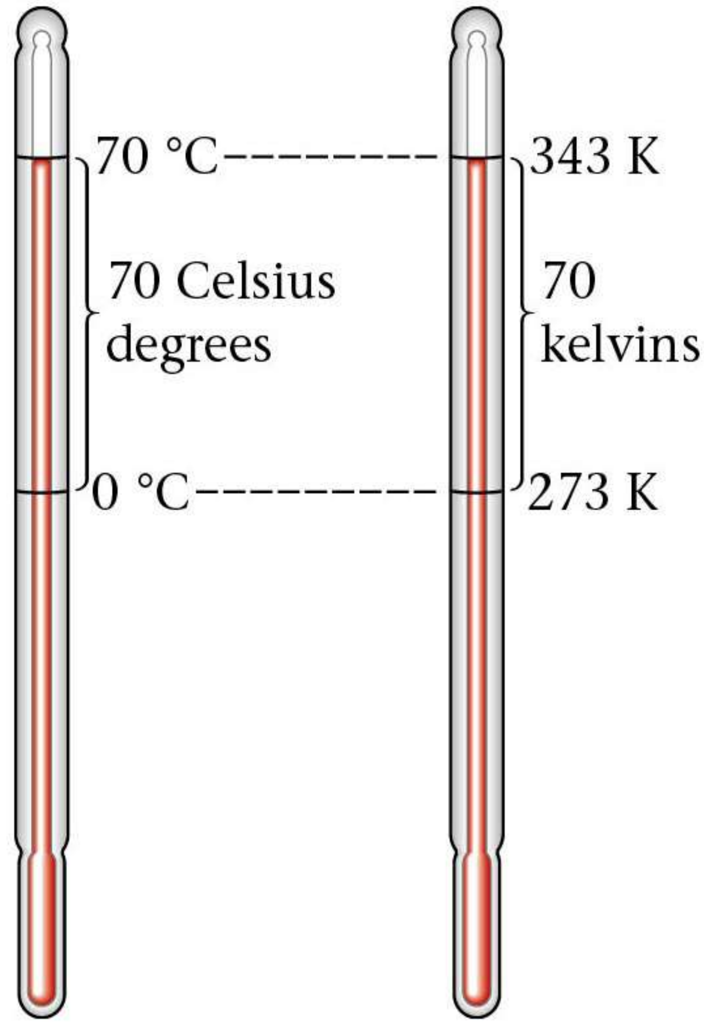
5.7 Temperature Conversion

- Big Three Temp Scales are **Fahrenheit**, **Celsius**, and **Kelvin**
- In science we use almost exclusively C and K



Converting Between K and C

- A degree C and K are the same amount; *they just differ by their starting points*
- They only differ by 273
- **$T_C + 273 = T_K$**



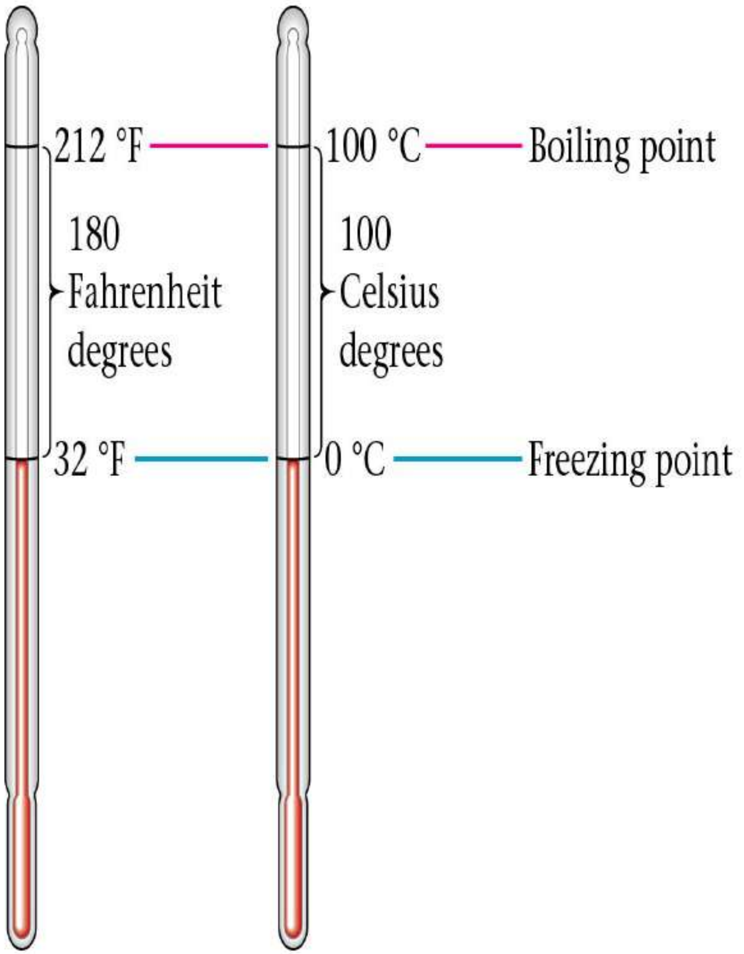
Conversion Problems

- *What is 70°C in kelvins?*
- $T_C + 273 = T_K$
- $70 + 273 = T_K$
- **343 K** $= T_K$
- *Nitrogen boils at 77 K. What is that in C?*
- $T_C + 273 = T_K$
- $T_C = T_K - 273$
- $T_C = 77 - 273$
- $\mathbf{T_C =}$ **-196 °C**

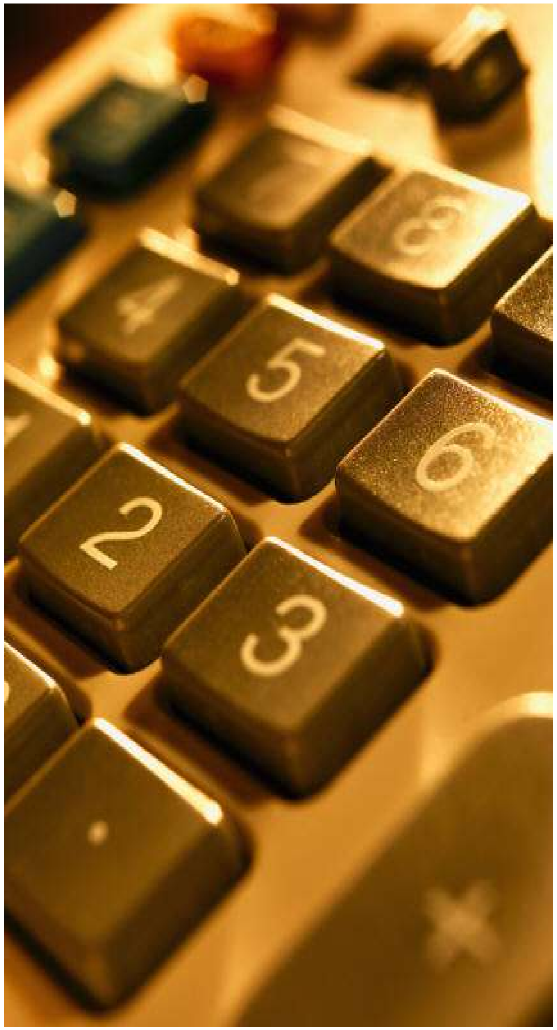
Converting Between C and F

- Here we have different size units and different starting points! yikes!
- short story:

$$T_F = 1.8T_C + 32$$

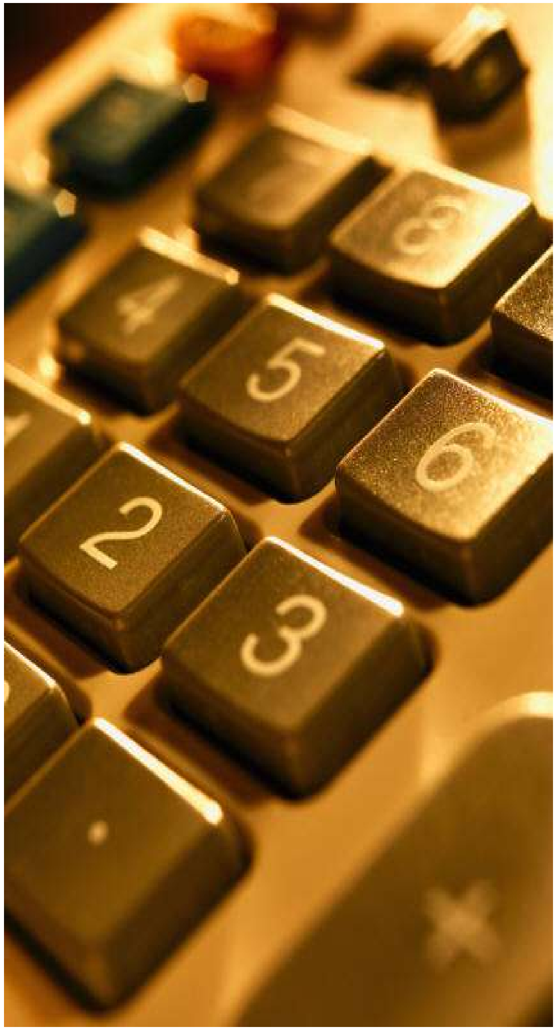


Examples



- *It's 28°C outside. What is that in F?*
- **$T_F = 1.8T_C + 32$**
- $T_F = 1.8(28) + 32$
- $T_F = 50. + 32$
- **$T_F = 82$ °F**
- *It's -40.°C in that lab freezer. What's that in F?*
- **$T_F = 1.8T_C + 32$**
- $T_F = 1.8(-40.) + 32$
- $T_F = -72 + 32$
- **$T_F = -40$°F (!)**

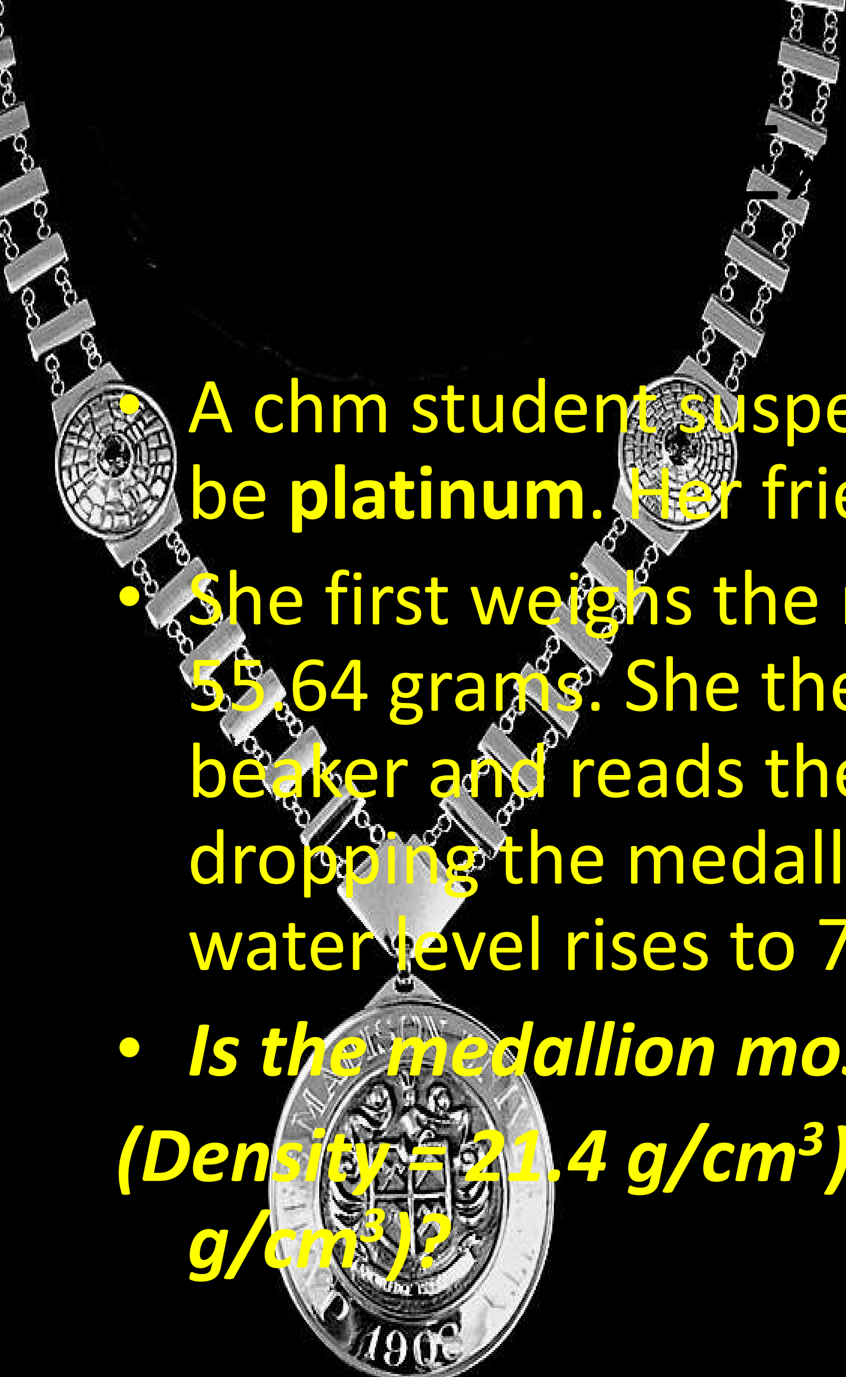
More Examples



- You have a 101°F fever. What is that in C?
- $T_F = 1.8T_C + 32$
- $101 = 1.8T_C + 32$
- $69 = 1.8T_C$
- $38° = T_C$

5.8 Density

- **Density** is just how much stuff is crammed into a certain space
- In science speak it's mass/volume:
- **$D = m/V$**
- Finding mass is no problem; how do you find volume?



- A chm student suspects that a medallion might be **platinum**. Her friend thinks it is **silver**.
- She first weighs the medallion and finds it to be 55.64 grams. She then places some water in a beaker and reads the volume at 75.2 mL. After dropping the medallion into the cylinder the water level rises to 77.8 mL.
- ***Is the medallion most likely Pt***

(Density = 21.4 g/cm^3) or Ag (Density = 10.5 g/cm^3)?

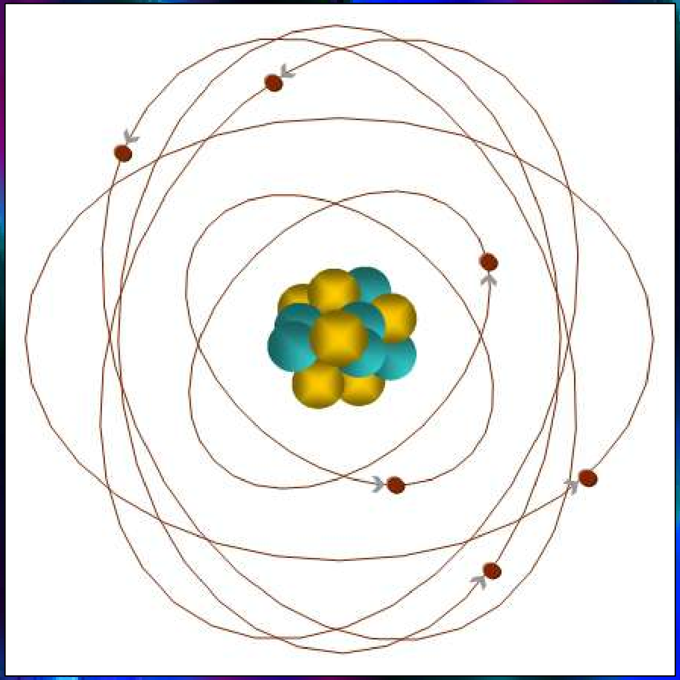
Example 2

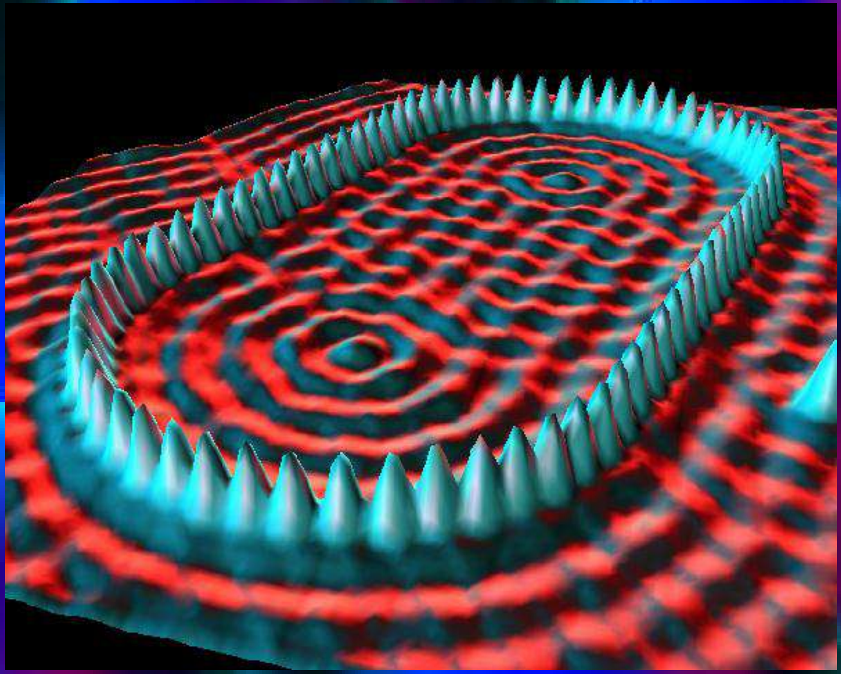
- A student needs 450 cm^3 of salt for an experiment.
  (Density$_{salt}$ = 2.16 g/cm^3)
  How many 1-lb boxes should he buy? *(1 lb = 454.5 g)*

- All the "stuff" in the universe is composed of matter.
- Matter is anything that has mass and takes up space (has volume).
- All matter is composed of a relatively small number of fundamental particles.
  - What are these particles called?

Atoms

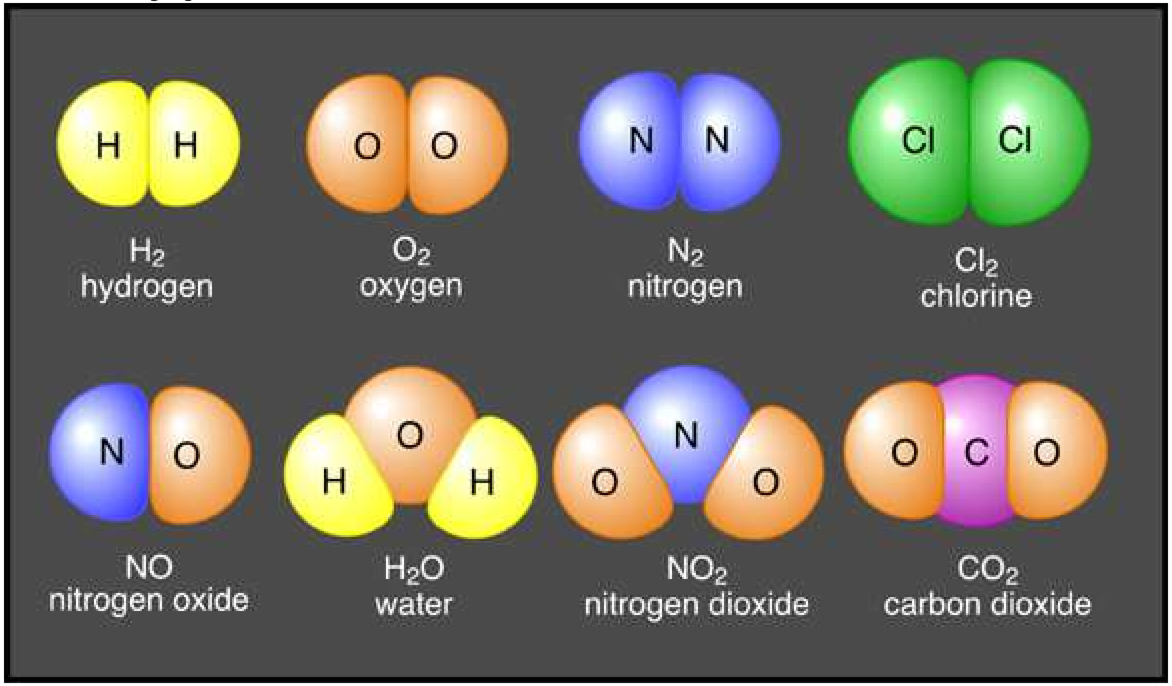
- Atoms are the smallest particles of an element that still retain the properties of that element.
- Question: Are all atoms the same?

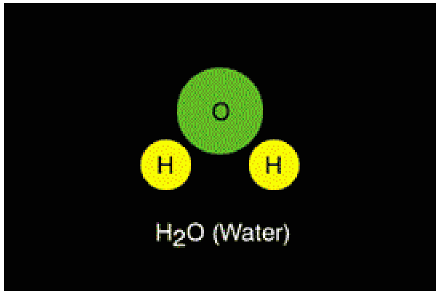


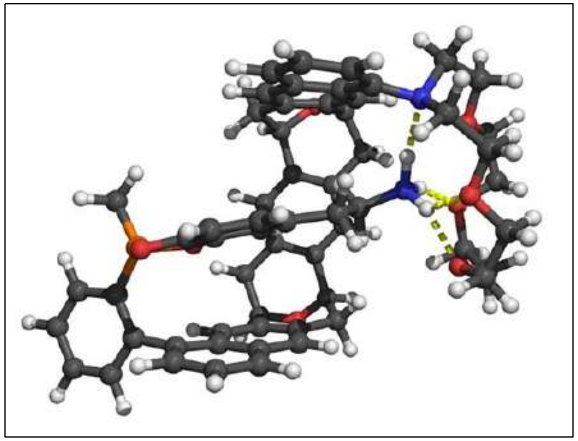


Molecules

- Particles composed of two or more atoms that are bonded together.
- Molecules can be formed from one or more types of atoms.







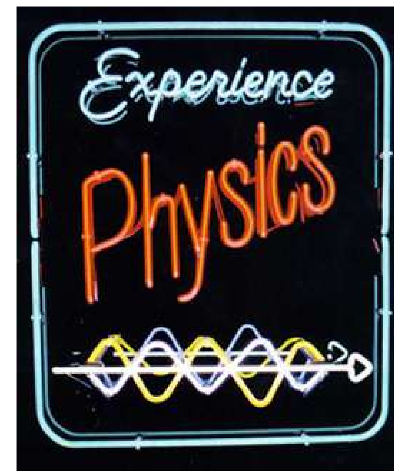
Substances

- Pure substances <u>always</u> have the same composition.
- Substances have only one type of molecule.
- Substances can be either elements or compounds.
- Pure water (only H_2O) is a substance. Naturally occurring water is not. Why not?
- Is air a substance? If not, what do we call it?

Elements

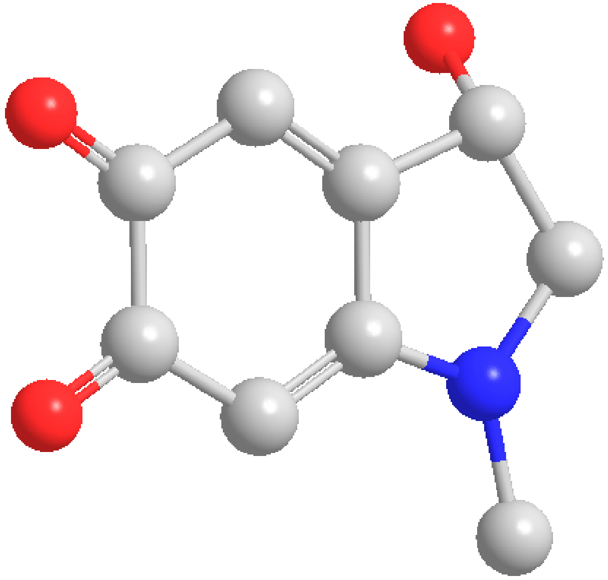
- Substances that contain only one type of atom are called elements.



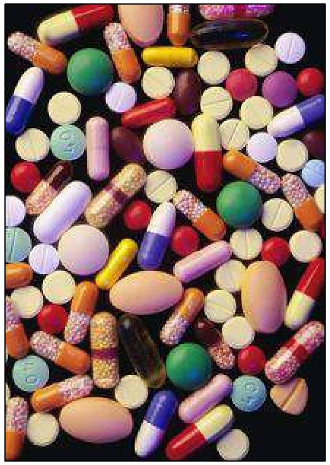


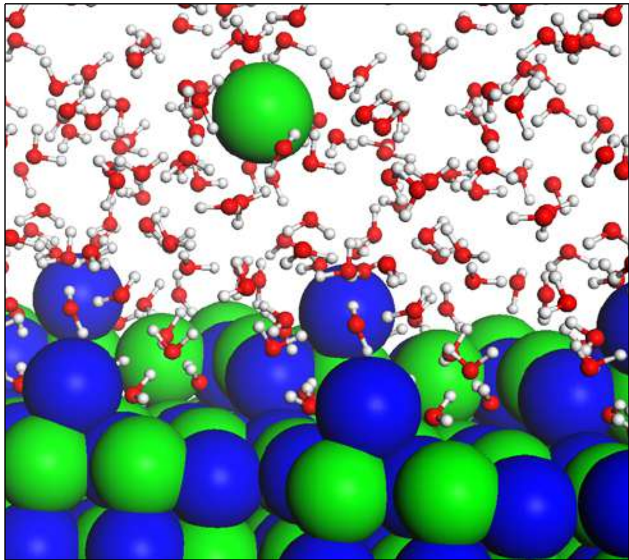



Compounds



- Substances composed of two or more different types of atoms bonded in a specific way.
- Compounds consist of the same particles throughout.



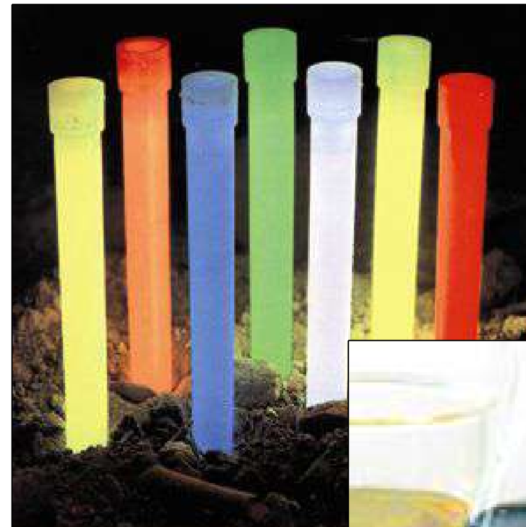


Mixtures

- A mixture is a collection of two or more pure substances.
- Mixtures have variable composition.
- Examples: Air, Water, Kool-Aid, Soda, Blood.
- Mixtures of metals are called <u>alloys</u>.
- There are two types of mixtures.
  - Homogeneous
  - Heterogeneous

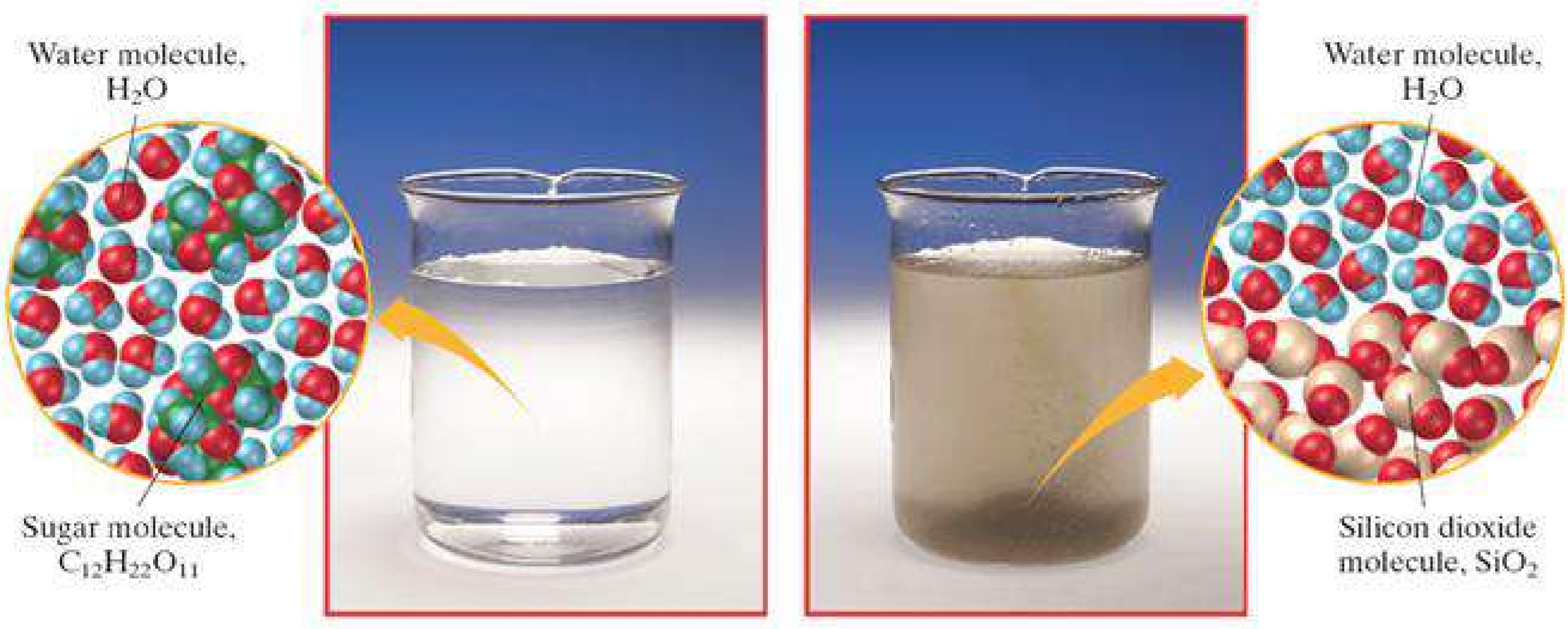
Homogeneous Mixtures

- Does not vary in composition from one region to another.
- Also known as a <u>solution</u>.
- Ex: Salt Water, Brass, Kool-Aid, Soda, Air



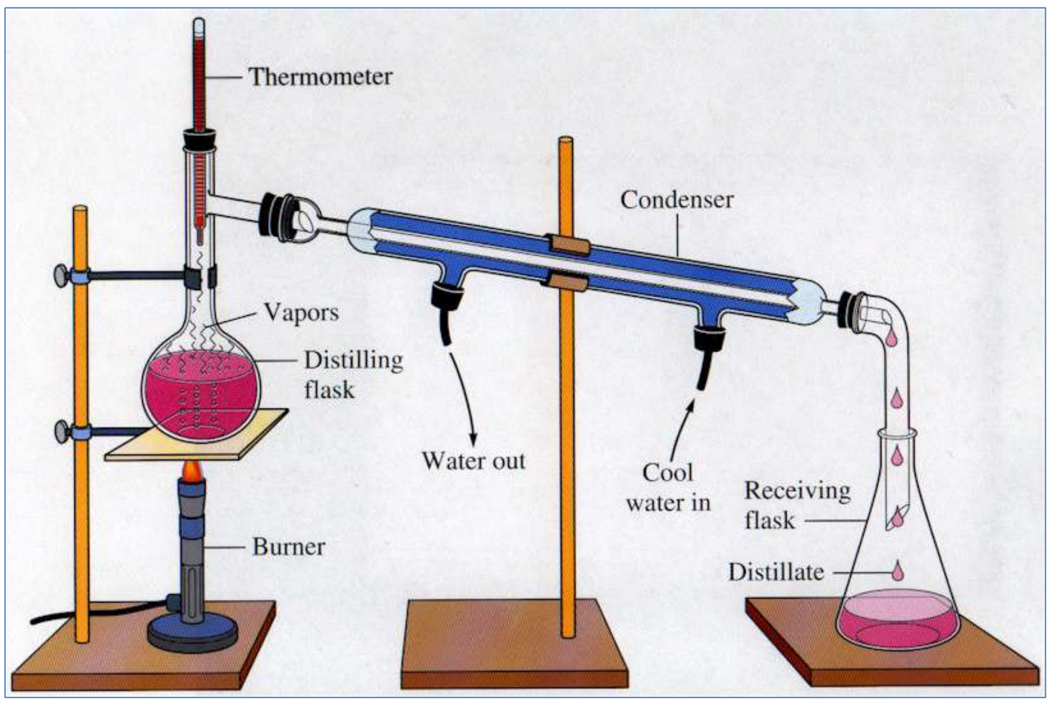
Heterogeneous Mixture

- A mixture of two or more pure substances that contains regions that are different than other regions.



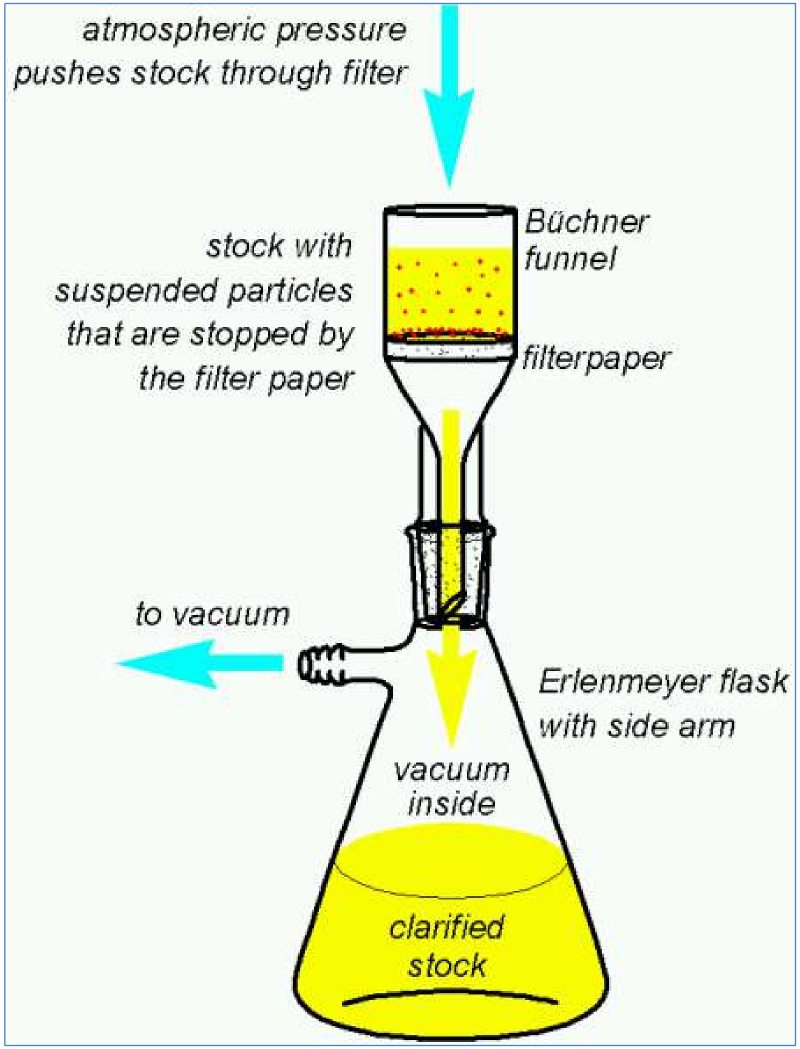
Separation of Mixtures

- Distillation: used to separate mixtures based on differences in volatility (boiling point).



Separation of Mixtures

- Filtration: separates mixtures of solids and liquids.



Separation of Mixtures

- Chromatography: Uses two phases (mobile and stationary) to separate mixtures based on the affinity of the components for the two phases.

Phases or States of Matter

- Solid- Has definite volume and shape.
- Liquid- Has definite volume but no definite shape.
- Gas- Has neither definite volume nor shape.

Melting Evaporating

Solid Liquid Gas

SolidificationCondensation

Physical Change

- Changes in a substance that do <u>not</u> change the composition or identity of the substance.
- Examples:
  - Changes of phase. (Solid, Liquid, Gas)
  - You accidentally break a glass into many pieces.
  - You step on a piece of chalk and it becomes powder.

Chemical Changes

- Also called Chemical Reactions
- Changes in a substance that <u>do</u> change the composition or identity of the substance.
- Examples:
  - Cooking food.
  - Burning anything.
  - Breathing, digesting, thinking, learning.
  - Everything that is not a physical change.

Chapter Questions/Homework

 # 28, 34, 36 ef, 52 ^{O}C only, 53, 67, 71