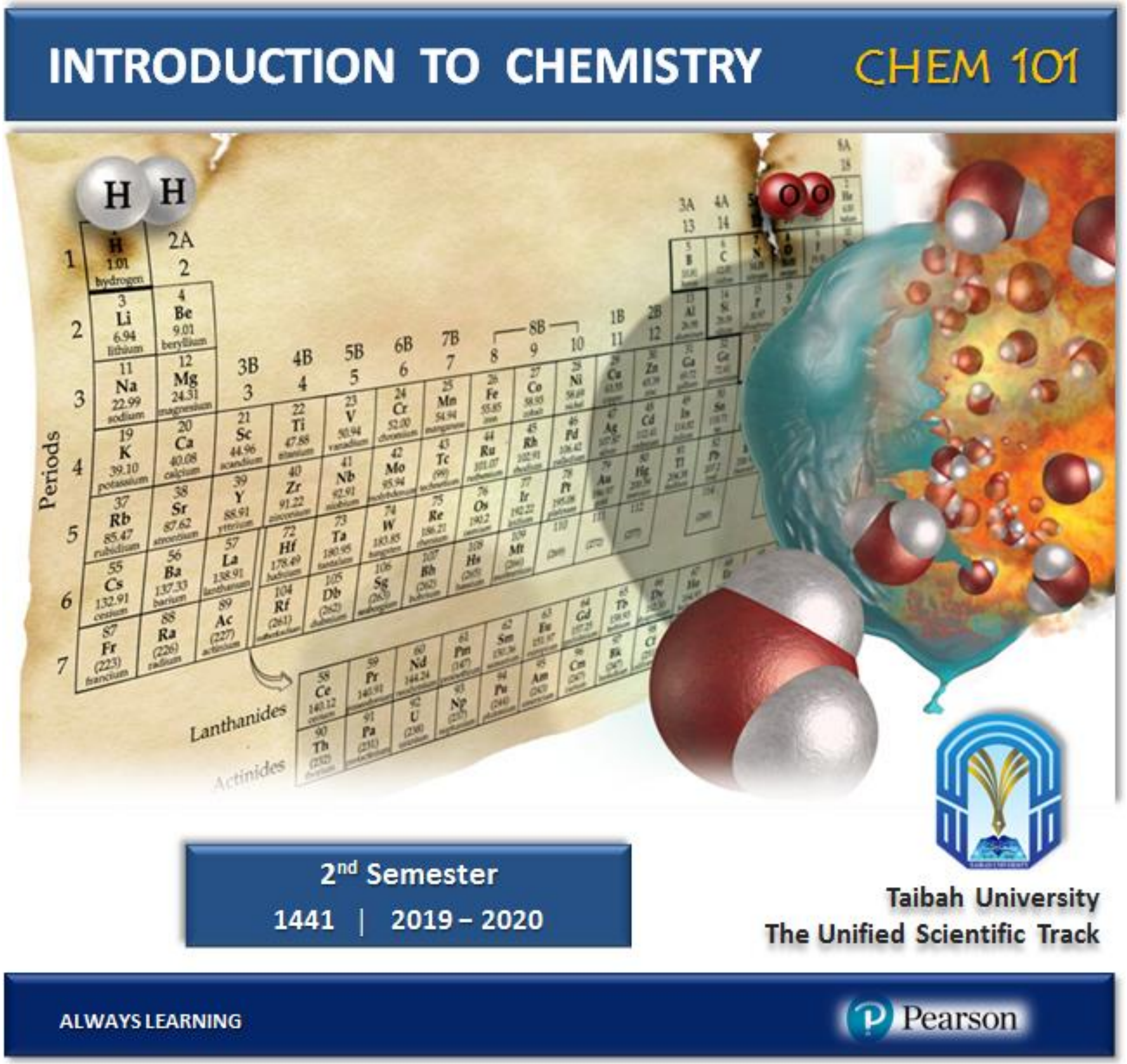


Lecture Presentation

Chapter 1

Matter and Measurements

Topic 01

Classification and States of Matter

The Chemical World

Chemistry:

The science that seeks understanding the properties and behavior of matter by studying atoms and molecules.

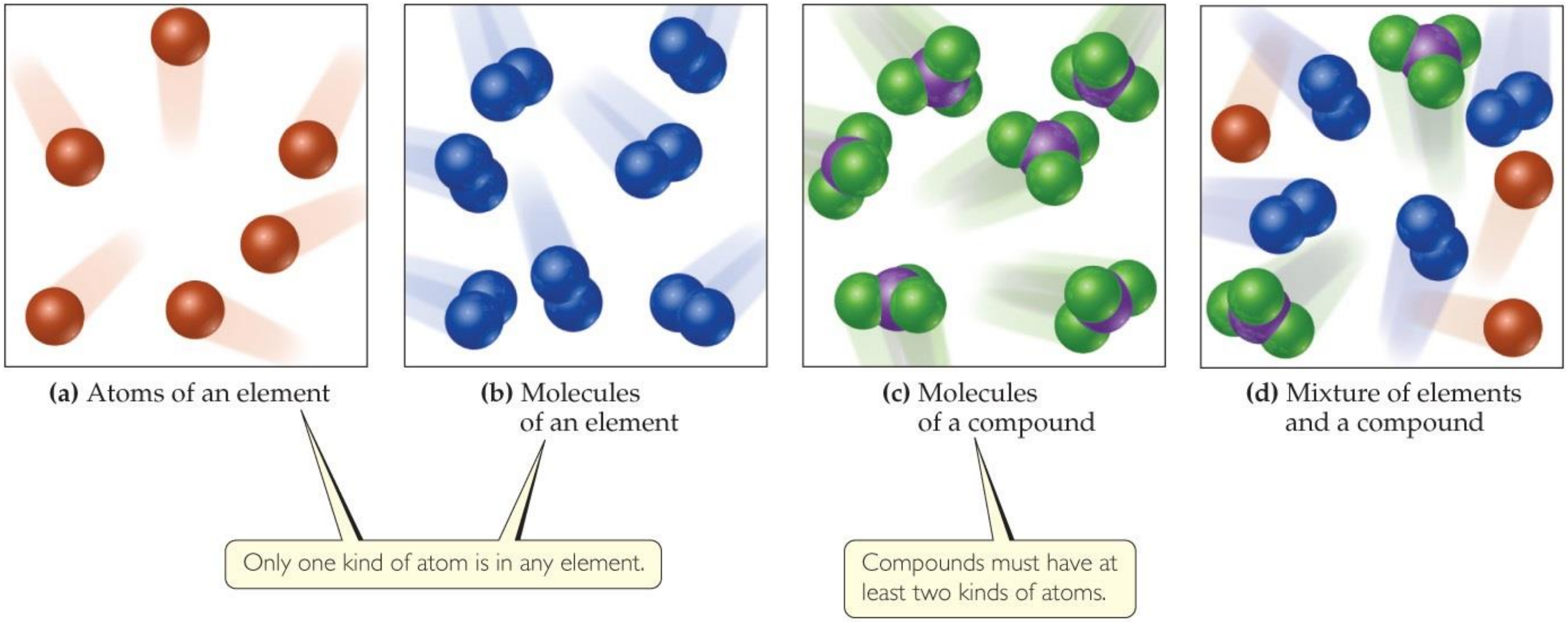
- Chemistry is central to understand many other scientific fields.

- Virtually, everything around you is composed of “chemicals”.

Chemistry: The Central Science

1.1 Atoms and Molecules

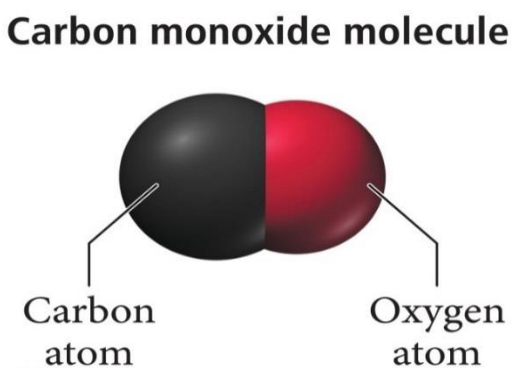
- **Atoms** are the building-blocks of **matter**.
- Each **element** is made of a unique kind of **atoms** (so far, 120 elements are identified in the universe, they are represented in the periodic table of elements).
- The **compound** is made of two or more atoms of different elements, bonded together to form **molecules** (molecules are the building-blocks of compounds).
- The properties of a substance are determined by the properties of its constituent molecules and atoms.



Important Note: some elements are present as "**molecules**" instead of "**free atoms**", they are called:

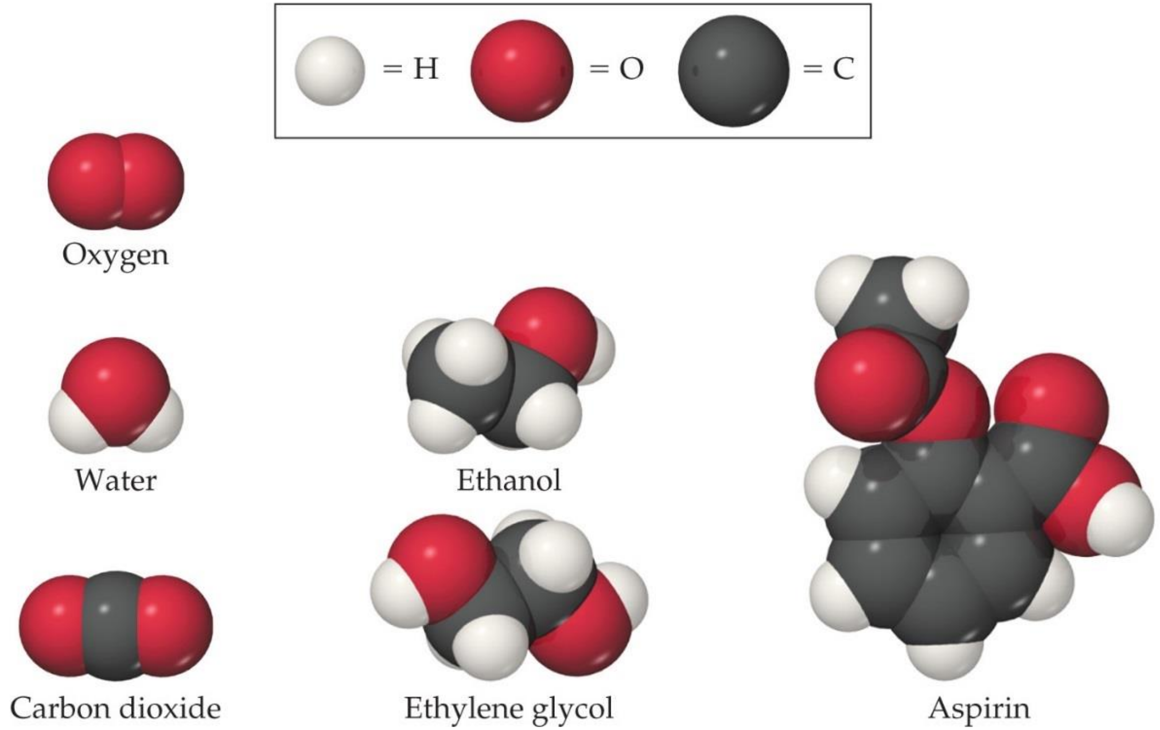
"**Molecular Elements**", such as: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2

1.1 Atoms and Molecules: Example 1



- ✓ The air contains **carbon monoxide** pollutant.
- ✓ Each molecule contains a carbon **atom** and an oxygen **atom** held together by a **chemical bond**.

1.1 Atoms and Molecules: Example 2



Note: Balls of different colors are used to represent **atoms** of different **elements**. Attached balls represent connections between atoms that are seen in nature. These groups of atoms are called **molecules**.

1.2 The Classifications of Matter

- **Matter** is anything that **occupies space** and has **mass**.

  Examples: your textbook, your desk, the air around you, and even your body, are all composed of matter. Matter is everything around us.

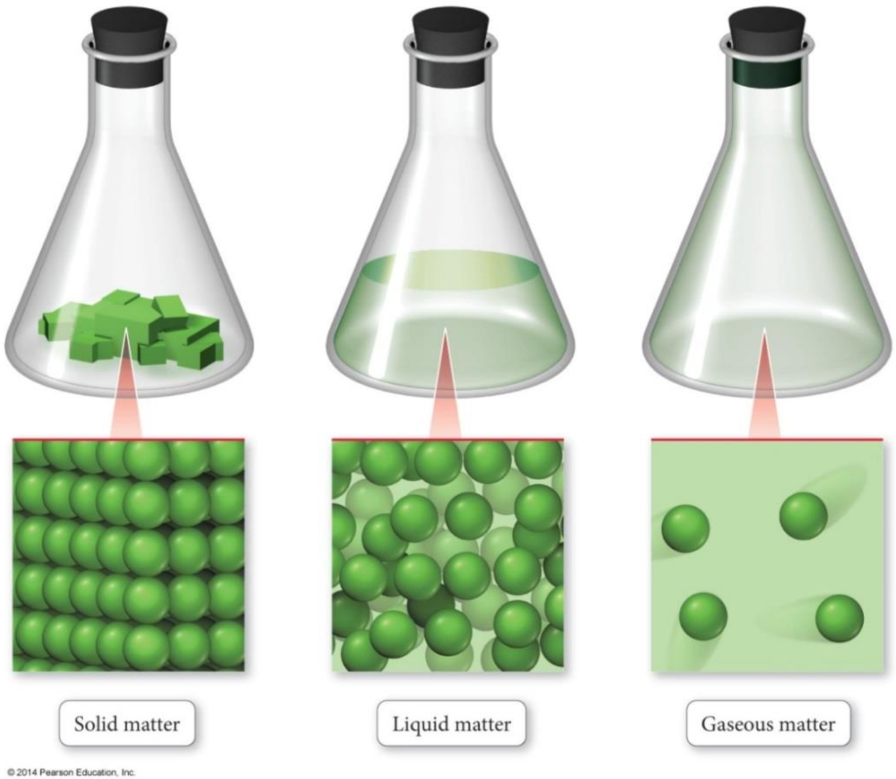
- **Matter can be classified according to:**

  1. **State** (the physical form)

  2. **Composition** (the components that make it up)

The States of Matter

- **Matter** can exist in one of three main states: **solid**, **liquid**, or **gas**.



The state of matter changes from solid to liquid to gas by **increasing temperature**, and vice versa!

Solid Matter

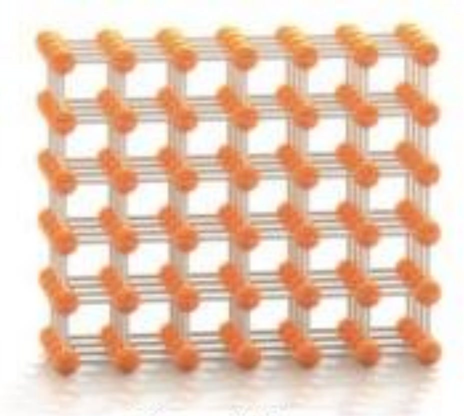
- **Solid Matter**: is composed of tightly packed particles (atoms or molecules). Solids retain their shapes because the particles are not free to move.

- Although the atoms and molecules vibrate in solids, they do not move around or past each other.

- Consequently, solid matter has a **fixed (definite) volume** and a **fixed (rigid) shape**.
  - Examples of solids: Ice, aluminum, iron, wood, salt, and diamond.

Solid Matter: Crystalline or Amorphous?

- **Crystalline Solids:** atoms or molecules are arranged in "patterns" with a long-range repeating order.
  - **Important Examples on crystalline solids:**
    - table salt (NaCl) and diamond.



Crystalline

- **Amorphous Solids:** atoms or molecules are not arranged in long-range patterns.
  - **Important Examples on amorphous solids:**
    - graphite, rubber, glass and plastic.

Amorphous

Liquid Matter

- **Liquid Matter:** is made of more loosely packed particles than in solids. Particles can move about within a liquid, but they are packed densely enough that volume is maintained.

- The ability of liquids to flow, makes them take the shapes of their containers.

- Liquids have **fixed volume** but **no fixed shape**.
  - Examples of liquids: water, oil, and gasoline.

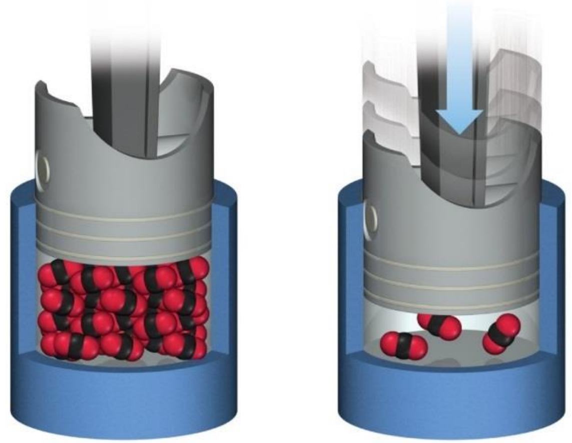
Gaseous Matter

- **Gaseous Matter**: is composed of particles packed so loosely that it has neither a defined shape nor a defined volume.
- Particles of gases (atoms or molecules) are free to move relative to one another.
- Gases have **no fixed volume** and **no fixed shape**, they take the volume and shape of their containers.

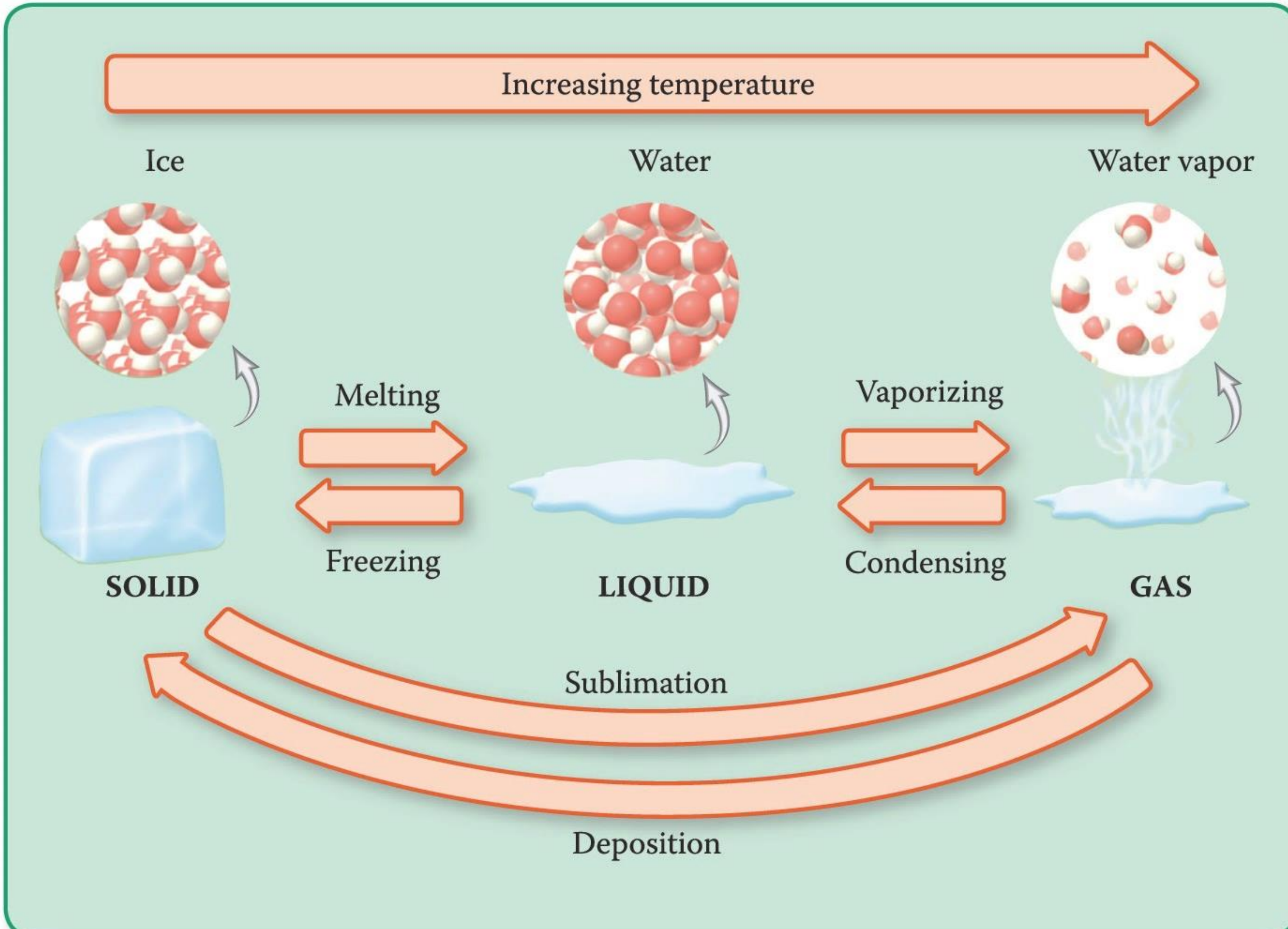
These qualities make gases **compressible**.

- Examples of gases:

  oxygen, nitrogen, CO_2, water vapor



Summary of State Changes of Matter



Classification of Matter According to its Composition

- **Matter can be divided into two classes:**

1. Mixtures: are composed of more than one substance and can be physically separated into its component substances.

2. Pure substances: are composed of only one substance and can NOT be physically separated.

Mixtures

There are **two types of mixtures**:

1. **Heterogeneous mixtures**
2. **Homogeneous mixtures**

- ✓ **Heterogeneous Mixture:** does NOT have uniform properties throughout.
  - (sand + water), (oil + water) or (gasoline + water) are examples on heterogeneous mixtures.
- ✓ **Homogeneous Mixture:** has uniform properties throughout.
  - (salt water), (sugar + water) and alloys are homogeneous mixtures.

Pure Substances

There are **two types of pure substances:**

1. **Compounds**
2. **Elements**

✓ **Compound:** can be chemically separated into individual elements.

There are millions of compounds in the universe.

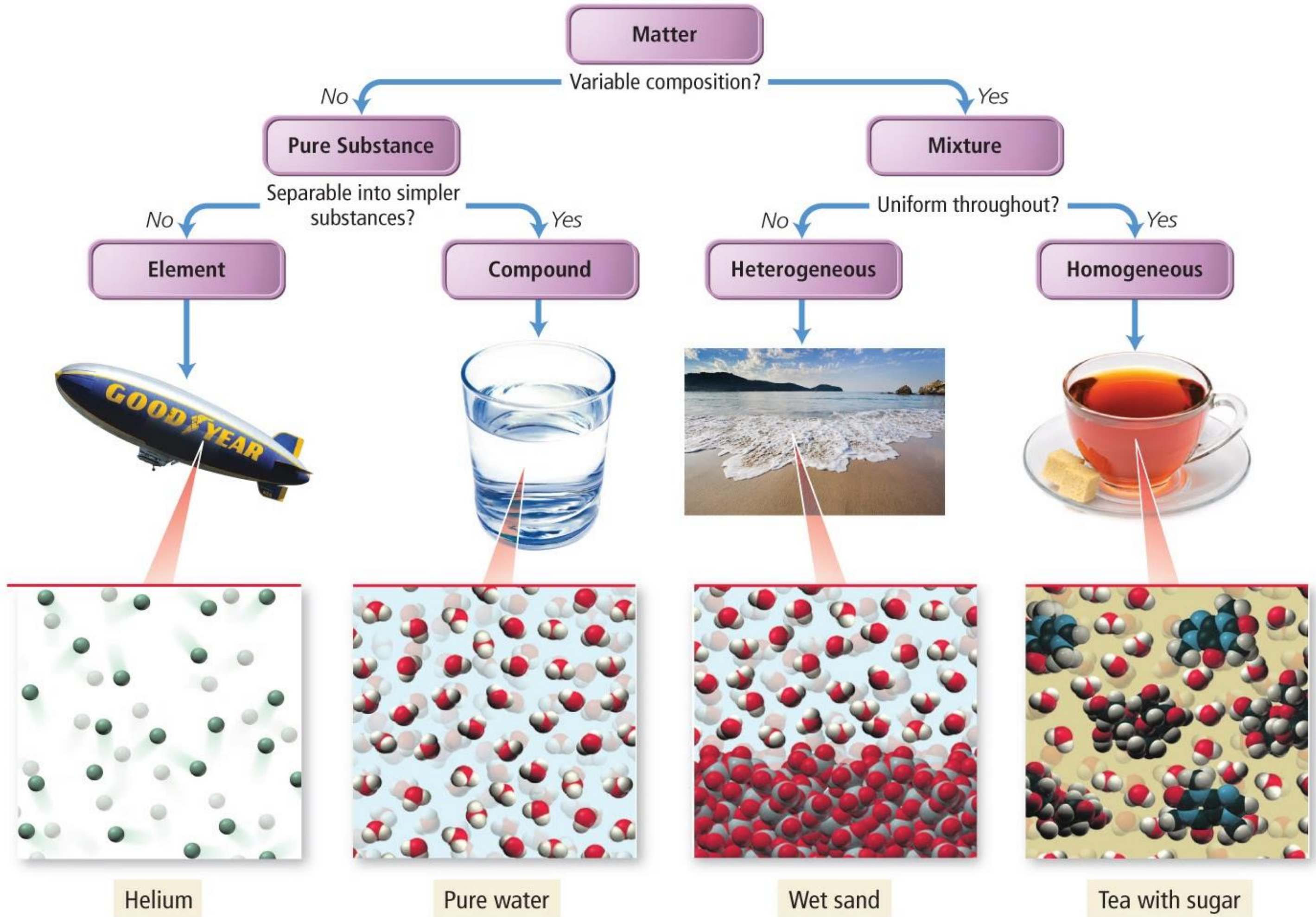
➢ Water is a compound that can be separated into hydrogen and oxygen.

✓ **Element:** cannot be broken down further by chemical reactions.

➢ Elements are the 120 members of the periodic table of elements, such as: Sodium, Iron, Gold, Silver, Hydrogen, Oxygen, Carbon …... etc

Summary of Types of Matter

Matter can be classified according to its composition into: **pure substances** *(elements or compounds)* and **mixtures** *(homogeneous or heterogeneous):*



Assessment

1- The process is which a solid substance is transformed directly into a gas is called _________ and it requires ___________ of temperature.

2- _________ is the physical process which changes a gas into a liquid, and it needs _______ of temperature.

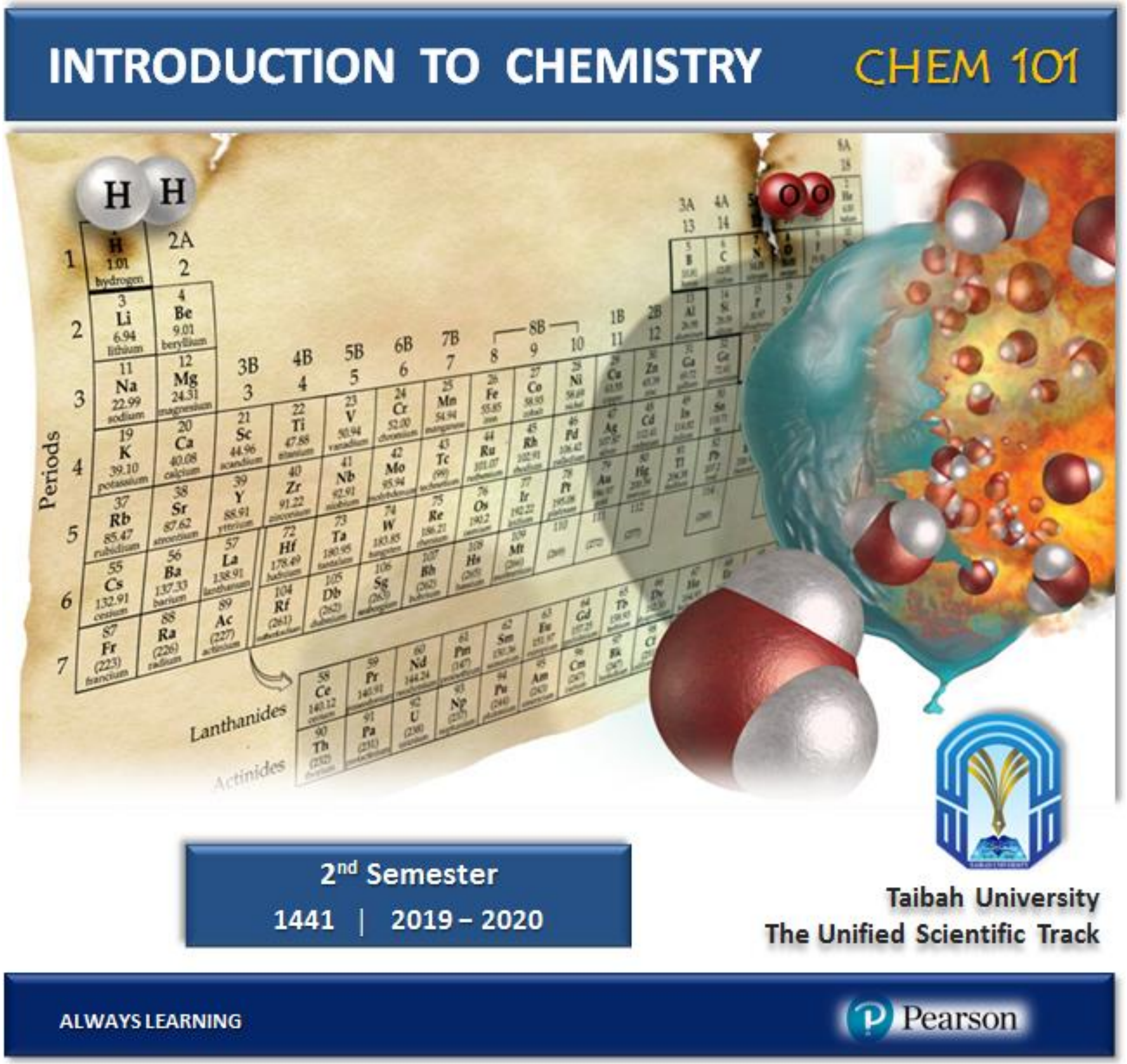
3- Which state of matter has a fixed volume but not a fixed shape.

4- A_______ matter is able to assume both the shape and volume of its container.

5- The ability of both ______ and ______ states of matter to flow makes them able to change their shape to the shape of their reservoir.

6- Classify each substance as a pure substance or a mixture, and indicate the type of each of them (element, compound or homogeneous, heterogeneous):

a. sweat
b. carbon dioxide
c. aluminum
d. salt
e. rust
f. wet sand
g. air
h. oxygen gas
i. bronze alloy
j. honey



Lecture Presentation

Chapter 1

Matter and Measurements

Topic 02

Physical and Chemical Changes & Properties

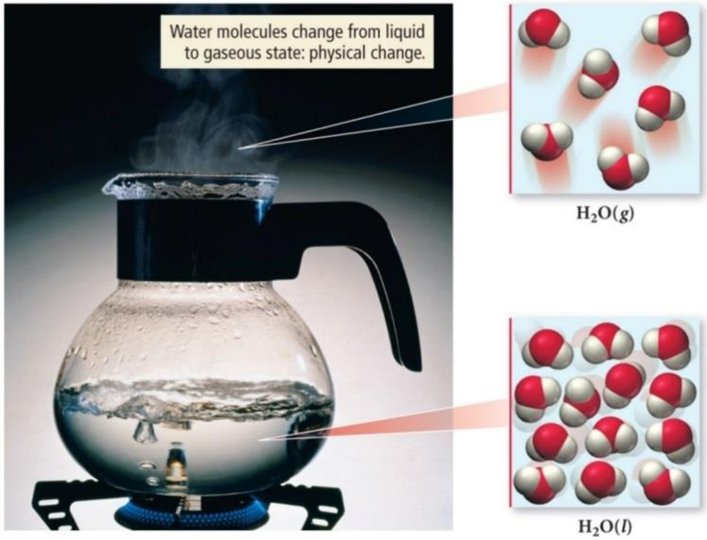
Physical Changes:

- A process that does NOT cause a substance to become a different substance (i.e. only the appearance (state or shape) is changed, but NOT the chemical composition).
- Physical changes are **reversible**.

Example 1: when water (H_2O) boils, it changes its state from liquid to gas.

- The gas remains composed of H_2O, so this is a "physical change".

Example 2: when a piece of paper is shredded, or a glass window is broken, only their shapes have changed, but their chemical compositions remains unchanged.

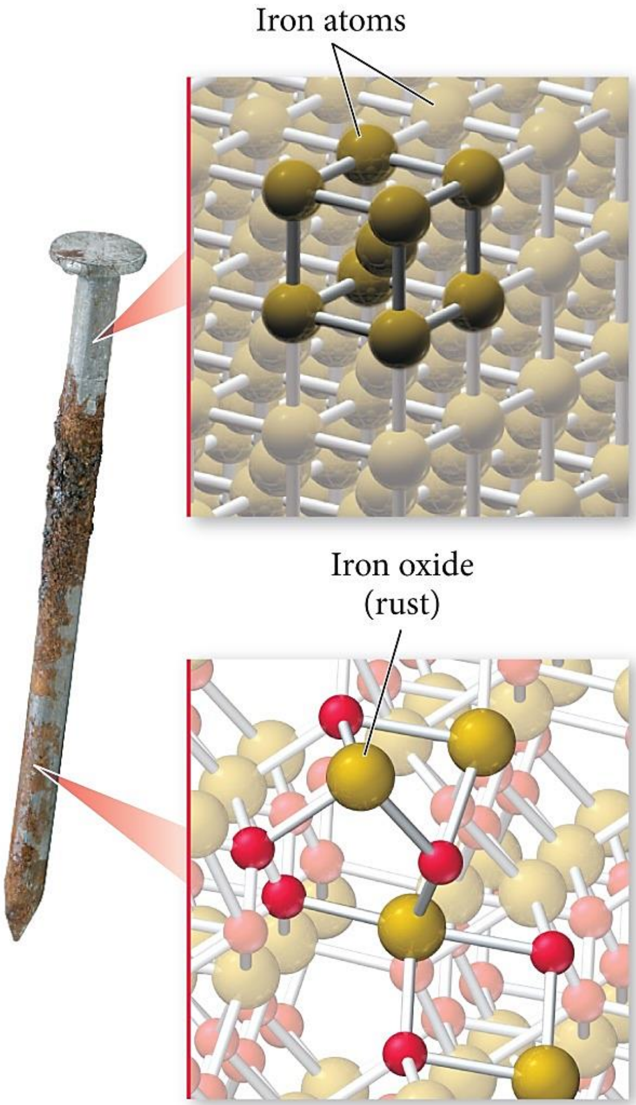


Chemical Changes:

- A process that causes a substance to change into a new substance with a new chemical composition.
- During a chemical change, atoms rearrange themselves to make different substances.
- Chemical changes are **irreversible**.

Example 1: rusting of iron is a chemical change: $4\,Fe + 3\,O_2 \rightarrow 2\,Fe_2O_3$

Example 2: burning of gasoline produces $CO_2 + H_2O$, so, it's a chemical change

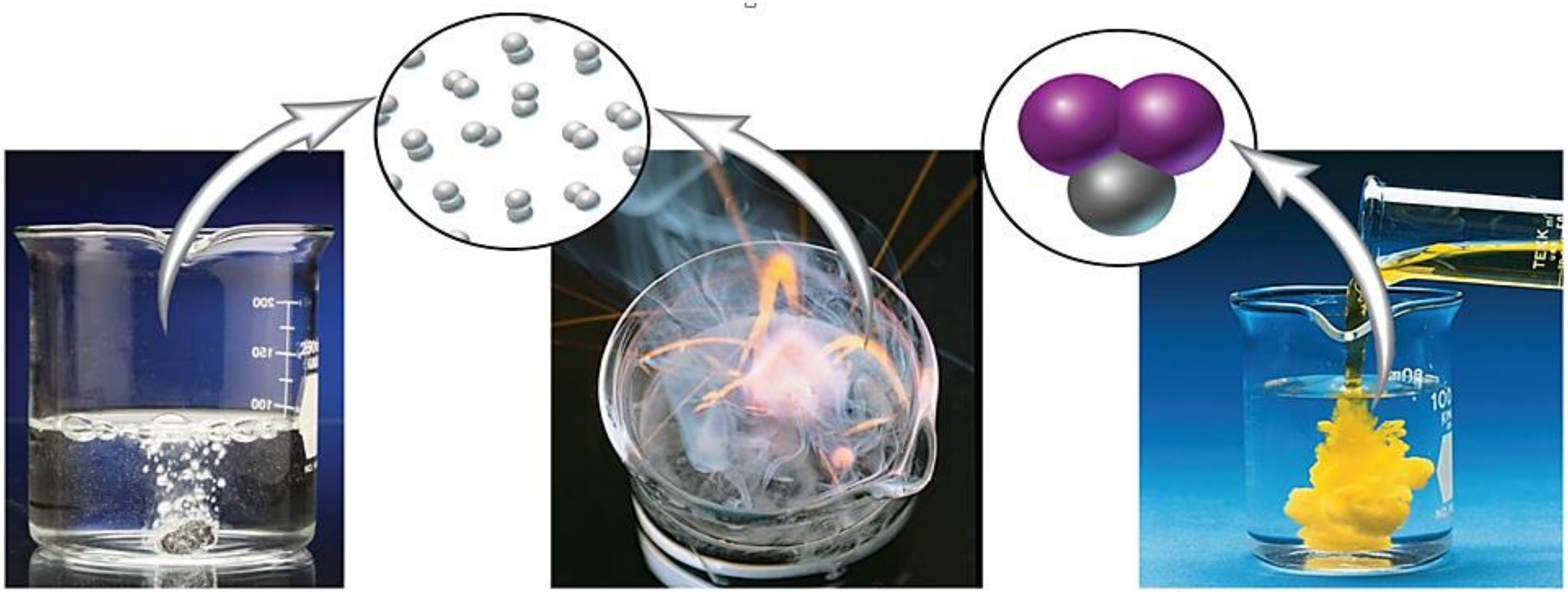


Evidences for Chemical Changes

a) Release of a gas (e.g. bubbles or smoke)

b) Emission of light or heat (e.g. burning of wood)

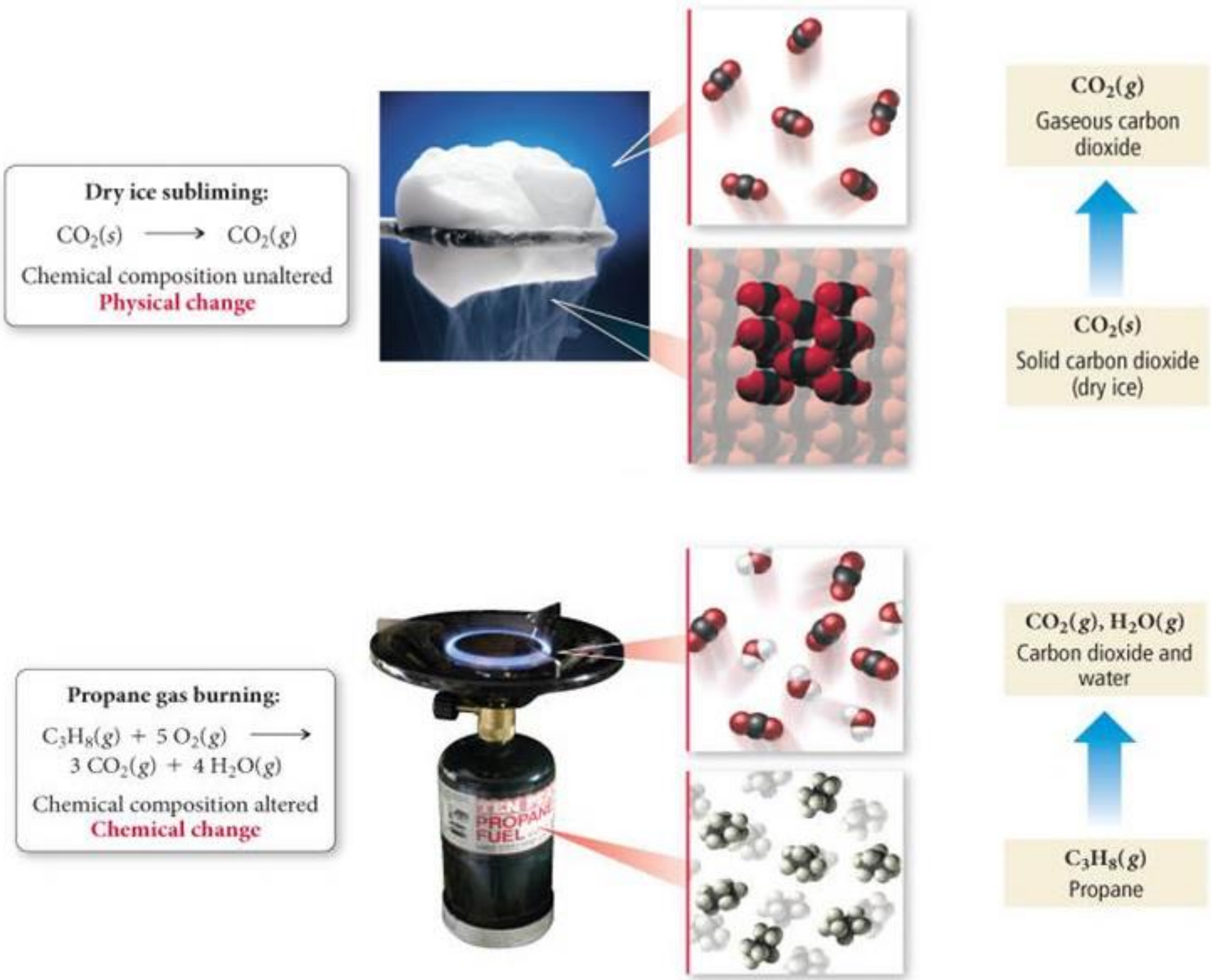
c) Permanent change in color (e.g. the brown layer of iron rust)



(a) (b) (c)

Physical and Chemical Changes: Examples

Physical Change versus Chemical Change



Physical and Chemical Properties of Matter

1. Physical Properties: any characteristic that can be measured without changing the substance's chemical identity or composition (i.e. without any chemical reactions).

Examples on Physical Properties:

- Color
- Odor
- Taste
- Density
- Melting Point
- Boiling Point
- Viscosity
- Temperature
- Hardness
- Metallic Luster
- Malleability
- Ductility

Physical and Chemical Properties of Matter

2. <u>Chemical Properties</u>: any characteristic that can be measured only by changing a substance's chemical identity or composition (i.e. in a chemical reaction).

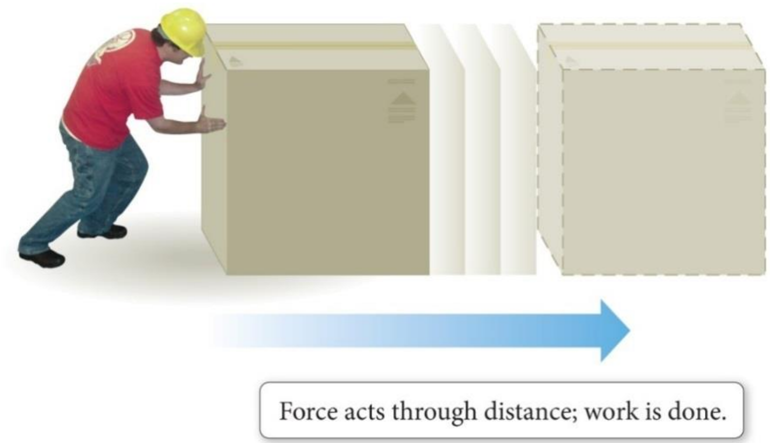
Examples on <u>Chemical Properties</u>:

- Reactivity with other chemicals (acids, water, oxygen, ….)
- Acidity and Basicity
- Flammability
- Chemical stability
- Toxicity
- Heat of combustion
- Oxidation state

1.4 Energy: A Fundamental Part of Physical and Chemical Changes

- **Energy** is the ability to do work.

- **Work** is defined as the action of a force through a distance.



- When you push a box across the floor or pedal your bicycle across the street, you have done **work**.

> **The law of conservation of energy:** energy can neither be created nor destroyed, but only changes from one form to another.

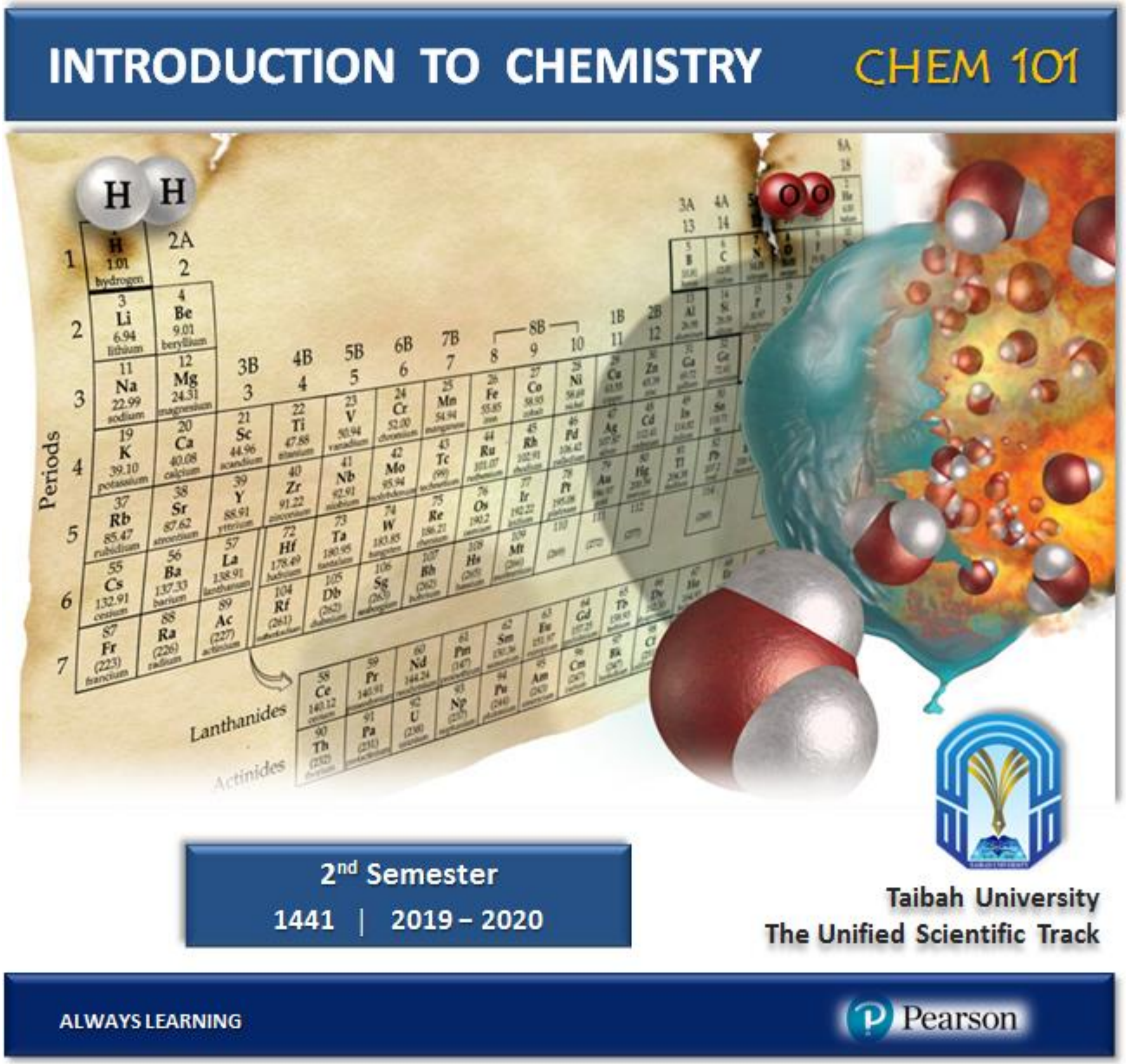
Potential and Kinetic Energy

- **Potential energy**, **PE**: is any form of stored energy; it results from position or composition (examples: chemical and nuclear energy)
- **Kinetic energy**, **KE**: is the energy matter has as a result of its motion (examples: thermal and electrical energy).
  - ✓ Energy can be converted between the two types.
  - ✓ All substances have both potential and kinetic energies, regardless to their physical states.
  - ✓ **Solids** have the lowest kinetic energy, and **gases** have the greatest kinetic energy.
  - ✓ As we increase the temperature of a substance, its kinetic energy increases.

Assessment

Identify the following as chemical or physical changes or properties:

1. blue color
2. melting point
3. density
4. reaction with water
5. flammability
6. hardness
7. toxicity
8. boiling point
9. reaction with acid
10. luster
11. perfume odor
12. sour taste
13. coal Burns
14. dry ice sublimes
15. Ag (Silver) tarnishes
16. milk sours
17. an apple is cut
18. fruit rot
19. heat changes H_2O to steam
20. pancakes cook
21. baking soda reacts to vinegar
22. grass grows
23. iron rusts
24. a tire is inflated
25. alcohol evaporates
26. food is digested
27. ice melts
28. paper absorbs water



Lecture Presentation

Chapter 1

Matter and Measurements

Topic 03

- Units of Measurements
- Density of Materials

1.5 The Units of Measurement

We use measurements in everyday life, for example:

- walking 2.25 **km** to the university campus,
- carrying a backpack with a mass of 12 **kg**, and
- observing when the outside temperature has reached 40**°C**.

1.5 The Units of Measurement

- **Units:** standard quantities used to specify measurements, they are critical in chemistry.

- The most common systems of units are:

  1. **The English system:** used in the United States.
  2. **The Metric system:** used in most countries.
  3. **The International System of Units (SI):** used by scientists, and it is based on the metric system.

Units in the Metric and SI Systems

- In the metric and SI systems, one unit is used for each type of measurement:

Measurement	Metric System	(SI) System
Length	meter (m)	meter (m)
Volume	liter (L)	cubic meter (m^3)
Mass	gram (g)	kilogram (kg)
Temperature	Celsius (°C)	Kelvin (K)
Time	second (s)	second (s)

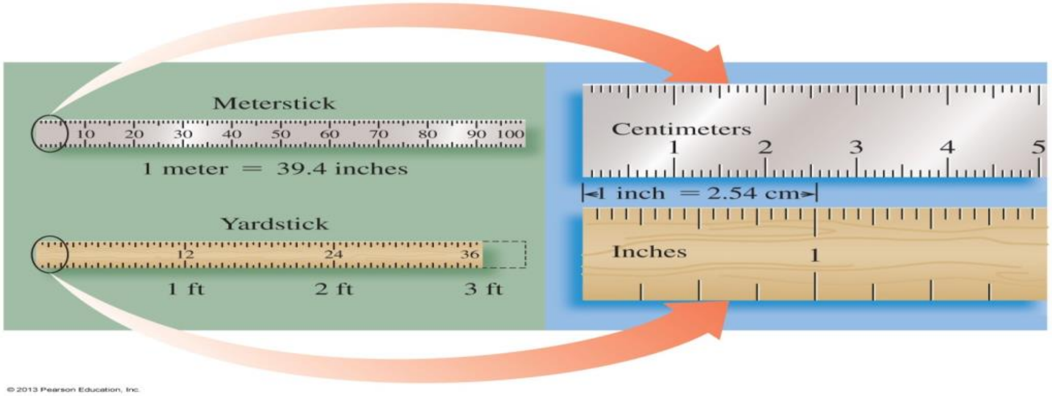
The Meter: A Measure of Length (or Distance)

Length (or distance):

- is measured using a meter stick.
- The unit **meter (m)** is used in both the metric and SI systems.
- Centimeters (cm) is used for smaller lengths.

➢ **Useful relationships between the units of length:**

- 2.54 cm = 1 in.
- 1 m = 100 cm
- 1 m = 39.4 in.
- 1 m = 1.09 yd



The Kilogram: A Measure of Mass

- The **mass** of an object is a measure of the quantity of matter within it.

- The SI unit of mass is **kilogram (kg):**

  1 kg = 2.21 lb (pound)

  1 gram = 1/1000 kg = (10^{-3} kg)

- **Weight** of an object is a measure of the gravitational pull on its matter:

  (weight ≠ mass)

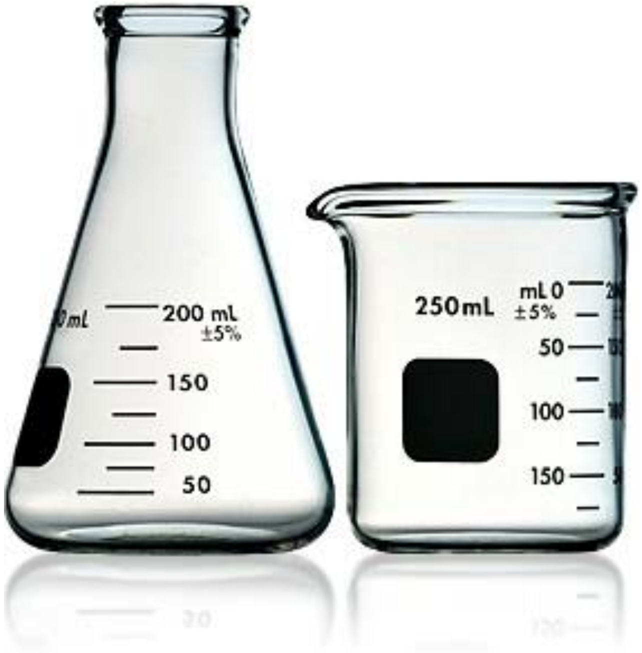
▲ A nickel (5 cents) weighs about 5 grams.

Units for Volume Measurement

- The common units for volume measurements are:

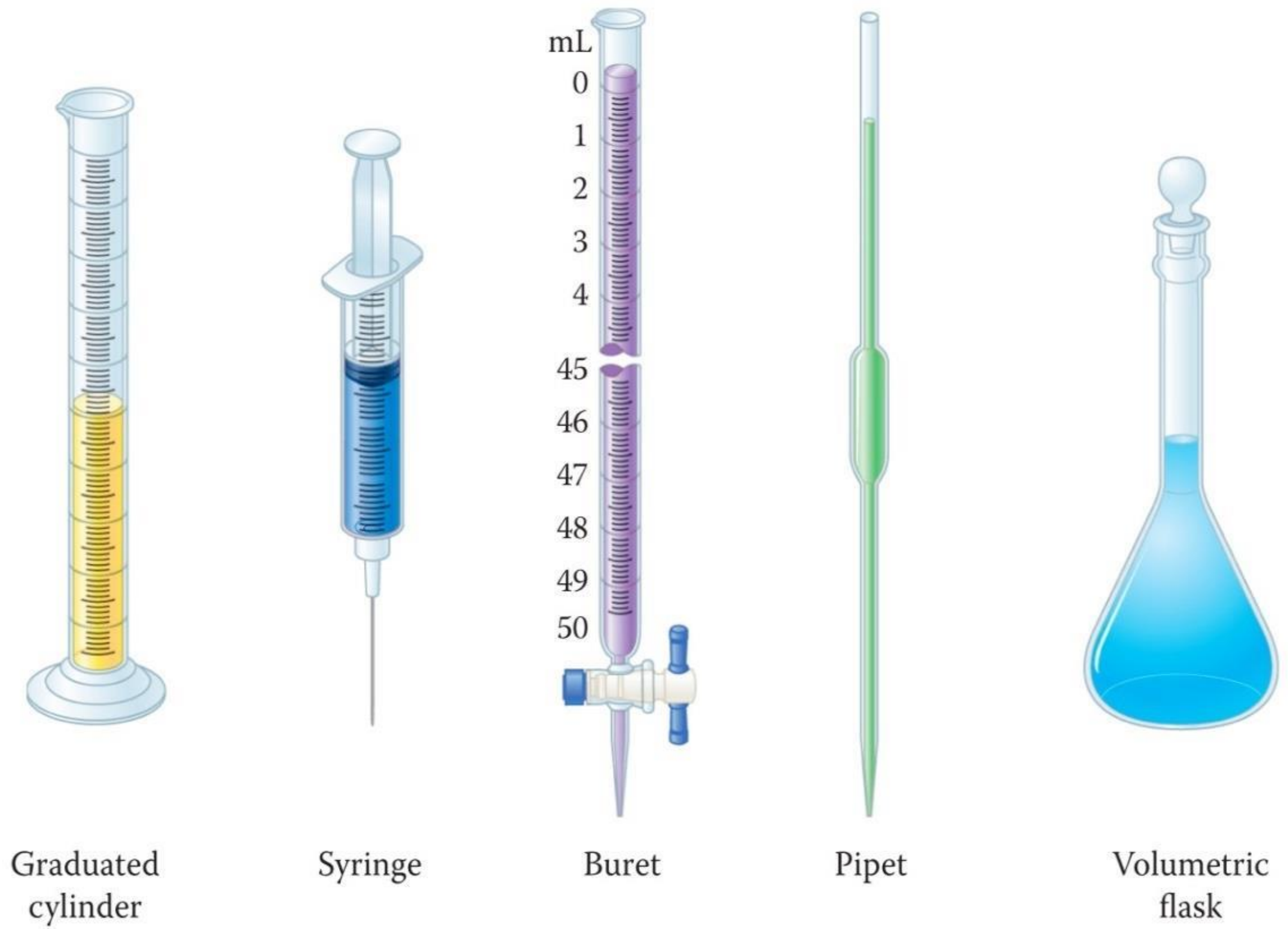
  Quart (qt), Liter (L), Milliliter (mL), and Cubic Meter (m^3)

- **Useful relationships between the units of volume:**
  - 1 L = 1000 mL
  - 1 L = 1.06 qt
  - 946 mL = 1 qt
  - 1000 L = 1 m^3



Some Lab Tools for Volume Measurement

- Volume is the amount of space occupied by a substance.



Units of Time Measurement

Time measurement:

- uses the unit **second (s)** in both the metric and SI systems.

Days, Hours, Minutes, Seconds

➢ **Useful relationships between the units of time:**

- 1 day = 24 h
- 1 h = 60 min
- 1 min = 60 s

Units of Temperature Measurement

- **Common Units of Temperature:**
  - **Fahrenheit (ºF) (English system).**
  - **Celsius (ºC) (Metric system).**
  - **Kelvin (K) (SI system).**

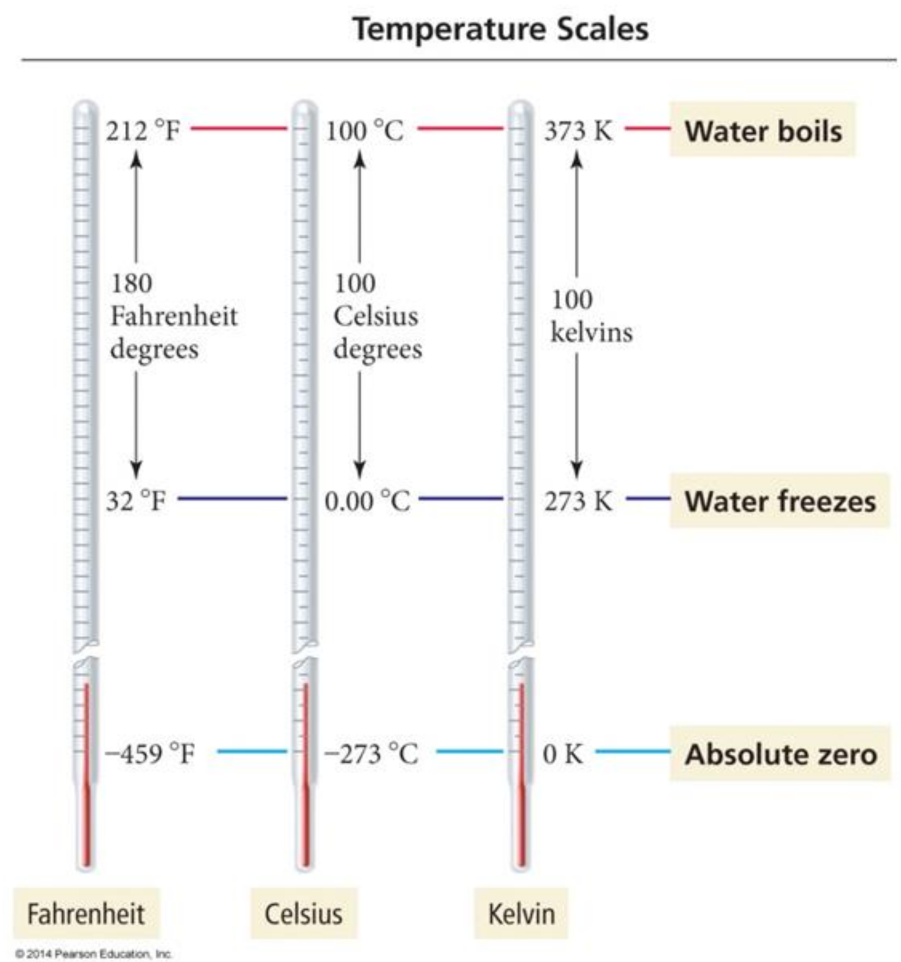
- **Example 1:** Boiling Point of Water:

  212 ºF = 100 ºC = 373.15 K

- **Example 2:** Freezing Point of Water:

  32 ºF = 0 ºC = 273.15 K

> **0 K** is called: "**absolute zero**".
> It is the lowest possible temperature in the universe!



Relationships Between Units of Temperature

- From °C to K:

$$K = °C + 273.15$$

- From K to °C:

$$°C = K - 273.15$$

- From °C to °F:

$$°F = [1.8 \times (°C)] + 32$$

- From °F to °C:

$$°C = \frac{(°F - 32)}{1.8}$$

Examples on Temperature Conversions

1- **How much does 350 ^{o}F equal in both ^{o}C and K?**

$$^{o}C = (350 - 32)/1.8 = 318/1.8 = 177\ ^{o}C$$

$$K = 177 + 273 = 450\ K$$

2- **Convert (– 40 ^{o}C) to ^{o}F:**

$$^{o}F = [1.8 \times (-40)] + 32 = -72 + 32 = -40\ ^{o}F$$

3- **Express 298 Kelvin in degree Celsius:**

$$^{o}C = 298 - 273 = 25\ ^{o}C$$

Prefix Multipliers: Changing The Value of The Unit

- The International System of Units (SI) uses the **prefix multipliers** with the standard units.

- These prefix multipliers **change the values of the units** (makes units **larger** or **smaller** than the initial unit).

- **Examples:**

- the kilometer has the prefix **kilo**, meaning **1000** meter **(or 10^3 m).**

- the millimeter has the prefix **milli**, meaning **1/1000** meter **(or 10^{-3} m).**

Prefix Multipliers: Increasing the Size of Unit

Prefixes that INCREASE the size of unit:

Prefix	Symbol	Factor	Example
Kilo	K	10^3	11 kilometer (km) = 11×10^3 m
Mega	M	10^6	3 megagram (Mg) = 3×10^6 g
Giga	G	10^9	8 gigasecond (Gs) = 8×10^9 s
Tera	T	10^{12}	5 teraliters (TL) = 5×10^{12} L
Peta	P	10^{15}	4 petameter (Pm) = 4×10^{15} m

Note: students shall memorize those prefixes: (names, symbols, and factors)!

Prefix Multipliers: Decreasing the Size of Unit

Prefixes that DECREASE the size of unit:

Prefix	Symbol	Factor	Example
deci	d	10^{-1}	3 deciliter (dL) = 3×10^{-1} L (0.3 L)
centi	c	10^{-2}	7 centimeters (cm) = 7×10^{-2} m (0.07 m)
milli	m	10^{-3}	12 millimoles (mmol) = 12×10^{-3} mol
micro	μ	10^{-6}	9 microliter (μL) = 9×10^{-6} L
nano	n	10^{-9}	4 nanograms (ng) = 4×10^{-9} g
pico	p	10^{-12}	8 picometer (pm) = 8×10^{-12} m
femto	f	10^{-15}	5 femtosecond (fs) = 5×10^{-15} s

Note: students shall memorize those prefixes: (names, symbols, and factors)!

Prefix Multipliers: A Summary with Example

10^{-15}	Peta	P	10^{15}
10^{-12}	Tera	T	10^{12}
10^{-9}	Giga	G	10^{9}
10^{-6}	Mega	M	10^{6}
10^{-3}	Kilo	K	10^{3}
Example For Conversion: To/From Meter (m)			
10^{-1}	Deci	d	10
10^{-2}	Centi	c	10^{2}
10^{-3}	Milli	m	10^{3}
10^{-6}	Micro	µ	10^{6}
10^{-9}	Nano	n	10^{9}
10^{-12}	Pico	p	10^{12}
10^{-15}	Femto	f	10^{15}

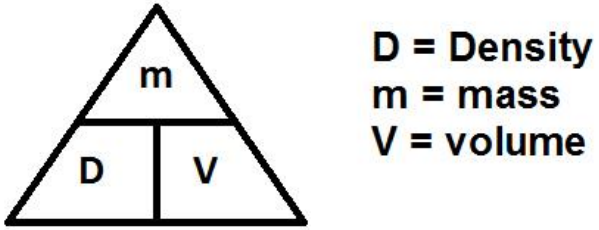
Density Of Materials

- **Material's Density** is its mass per unit volume.
- Is measured in g/L for gases.
- Is measured in g/cm^3 or g/mL for solids and liquids.

➢ **Density Expression**:

$$D = \frac{\text{mass}}{\text{volume}} = \frac{g}{mL} \text{ or } \frac{g}{cm^3} = g/cm^3$$

Useful Note: 1 mL = 1 cm^3



Density Formula

Calculating Density - Example

If a **0.258 g** sample of HDL has a volume of **0.215 cm³**, what is the density (in **g/cm³**), of the HDL?

Step 1: State the given and needed quantities:

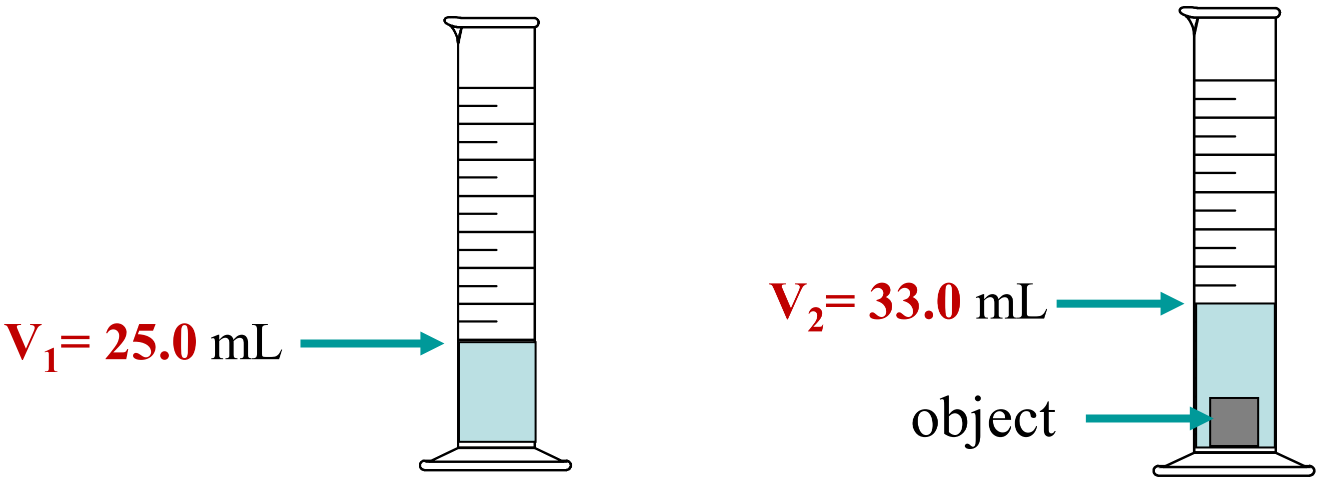
Given	Needed
0.258 g HDL 0.215 cm^3 HDL	density of HDL in g/cm^3

Step 2: Use the relation: $\text{Density} = \dfrac{\text{mass of substance}}{\text{volume of substance}}$

$$\text{Density} = \frac{0.258\ \text{g}}{0.215\ \text{cm}^3} = \frac{1.20\ \text{g}}{\text{cm}^3} = 1.20\ \text{g/cm}^3$$

Calculating Density Using Volume Displacement

What is the density (in g/cm^3) of a **48.0 g** sample of a metal if the level of water in a graduated cylinder rises from **25.0 mL** to **33.0 mL** after the metal is added?



Answer:

Step 1: Volume of metal = $V_2 - V_1$ = 33.0 – 25.0 = 8.0 mL

Step 2: Density = mass/volume = 48.0 g / 8.0 mL = 6.0 g/mL

Assessment

1- Do the following conversions:

a.	55 m	=	.... km	=	.... cm
b.	11 s	=	.... ms	=	.... ks
c.	2.7 g	=	.... pg	=	.... ng
d.	3.6 L	=	.... mL	=	.... μL

2- Express the temperature −56 °F in both °C and K.

3- Perform each of the following unit conversions:

a. 13.53 m to yd

b. 2.87 kg to lb

c. 2.45 L to qt

d. 123.7 mm to in

4- Calculate the density of penny that has a mass of 2.49 g and a volume of 0.349 cm^3.

▲ Periodic Table of the Elements

Key

Atomic number → 6; Symbol → C; Name → Carbon; atomic mass → 12.01

An element

Main group (groups 1A–2A, 3A–8A); Transition metals (groups 3B–2B); Period number; Group number

Period	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18	Period
1	1 H Hydrogen 1.008																	2 He Helium 4.003	1
2	3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18	2
3	11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95	3
4	19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.41	31 Ga Gallium 69.72	32 Ge Germanium 72.64	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80	4
5	37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9	46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3	5
6	55 Cs Cesium 132.9	56 Ba Barium 137.3	57 La Lanthanum 138.9	72 Hf Hafnium 178.5	73 Ta Tantalum 180.9	74 W Tungsten 183.8	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.1	79 Au Gold 197.0	80 Hg Mercury 200.6	81 Tl Thallium 204.4	82 Pb Lead 207.2	83 Bi Bismuth 209.0	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)	6
7	87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (267)	105 Db Dubnium (268)	106 Sg Seaborgium (271)	107 Bh Bohrium (272)	108 Hs Hassium (270)	109 Mt Meitnerium (276)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (280)	112 Cn Copernicium (285)	113 Nh Nihonium (284)	114 Fl Flerovium (289)	115 Mc Moscovium (288)	116 Lv Livermorium (293)	117 Ts Tennessine (293)	118 Og Oganesson (294)	7

Lanthanides 6	58 Ce Cerium 140.1	59 Pr Praseodymium 140.9	60 Nd Neodymium 144.2	61 Pm Promethium (145)	62 Sm Samarium 150.4	63 Eu Europium 152.0	64 Gd Gadolinium 157.3	65 Tb Terbium 158.9	66 Dy Dysprosium 162.5	67 Ho Holmium 164.9	68 Er Erbium 167.3	69 Tm Thulium 168.9	70 Yb Ytterbium 173.0	71 Lu Lutetium 175.0	6
Actinides 7	90 Th Thorium 232.0	91 Pa Protactinium 231.0	92 U Uranium 238.0	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (262)	7