General Chemistry I
CHEM 100

Chemistry Department

University of Ha'il

Chapter 1

Fundamentals of Chemistry and Measurement

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FIRST SEMESTER						
CODE	COURSE TITLE	*Cr.	Lecture		Lab	
			*Cr.	#Con.	*Cr.	#Con.
CHEM 100	General Chemistry I	4	3	3	1	2

Component	% (Lec)	% (Lab)
Midterm Exam	25	-
Final Exam	40	-
HW + Quizzes	10	-
Lab exams & reports	-	25
Total	75	25

Attendance

المادة التاسعة:

لائحة الدراسة |18

على الطالب المنتظم حضور المحاضرات والدروس العملية، ويحرم من دخول الاختبار النهائي فيها إذا قلت نسبة حضوره عن النسبة التي يحددها مجلس الجامعة، على ألا تقل عن (٧٥%) من المحاضرات والدروس العملية المحددة لكل مقرر خلال الفصل الدراسي، ويعد الطالب الذي حرم من دخول الاختبار بسبب الغياب راسبًا في المقرر، ويرصد له تقدير محروم (ح) أو (DN).

المادة العاشرة:

لائحة الدراسة |19

يجوز لمجلس الكلية أو من يفوضه – استثناء – رفع الحرمان والسماح للطالب بدخول الاختبار، شريطة أن يقدم الطالب عذرًا يقبله المجلس، ويحدد مجلس الجامعة نسبة الحضور على ألا تقل عن (50٪) من المحاضرات والدروس العملية المحددة للمقرر.

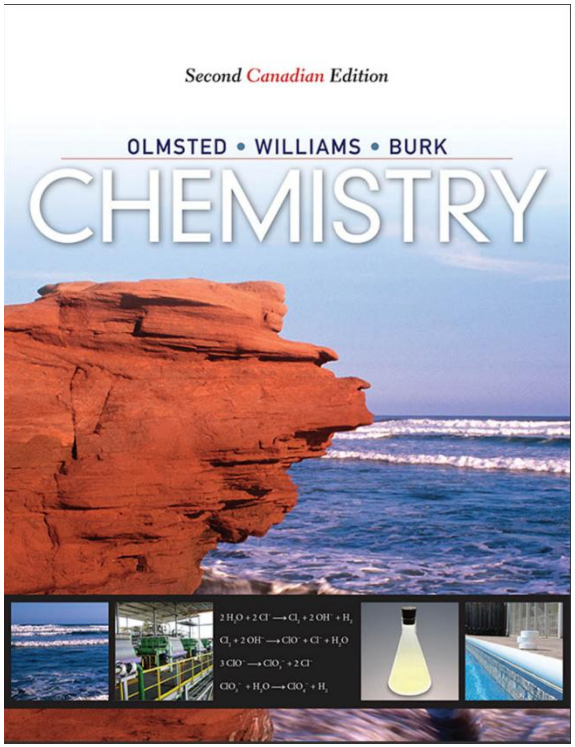

لائحة الدراسة والاختبارات
للمرحلة الجامعية
وقواعدها التنفيذية بجامعة حائل

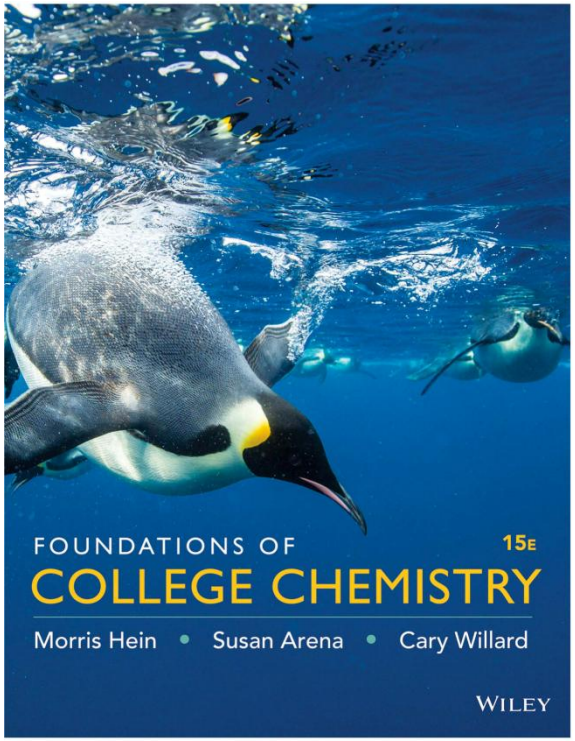
مجلس جامعة حائل
العام 1437 – 1438هـ

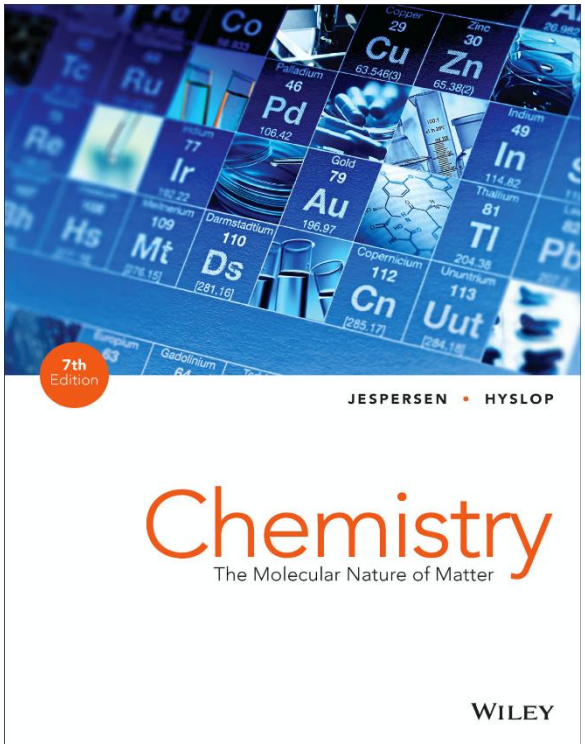
إعداد /
عمادة القبول والتسجيل
الإدارة القانونية

- 7 Abs ≈ 25%
- From 8 unexcused Abs = DN
- 15 Abs = 50% = DN (even with excuses)

Textbooks


Second Canadian Edition
OLMSTED • WILLIAMS • BURK
CHEMISTRY


FOUNDATIONS OF
15E
COLLEGE CHEMISTRY
Morris Hein • Susan Arena • Cary Willard
WILEY


7th Edition
JESPERSEN • HYSLOP
Chemistry
The Molecular Nature of Matter
WILEY

Atomic Theory

- Most significant theoretical model of nature

Atoms

- Tiny submicroscopic particles
- Make up all chemical substances
- Make up everything in macroscopic world
- Smallest particle that has all properties of given element

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Discovery of Subatomic Particles

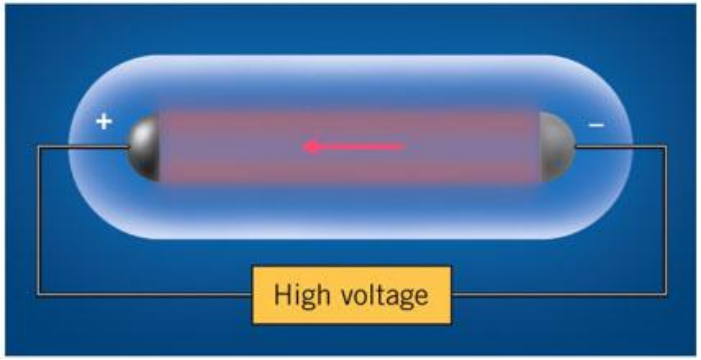
- Late 1800s and early 1900s
- **Key Experiment:**

Cathode ray tube experiments showed that atoms are made up of subatomic particles

- **Major Discovery:**

Discovered negatively charged particles moving from the cathode to the anode

- **Cathode** – negative electrode
- **Anode** – positive electrode

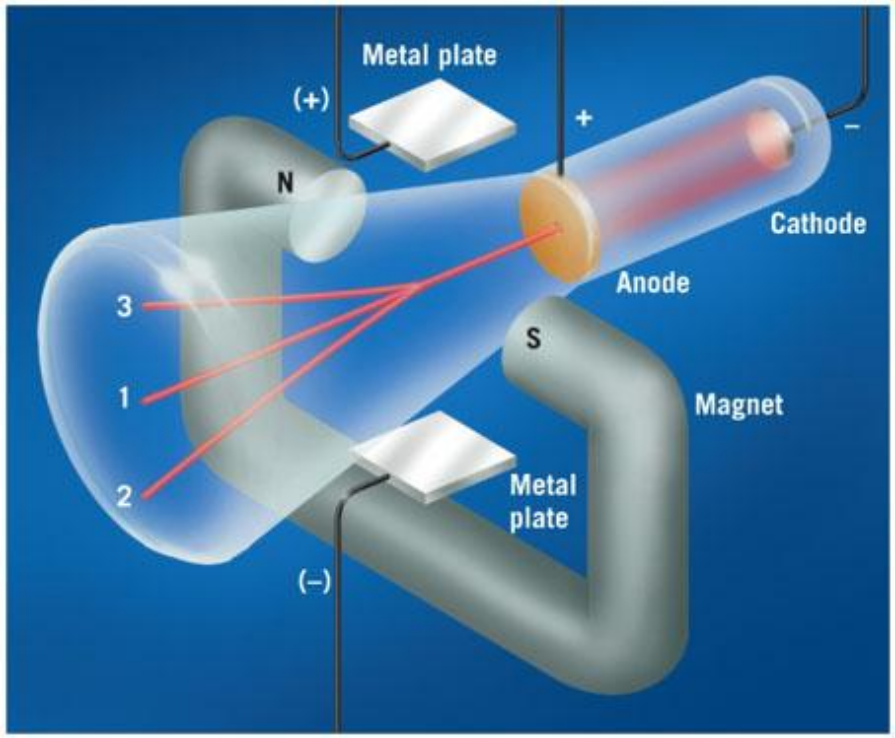


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Discovery of Electron

JJ Thomson (1897)

- Modified cathode ray tube
- Made quantitative measurements on cathode rays
- Discovered negatively charged particles
  - Electrons (e^-)

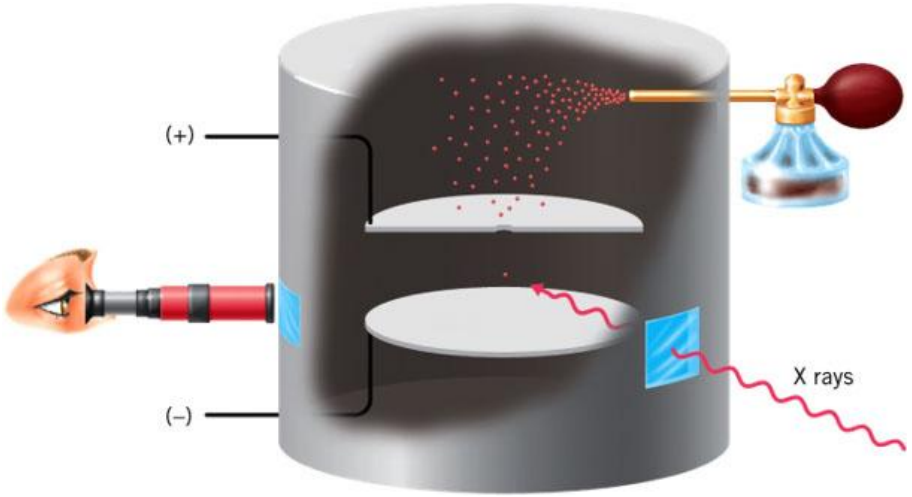


- Determined charge to mass ratio (e/m) of these particles
  - $e/m = -1.76 \times 10^8$ **coulombs/gram**

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Millikan Oil Drop Experiment

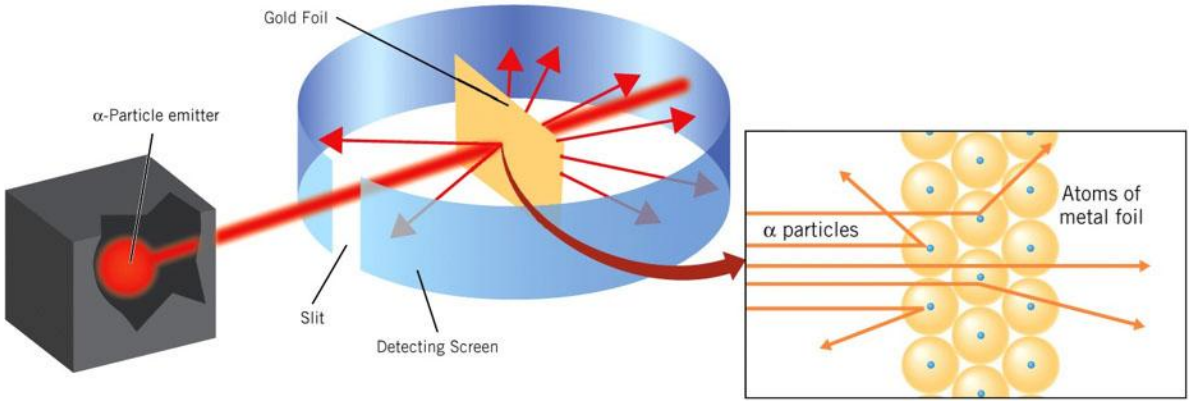
- Determining charge on Electron
- Calculated charge on electron
  - $e^- = -1.60 \times 10^{-19}$ Coulombs
- Combined with Thomson's experiment to get mass of electron
  - $m = 9.09 \times 10^{-28}$ g



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Discovery of Atomic Nucleus

Rutherford's Alpha Scattering Experiment



- Most alpha (α) rays passed right through gold
- A few were deflected off at an angle
- 1 in 8000 bounced back towards alpha ray source
- Gave us current model of **nuclear atom**

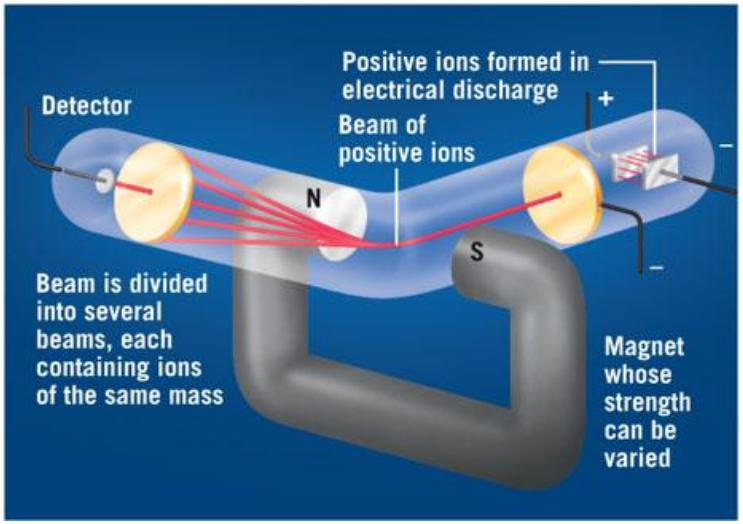
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Discovery of Proton

- Discovered in 1918 in Ernest Rutherford's lab
- Detected using a mass spectrometer
  - Hydrogen had mass 1800 times the electron mass
  - Masses of other gases whole number multiples of mass of hydrogen

Proton

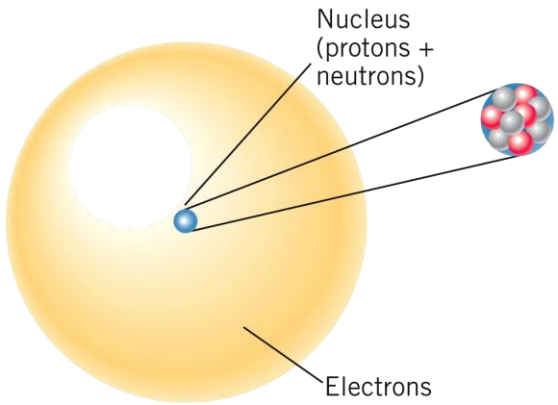
- Smallest positively charged particle



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Rutherford's Nuclear Atom

- Demonstrated that nucleus:
  - has almost all of mass in atom
  - has all of positive charge
  - is located in very small volume at center of atom
- Very tiny, extremely dense core of atom
- Where **protons** $\left({}_{1}^{1}p\right)$ and **neutrons** $\left({}_{1}^{1}n\right)$ are located



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Discovery of Neutron

- First postulated by Rutherford and coworkers
  - Estimated number of positive charges on nucleus based on experimental data
- Nuclear mass based on this number of protons always far short of actual mass
  - About ½ actual mass
- Therefore, must be another type of particle
  - Has mass about same as proton
  - Electrically neutral
- Discovered in 1932 by James Chadwick

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Atomic Structure

- **Electrons** $\left({}_{-1}^{0}e, \text{or } e^{-}\right)$
  - Very low mass
  - Occupy most of atom's space
- Balance of attractive and repulsive forces controls atom size
- **Attraction** between **protons** $\left({}_{1}^{1}p\right)$ and **electrons** $\left({}_{-1}^{0}e\right)$ holds electrons around nucleus
- **Repulsion** between electrons helps them spread out over volume of atom
- In neutral atom
- Number of electrons must equal number of protons
- Diameter of atom ~10,000 × diameter of nucleus

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Properties of Subatomic Particles

Three kinds of subatomic particles of principal interest to chemists

Particle	Mass (g)	Electrical Charge	Symbol
Electron	9.109×10^{-28}	-1	${}^{0}_{-1}e$
Proton	1.673×10^{-24}	$+1$	${}^{1}_{1}\mathrm{H}$, ${}^{1}_{1}p$
Neutron	1.675×10^{-24}	0	${}^{1}_{0}n$

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Atomic Notation (1 of 2)

Atomic number (Z)

- Number of protons that atom has in nucleus
  - Unique to each type of element
  - Element is substance whose atoms all contain identical number of protons
- **Z = number of protons**

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Atomic Notation (2 of 2)

Mass number (A)

- A = (number of protons) + (number of neutrons)
- $A = Z + N$
- For charge neutrality, number of electrons and protons must be equal

Atomic Symbols

- Summarize information about subatomic particles
- Every isotope defined by two numbers Z and A
- Symbolized by ${}^{A}_{Z}X$

Example: What is the atomic symbol for helium?

He has $2e^-$, $2n$ and $2p$ $Z = 2, A = 4$

${}^{4}_{2}\mathrm{He}$

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Isotopes

- Most elements are mixtures of two or more stable isotopes
- Each isotope has slightly different mass
- Chemically, isotopes have virtually identical chemical properties
- Relative proportions of different isotopes are essentially constant
- Isotopes distinguished by **mass number (*A*)**:

example:

  - Three isotopes of hydrogen (**H**)
  - Four isotopes of iron (**Fe**)

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Your Turn!

An atom of $^{206}_{82}Pb$ has ___ protons, ___ neutrons, and ___ electrons.

A. 82, 206, 124

B. 124, 206, 124

C. 124, 124, 124

D. 82, 124, 82

E. 82, 124, 124

Answer: D

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Learning Check:

- Fill in the blanks:

symbol	neutrons	protons	electrons
$^{131}_{53}I$	_____	_____	_____
$^{81}_{35}Br$	_____	_____	_____
_____	36	29	29

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Learning Check:

- Fill in the blanks:

symbol	neutrons	protons	electrons
$^{131}_{53}I$	78	53	53
$^{81}_{35}Br$	46	35	35
$^{65}_{29}Cu$	36	29	29

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Atomic Mass Scale

- Need uniform mass scale for atoms

Atomic mass units (symbol **u**)

- Based on carbon:
- 1 atom of carbon-12 = 12 u (exactly)
- $1\ \text{u} = \frac{1}{12}$ mass 1 atom of carbon-12 (exactly)

Why was ^{12}C selected?

- Common
- Most abundant isotope of carbon
- All atomic masses of all other elements ~ whole numbers
- Lightest element, H, has mass ~1 u

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Calculating Atomic Mass

- Generally, elements are mixtures of isotopes

Example: Hydrogen

Isotope	Mass	% Abundance
^{1}H	1.007825 u	99.985
^{2}H	2.0140 u	0.015

How do we define atomic mass?

- Average of masses of all stable isotopes of given element

How do we calculate average atomic mass?

- Weighted average
- Use isotopic abundances and isotopic masses

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Your Turn!

Answer: B

A naturally occurring element consists of **two** isotopes. The data on the isotopes:

isotope #1	68.5257 u	60.226%
isotope #2	70.9429 u	39.774%

Calculate the average atomic mass of this element.

A. 70.943 u
B. 69.487 u
C. 69.526 u
D. 69.981 u
E. 69.734 u

$0.60226 \times 68.5257\ \text{u} = 41.270\ \text{u}$

$0.39774 \times 70.9429\ \text{u} = 28.217\ \text{u}$

$69.487\ \text{u}$

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Matter and Its Classifications

Matter

- Anything that has mass and occupies space

Mass

- How much matter given object has
- Measure of object's momentum, or resistance to change in motion

Weight

- Force with which object is attracted by gravity

Example: Mass versus Weight

Astronaut on moon and on earth

- Weight on moon = 1/6 weight on earth
- Same mass regardless of location

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Matter and Elements

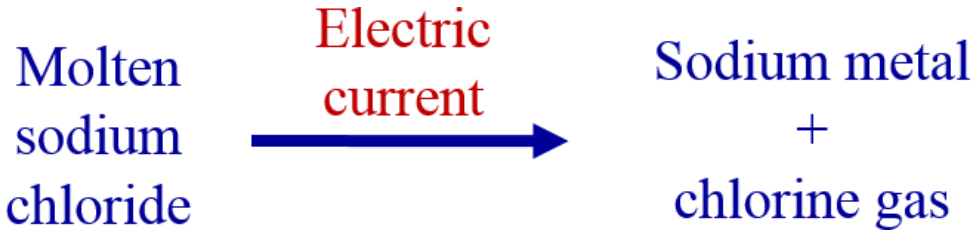
Chemical Reactions

- Transformations that alter chemical compositions of substances

Decomposition

- A subset of chemical reactions
- Chemical reaction where one substance is broken down into two or more simpler substances

Example:



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Elements

- Substances that can't be decomposed into simpler materials by chemical reactions
- Substances composed of only one type of atom
- Simplest forms of matter that we can work with directly
- More complex substances composed of elements in various combinations


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Chemical Symbols for Elements

Chemical Symbol

- One or two letter symbol for each element name
- First letter capitalized, second letter lower case

Example:

C = carbon	S = sulfur
Ca = calcium	Ar = argon
Br = bromine	H = hydrogen
Cl = chlorine	O = oxygen

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Chemical Symbols for Elements

- Used to represent elements in chemical formulas

Example: Water = H_2O

Carbon dioxide = CO_2

- Most based on English name
- Some based on Latin or German names

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Chemical Symbols

English Name	Chemical Symbol	Latin Name
Sodium	Na	Natrium
Potassium	K	Kalium
Iron	Fe	Ferrum
Copper	Cu	Cuprum
Silver	Ag	Argentum
Gold	Au	Aurum
Mercury	Hg	Hydrargyrum
Antimony	Sb	Stibium
Tin	Sn	Stannum
Lead	Pb	Plumbum
Tungsten	W	Wolfram (German)

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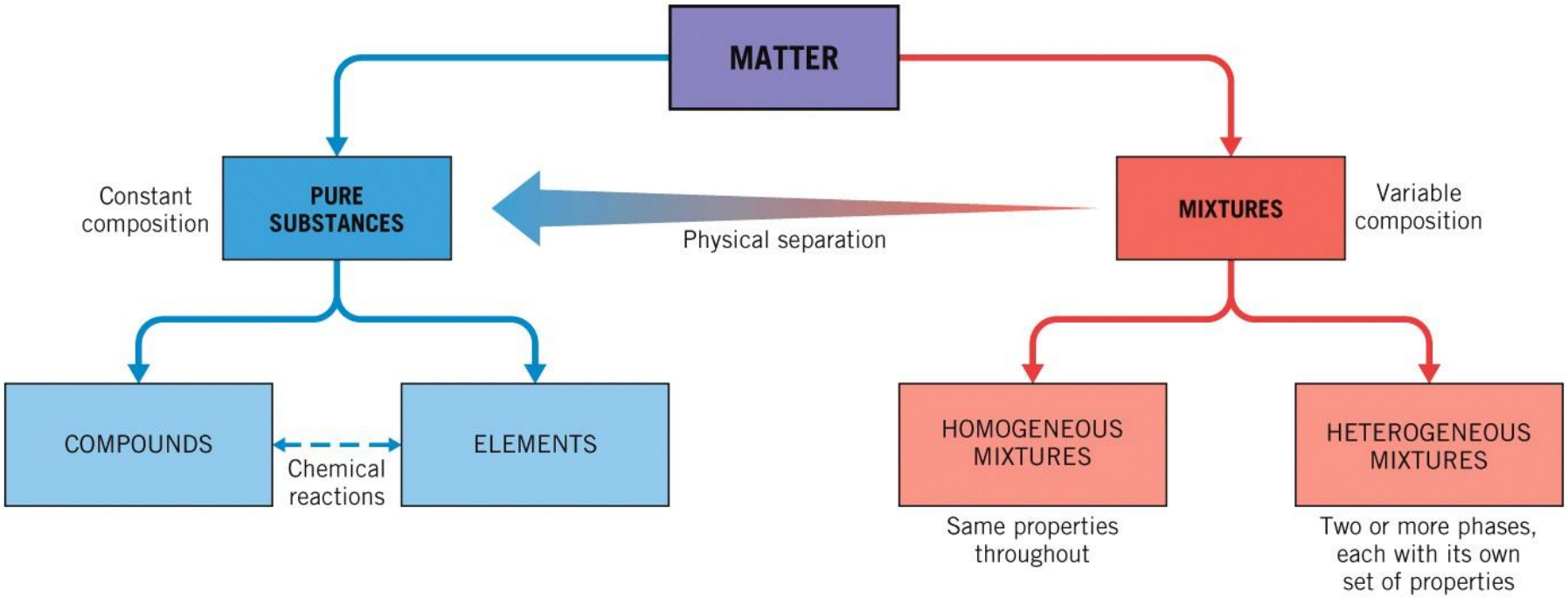
Compound

- Formed from two or more atoms of different elements
- Always combined in same fixed ratios by mass
- Can be broken down into elements by some chemical changes

Example: Water decomposed to elemental hydrogen and oxygen

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Relationship of Elements, Compounds, and Substances



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Learning Check: Classification

	Chicken Noodle Soup	Ice (H_2O)	Liquid Dish Soap	Table Salt (NaCl)
Pure Substance		X		X
Element				
Compound		X		X
Heterogeneous Mixture	X			
Homogeneous Mixture			X	

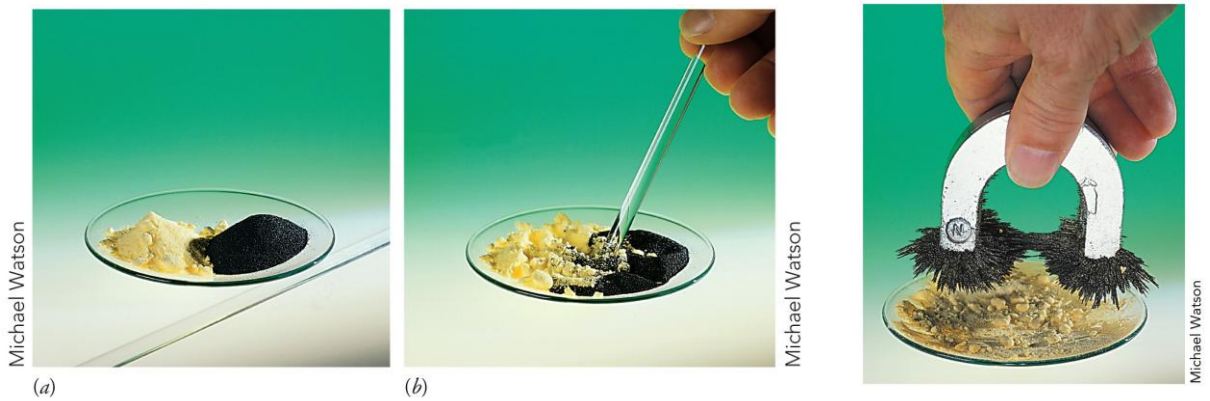
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Physical Change

- No new substances formed
- Substance may change state or the proportions

Example: Ice melting

- Sugar or salt dissolving
- Stirring iron filings and sulfur together



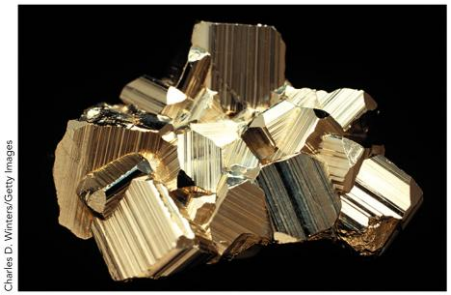
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Chemical Change

- Formation of new substance or compound
- Involves changing chemical makeup of substances
- New substance has different physical properties
- Can′t be separated by physical means

Example:

- Fool′s gold
- Compound containing sulfur and iron
  - No longer has same physical properties of free elements
  - Can′t be separated using magnet



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Learning Check 1:

- For each of the following, determine if it represents a chemical or physical change:

	Chemical	Physical
Magnesium burns when heated	X	
Magnesium metal tarnishes in air	X	
Magnesium metal melts at 922 K		X
Grape Kool-Aid lightens when water is added		X

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Physical Properties

Properties that can be observed without changing the chemical makeup of a substance

Example:

- Color
- Electrical conductivity
- Melting point and boiling point

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Chemical Properties

- Describes how a substance undergoes a chemical reaction – its 'reactivity'
  - Involves changing chemical makeup of substances
  - New substance has different physical properties
  - Can't be separated by physical means

Example:

- Fool's gold
- Compound containing sulfur and iron
  - No longer has same physical properties of free elements
  - Can't be separated using magnet

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Intensive versus Extensive Properties

Intensive properties

- Independent of sample size
- Used to identify substances

Example: Color
Density
Boiling point
Melting point
Chemical reactivity

Extensive properties

- Depend on sample size

Example: volume and mass

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Observations

- Fall into two categories

1. **Quantitative** observations
   - Numeric data
   - Measure with instrument

   Example: Melting point, boiling point, volume, mass

2. **Qualitative** observations
   - Do not involve numerical information

   Example: Color, rapid boiling, white solid forms

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International System of Units (SI)

- Standard system of units used in scientific and engineering measurements
- Metric
- Seven Base Units

TABLE 1.2 The SI Base Units

Measurement	Unit	Abbreviation
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electric current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

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Laboratory Measurements

1. Length

- SI Unit is meter (m)

2. Volume

- SI unit for Volume = m^3

3. Mass

- SI unit is kilogram (kg)

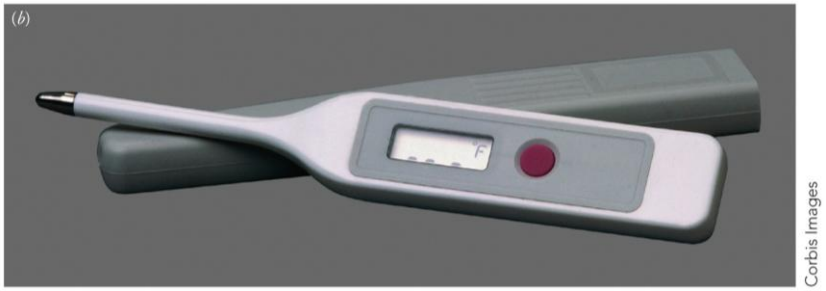
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4. Temperature

- Measured with thermometer
- Three common scales

A. Fahrenheit scale

- Common in US
- Water freezes at 32 °F and boils at 212 °F
- 180 degree units between melting and boiling points of water

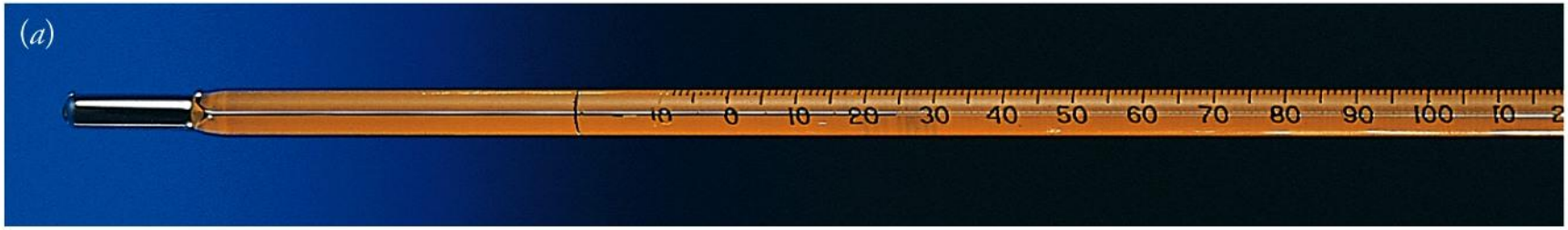


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4. Temperature

B. Celsius scale

- Most common for use in science
- Water freezes at 0 °C
- Water boils at 100 °C
- 100 degree units between melting and boiling points of water



Michael Watson

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4. Temperature

C. Kelvin scale

- SI unit of temperature is **kelvin (K)**
  - **Note:** No degree symbol in front of K
- Water freezes at 273.15 K and boils at 373.15 K
  - 100 degree units between melting and boiling points
- Only difference between Kelvin and Celsius scale is zero point

Absolute Zero

- Zero point on Kelvin scale
- Corresponds to nature's lowest possible temperature

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Temperature Conversions

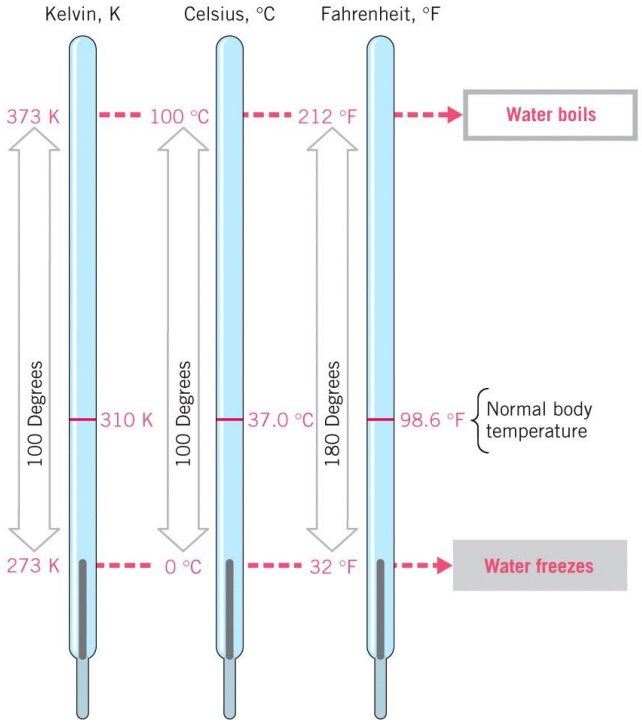
How to convert between °F and °C?

$$°F = 1.8\,(°C) + 32$$

Example: 100 °C = ? °F

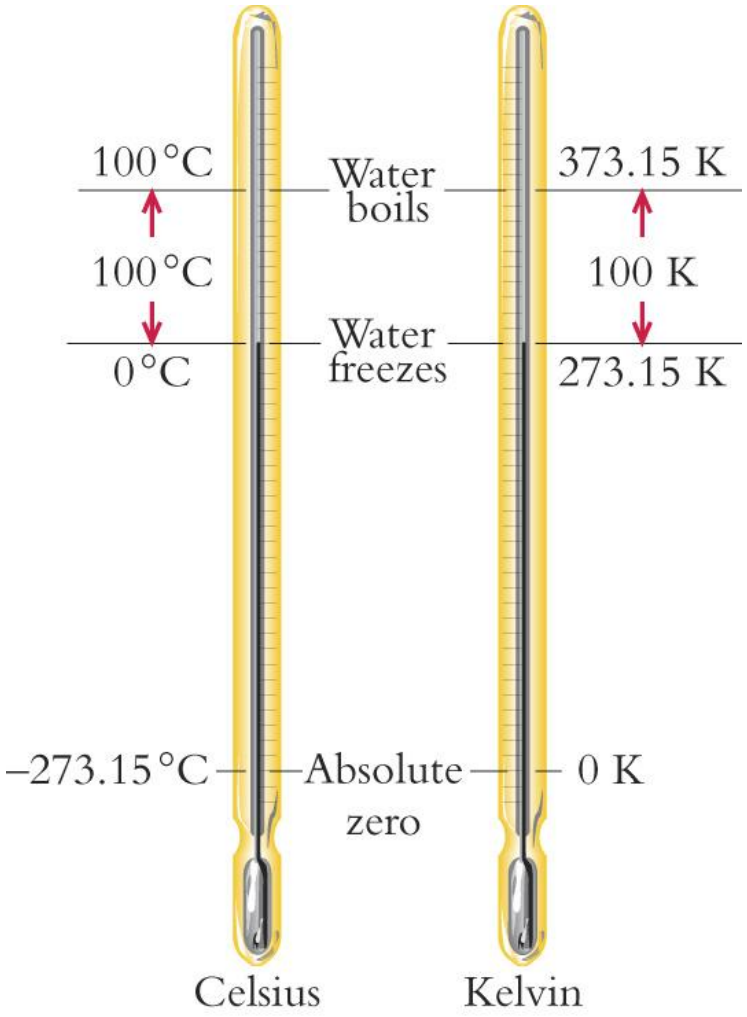
$$°F = 1.8\,(100) + 32$$

$$t_F = 212\ °F$$



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Temperature Conversions



$$T(K) = T(°C) + 273.15$$

$$T\,(°C) = T\,(K) - 273.15$$

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Learning Check: T Conversions

1. Convert 121 °F to the Celsius scale.

$$°F = 1.8\,(°C) + 32 \qquad °C = \frac{(°F - 32)}{1.8}$$

$$°C = \frac{(121-32)}{1.8} = \mathbf{49\ °C}$$

2. Convert 121 °F to the Kelvin scale.

- We already have in °C so…

$$T_K = (t_C + 273.15\ °C)\frac{1\ K}{1\ °C} = (49 + 273.15\ °C)\frac{1\ K}{1\ °C} \qquad \mathbf{T_K = 332\ K}$$

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Learning Check: T Conversions

3. Convert 77 K to the Celsius scale.

$$T_K = (t_C + 273.15\ °C)\frac{1\ K}{1\ °C} \qquad t_C = (T_K - 273.15\ K)\frac{1\ °C}{1\ K}$$

$$t_C = (77\ K - 273.15\ K)\frac{1\ °C}{1\ K} \quad = -196\ °C$$

4. Convert 77 K to the Fahrenheit scale.

- We already have in °C so

$$°F = 1.8\,(-196) + 32 \quad = -321\ °F$$

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Significant Figures

- Scientific convention:
  - **All digits in measurement up to and including first estimated digit are significant.**
- Number of certain digits plus first uncertain digit
- Digits in measurement from first non-zero number on left to first estimated digit on right

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Rules for Significant Figures

1. All non-zero numbers are significant.

Example: 3.456 has **4** sig. figs.

3. Zeros between non-zero numbers are significant.

Example: $20,089$ or 2.0089×10^4 has **5** sig. figs

3. Trailing zeros always count as significant if number has decimal point

Example: 5.00×10^2 has **3** sig. figs

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Rules for Significant Figures

4. Final zeros on number without decimal point are NOT significant

Example: $104,956,000$ or 1.04956×10^{8} has **6** sig. figs

5. Final zeros to right of decimal point are significant

Example: 3.00 has **3** sig. figs.

6. Leading zeros, to left of first nonzero digit, are never counted as significant

Example: 0.00012 or 1.2×10^{-4} has **2** sig. figs

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Learning Check

How many significant figures does each of the following numbers have?

	Scientific Notation	# of Sig. Figs.
1. 413.97	4.1397×10^{2}	**5**
2. 0.0006	6×10^{-4}	**1**
3. 5.120063	5.120063	**7**
4. 161,000	1.61×10^{5}	**3**
5. 3600.	3.600×10^{3}	**4**

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Scientific Notation

- Clearest way to present number of significant figures unambiguously
  - Report number between 1 and 10 followed by correct power of 10
  - Indicates only significant digits

Example: 75,000 people attend a concert

- If a rough estimate
  - Uncertainty ± 1000 people
  - 7.50×10^4
- If number estimated from aerial photograph
  - Uncertainty ±100 people
  - 7.50×10^4

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Learning Check:

Round each of the following to three significant figures. Use scientific notation where needed.

1. 37.459 — **37.5 or 3.75×10^{1}**
2. 5431978 — **5.43×10^{6}**
3. 132.7789003 — **133 or 1.33×10^{2}**
4. 0.00087564 — **8.77×10^{-4}**
5. 7.665 — **7.66**

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Density

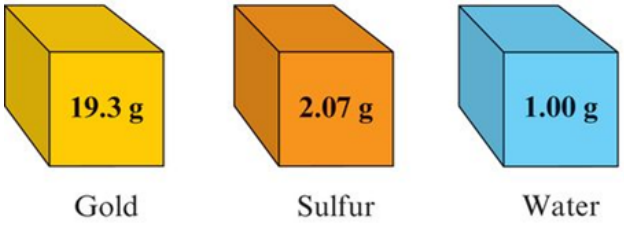
- Ratio of object's mass to its volume

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

- Intensive property (size independent)
  - Determined by taking ratio of two extensive properties (size dependent)
  - Frequently ratio of two size dependent properties leads to size independent property
  - Sample size cancels
- Units
  - **g/ml** or **g/cm^3**

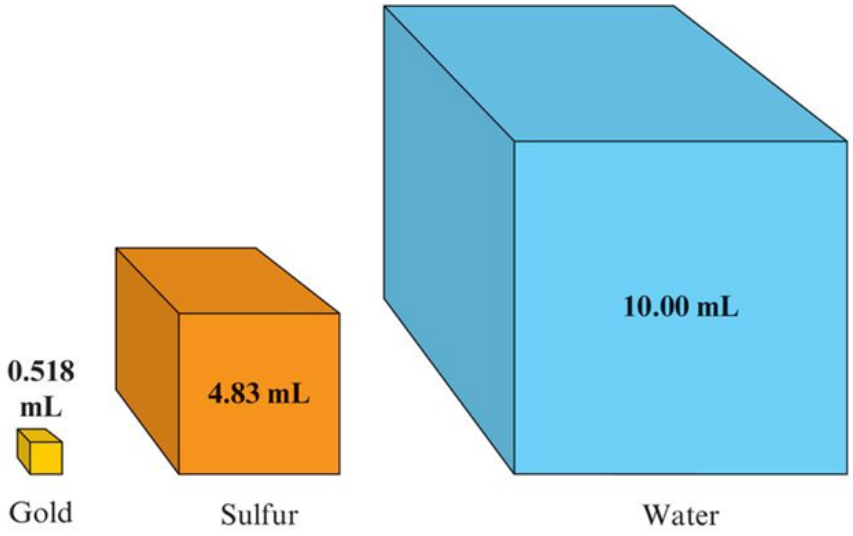
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Density



Volume, 1.00 cm^3

Comparing masses of equal volumes



Mass, 10.0 g

Comparing volumes of equal masses

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Learning Check:

1. Glass has a density of $2.2\,\text{g/cm}^3$. What is the volume occupied by 35 g of glass?

$$V = \frac{m}{d} = \frac{35\,\text{g}}{2.2\,\text{g}/\text{cm}^3} = \mathbf{16\ cm^3}$$

2. What is the mass of $4.00\,\text{cm}^3$ of glass?

$$m = d \times V = 2.2\,\text{g}/\text{cm}^3 \times 400.\text{cm}^3 = \mathbf{8.8 \times 10^2\ g}$$

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Density

- Most substances expand slightly when heated
  - Same mass
  - Larger volume
  - Less dense
- Density decreases slightly as temperature increases
- Liquids and solids
  - Change is very small
  - Can ignore except in very precise calculations
- Density useful to transfer between mass and volume of substance

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Thanks