ETHIOPIAN EDUCATION ®

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Unit 1

Fundamental Concepts of Chemistry

Unit Outline:

1.1 Definition of Chemistry
1.2 Measurements and units in chemistry
1.3 Uncertainty in measurements
1.4 Decimal place and rounding off
1.5 Significant figures
1.6 The scientific method
1.7 Some experimental skills in chemistry

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Introduction

Chemistry is the science of molecules and their transformations. It is the science not so much of the one hundred and ten elements, but of the infinite variety of molecules that may built from them.

Chemistry is useful to almost every one by

- Increasing crop production
- Manufacturing drugs and medicines
- Making building materials such as cement, paints, plastics.
- Converting natural resources into new substances.
- Preserving food and chemicals.
- Protecting the environment
- Treating drinking water with chlorine.
- Manufacturing cosmetics, soaps, detergents, clothes, shoes, ...

In general, chemistry is important and interesting science.

1.1 Definition of Chemistry

Chemistry is the science of matter and its transformation.

Matter is anything that has mass and occupies space. It undergoes two changes. These are

Chemical change

- It is a change in which new substance with new property is formed.
- It is difficult to reverse chemical change.
- It is accompanied by heat change

Example

All chemical reactions.

Physical Change

- It is a change in which no new substance is formed.
- It is easily reversed.
- It is accompanied by small or no heat change.

Example

- Melting of ice
- Powderingofchalk.
- FormationofCloud.

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All measurable properties of matter fall into two categories: **Extensive and intensive properties.**

Extensive property depends on how much matter is being considered.

Example
Mass, length, Volume, etc

Intensive Property does not depend on the amount of matter being considered.

Example
Temperature, density, melting point etc

The Scope of chemistry is the extent of its applications in production of new substances, discovering new substances, understanding matter and improving the living conditions for the society.

Illustrative Example

1. Which of the following is not chemical change?
   A. Souring of milk
   B. Decaying of food
   C. Rusting of iron
   D. tearing of paper

Explanation: Tearing of paper is change in size or shape, thus it is physical change.

Answer: D

2. All of the following are chemical properties. EXCEPT
   A. Fermentation of sugar.
   B. Reactivity of Sodium with water.
   C. Water boils at 100°C.
   D. Gold does not react with water.

Answer: C

3. Classify the following as intensive or extensive properties.
   a. Volume
   b. Color
   c. Area
   d. Thickness
   e. Density
   f. Freezing point
   g. Length
   h. Conductivity
   i. Specific Gravity

Answer

- a, c d, and g are extensive properties.
- b, e, f, h and i are intensive properties.

4. Write the main uses of chemical changes.

Explanation: To produce new substances and energy.

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Major Fields of Chemistry

Organic Chemistry

- Is the study of carbon compounds except carbonates, hydrogen carbonates, cyanides, cyanate, thiocyanate, and oxides of carbon.

Inorganic Chemistry

- Is the chemistry of all elements and inorganic compounds.

Analytical Chemistry

- Is concerned with the quantitative determination and qualitative identification of a sample of matter.

Physical Chemistry

- Is the study of the relationship between matter and energy.

Bio Chemistry

- Deals with processes and the structure of substances in living things.

Other branches of chemistry include:

- Textile chemistry
- Polymer Chemistry
- Water Chemistry
- Agricultural Chemistry
- Environmental Chemistry
- Food Chemistry.
- Geochemistry
- Astrochemistry

1.2 Measurements and Units in Chemistry

The study of chemistry depends heavily on measurement.

Measurement is the act of finding the magnitude of a given quantity interms of units.

Measured quantities are classified into two: **Basic and derived quantities.**

Table 1.1. Basic quantities and their basic SI units

Basic quantity	Name of unit	Symbol of unit
Length	Meter	m
Mass	Kilogram	Kg
Time	Second	S
Electrical current	Ampere	A
Temperature	Kelvin	K
Amount of substance	Mole	mol
Luminous intensity	Candela	Cd

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Mass is the measure of the quantity of matter in an object.

Its SI unit is kilogram. Other Units of mass are gram (g), milligram (mg).

$$\boxed{1\ kg = 1000g = 10^6 mg}$$

Temperature is the measure of the average kinetic energy of molecules.

Its SI unit is Kelvin. Other units of temperature are degree Celsius (°C), degree Fahrenheit (°F).

Lowest attainable temperature is -273.15°C or 0K.

$K - 0 = °C - (-273.15)$

$K = °C + 273.15$

$$\boxed{\begin{aligned} K &= °C + 273 \\ °C &= K - 273 \end{aligned}}$$

The relationship between degree Celcius and degree Fahrenheit is obtained from boiling and freezing points of water.

Boiling point of water is 100°C or 212°F.

Freezing point of water is 0°C or 32°F.

Subtract freezing point from both sides:

$$\frac{°C - 0}{°F - 32} = \frac{100 - 0}{212 - 32}$$

$$\Rightarrow \frac{°C}{°F - 32} = \frac{100}{180} = \frac{5}{9}$$

$$\therefore \boxed{\begin{aligned} °C &= \frac{5}{9}(°F - 32) \\ °F &= \frac{9}{5}°C + 32 \end{aligned}}$$

Example

Convert 27°C into °F and K.

Solution:

- $°F = \frac{9}{5}°C + 32 = \frac{9}{5} \times 27 + 32 = 80.6°F$
- $K = °C + 273 = 27 + 273 = 300K$

Mole is the amount of substance that contains as many elementary particles as there are atoms in 12g of carbon isotope.

$$\boxed{1\ mole = 6.02 \times 10^{23}\ particles\ (atoms, molecules, ions, formula\ units)}$$

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$$\text{Number of mole} = \frac{\text{given mass}}{\text{molar mass}} \qquad \boxed{n = \frac{m}{M}}$$

- Number of atoms = $n \times 6.02 \times 10^{23}$
- Number of atoms in a molecule = n × sum of subscripts$\times 6.02 \times 10^{23}$

Example

Calculate the number of atoms in 8g of CH_4.

Solution:

$$n = \frac{m}{M} = \frac{8g}{16\ g/mol} = 0.5mol$$

Number of atoms = n ×sum of subscripts $\times 6.02 \times 10^{23}$

$= 0.5\ mol \times (1 + 4) \times 6.02 \times 10^{23}$ atoms/mol

$= 1.505 \times 10^{24}$ atoms

Illustrative Examples

5. Which of the following is not basic SI unit?

A. Second B. Kelvin C. Gram D. Meter

Answer: C

6. -40°F is _____ K

A. -40 B. 233 C. 273 D. 300g

Solution:

-40°F = -40°C

$\Rightarrow K = °C + 273 = -40 + 273 = 233K$

Answer: B

7. How many atoms are present in 22 grams of carbon dioxide, CO_2? (EUEE)

A. 3.01×10^{23}

B. 6.02×10^{23}

C. $2 \times 6.02 \times 10^{23}$

D. $1.5 \times 6.02 \times 10^{23}$

Solution:

$$n = \frac{m}{M} = \frac{22g}{44g/mol} = 0.5mol$$

Number of atoms = n × sum of subscripts $\times 6.02 \times 10^{23}$

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$= 0.5 \times (2 + 1) \times 6.02 \times 10^{23}$

$= 1.5 \times 6.02 \times 10^{23}$ atoms

Answer: B

8. One mole of $(NH_4)_2HPO_4$ contains _____ moles of hydrogen atoms.

A. One B. Five C. Six D. Nine

Solution:

Number of moles of atoms in a compound = number of mole of compound × subscript of the atom

$\Rightarrow (NH_4)_2PO_4 = N_2H_9PO_4$

number of H - atom = $n(NH_4)_2HPO_4 \times 9$

$= 1 \text{ mol} \times 9 = 9 \text{ moles}$

Answer: D

9. How many moles are there is 159g of alanine, $C_3H_7NO_2$? (EUEE)

Solution:

$M, C_3H_HNO_2 = 3 \times 12 + 7 \times 1 + 1 \times 14 + 2 \times 16 = 89 g/mol$

$$n = \frac{m}{M} = \frac{159g}{89g/mol} = 1.8mol$$

10. What is the lowest attainable temperature in °C and K?

Answer: -273.15°C or 0K

Table 1.2. Some derived quantities and their SI Units

Derived quantity	Name of SI unit	Symbol of Unit
Area	Square meter	m^2
Volume	Cubic meter	m^3
Density	Kilogram per cubic meter	kg/m^3
Pressure	Pascal	Pa
Concentration	Mole per liter	mol/L
Weight	Newton	N
Force	Newton	N
Work	Joule	J
Speed	Meter Per second	m/s

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Volume is the space occupied by an object. Its SI unit is cubic meter (m^3). Other units of volume are liter (L), milliliter (mL), cubic centimeter (cm^3) and cubic decimeter (dm^3).

$1m^3 = 10^3dm^3 = 10^3L = 10^6cm^3 = 10^6mL$

Note

Volumetric flask, graduated cylinder, pipet and buret are laboratory equipments used to measure volume.

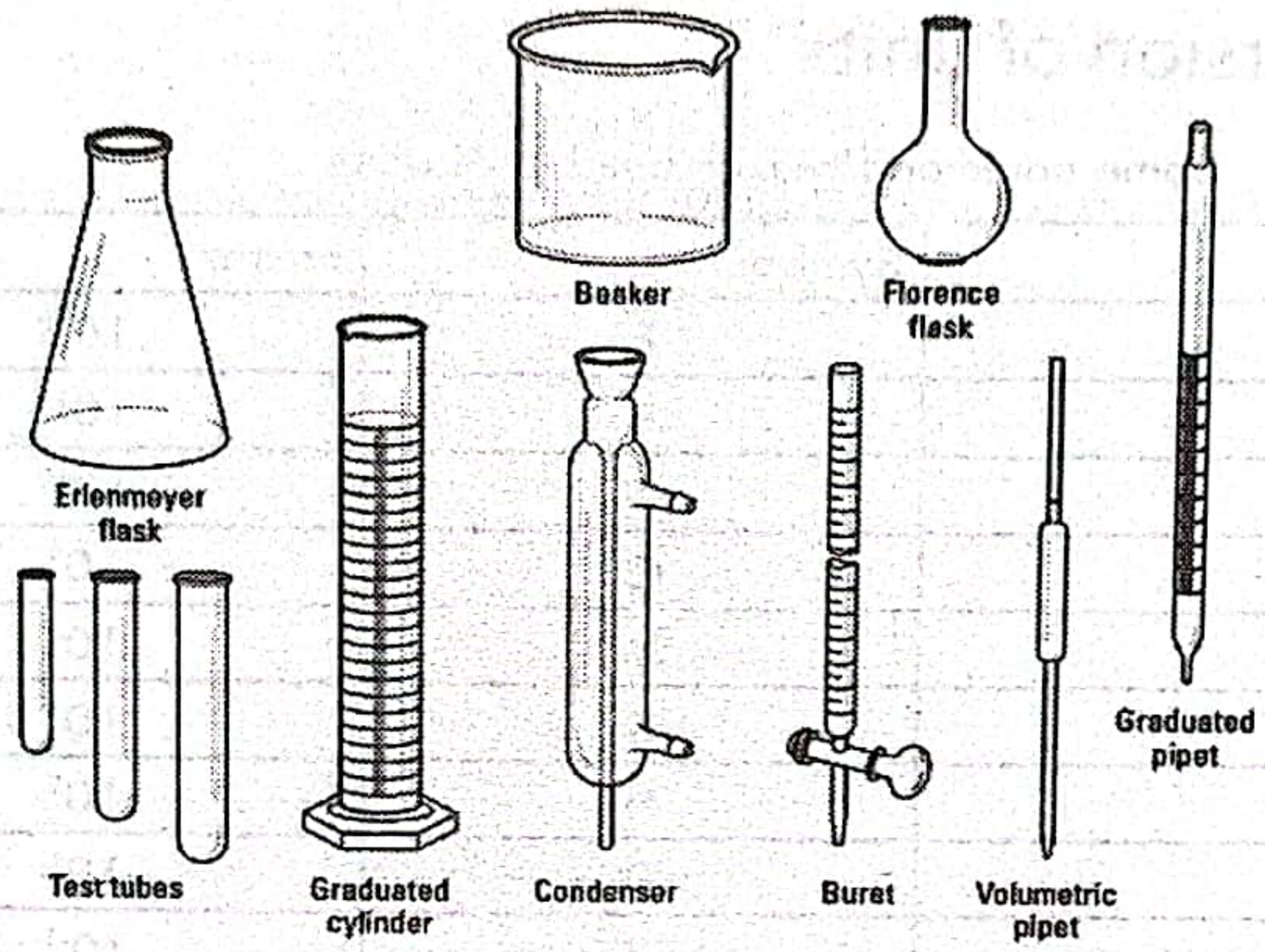


Figure 1.1 Thesepiecesofglasswareareusedtomeasurevolumeofliquids

Pressure is force per area.

$$P = \frac{F}{A}$$

The SI unit of pressure is Pascal (Pa). Other units of pressure are atmosphere (atm), millimeter of mercury (mmHg), centimeter of mercury (cmHg) and torr.

1 atm = 76cmHg = 760 mmHg = 760 torr = 101325 Pa

→ 1 mmHg = 1 torr

1 kPa = 1000Pa

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Density (ρ): is mass divided by volume.

$$\rho = \frac{m}{V}$$

The SI unit of density is kilogram per cubic meter (Kg/m^3). Other units are gram per mililiter (g/mL), gram per liter (g/L).

$$1Kg/m^3 = 1g/L = 0.001\ g/mL$$

Common Prefixes used in SI units and Conversion of units

Table 1.3. Some common Prefixes used in SI units

Prefix	Symbol	Factor
Exa	E	10^{18}
Peta	P	10^{15}
Tera	T	10^{12}
Giga	G	10^{9}
Mega	M	10^{6}
Kilo	K	10^{3}
Hecto	H	10^{2}
Deca	D	10^{1}
Deci	d	10^{-1}
Centi	c	10^{-2}
Milli	m	10^{-3}
Micro	μ	10^{-6}
nano	n	10^{-9}
Pico	P	10^{-12}
Femto	F	10^{-15}
Atto	a	10^{-18}

Conversion of Units

➥ To convert one unit into another, first find the conversion factor.

Example

.....

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Convert 160cm into dm.

Solution:

$$\frac{1\text{ cm}}{1\text{ dm}} = \frac{10^{-2}\text{m}}{10^{-1}\text{m}}$$

$$\Rightarrow \frac{1\text{ cm}}{1\text{ dm}} = \frac{10^{-2}}{10^{-1}}$$

$$1\text{ cm} = 10^{1}\text{ dm}$$

$\Rightarrow$ The conversion factor is 10^{1}.

$$\therefore 160\text{cm} = 160\,\text{cm} \times 10^{1}\,\frac{\text{dm}}{\text{cm}} = 16\text{dm}$$

Illustrative Examples

11. How many milligrams are in 0.2 decigrams? (EUEE)

A. 20 B. 2000 C. 0.002 D. 0.00002

Solution:

$$\frac{1\text{ dg}}{1\text{ mg}} = \frac{10^{-1}\text{g}}{10^{-3}\text{g}}$$

$$\frac{1\text{ dg}}{1\text{ mg}} = 10^{2}$$

$$\Rightarrow 1\text{ dg} = 10^{2}\text{ mg}$$

$$\therefore 0.2\text{dg} = 0.2 \times 10^{2}\text{mg} = 20\text{mg}$$

Answer: A

12. What fraction of a second is a picosecond, ps? (EUEE)

A. 10^{-15} B. 10^{-12} C. 10^{-10} D. 10^{-2}

Answer: B

13. Convert the following

a. 100cm^3 into dm^3

b. 6.54m^3 to μL

c. 0.65 ng to μg

Solution:

a. $$\frac{1\text{cm}^3}{1\text{dm}^3} = \frac{10^{-6}\text{m}^3}{10^{-3}\text{m}^3} \Rightarrow 1\text{cm}^3 = 10^{-3}\text{dm}^3$$

$$\therefore 100\text{cm}^3 = 100 \times 10^{-3}\text{ dm}^3 = 0.1\text{dm}^3$$

b. $$\frac{1\text{m}^3}{1\mu\text{L}} = \frac{10^{3}\text{L}}{10^{-6}\text{L}} \Rightarrow 1\text{m}^3 = 10^{9}\,\mu\text{L}$$

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$\therefore 6.54m^3 = 6.54 \times 10^9 \mu L$

c. $\frac{1ng}{1\mu g} = \frac{10^{-9}g}{10^{-6}g} \Rightarrow 1ng = 10^{-3}\mu g$

$\therefore 0.65\ ng = 0.65 \times 10^{-3}\mu g = 6.5 \times 10^{-4}\ \mu g$

14. From the seven basic SI units, which one has a prefix?

Explanation: The kilogram has a prefix kilo (10^3). The rest have no prefixes.

15. Given that the density of gold is 19.3 g/cm^3. Calculate the volume of a gold hugget with a mass of 5.98g.

Solution:

$$\rho = \frac{m}{V} \Rightarrow V = \frac{m}{\rho} = \frac{5.98g}{19.3g/cm^3} = 0.310cm^3$$

1.3 Uncertainty In Measurements

➥ Uncertainity in measurement can be caused by

1. Observational error
2. Instrumental error
3. Variability in the item measured.
4. Environment where the measurement is being made.

➥ Chemistry uses two types of number: **exact and inexact**.

Exact numbers include numbers with defined values and those that are obtained by counting.

Example

There are 1000m in 1km.

Inexact numbers are measured numbers.

Example

The length of a table.

Precision and Accuracy in measurements

Precision tells us how closely multiple measurements of the same thing are to one another.

Accuracy tells us how close a measurement is to the **true value**.

Example

The true mass of an aspirin tablet is 0.370g. Three students are asked to determine the mass of the tablet. Each student

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weighs the aspirin tablet three times. The results (in grams) are:

Trial	Student A	Student B	Student C
1	0.335	0.357	0.369
2	0.331	0.375	0.373
3	0.333	0.338	0.371
Average	0.333	0.357	0.371

- Student A's results are precise but not accurate.
- Student B's results are neither precise nor accurate.
- Student C's results are both precise and accurate.

Errors = |True value - measured value|

There are two types of errors in measurement. These are:

Random errors - vary in direction and magnitude.

- mainly affect the precision of measurement

Systematic errors - lie in one direction of the true value, that is all above or all below the true value.

- caused by instruments and personal errors.

Illustrative Example

16. The difference between an accepted value and an experimental value is called a (an).
    A. Error
    B. Measured value
    C. Percent error
    D. Precision

    Answer: A

17. The true value for the density of water at 4°C and one atmosphere is 1g/mL. Student A and B determined the density of water at 4°C and 1 atm and obtained the following values in g/mL.

Student	Trials						Average
A	1	2	3	4	5	6	
	0.91	0.92	0.90	0.88	0.91	0.90	0.90
B	1.10	0.99	1.01	0.97	0.91	1.03	1.00

a. Which set of data is the most precise?

b. Which set of data is the most accurate?

c. Which student has made random and which one has made systematic error.

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Explanation:

a. Student A has made precise but less accurate measurement.
b. Student B has made accurate but less precise measurement.
c. Student A has made systematic error and student B has made random error.

18. Indicate which of the following allow you to give exact numbers when you measure them.
    a. The mass of a bar copper.
    b. The number of centimeters in a meter.
    c. The area of a rectangle.
    d. The number of hours in a day.
    e. The thickness of a black board.

Explanation:

- b and d can be expressed by exact numbers.
- a, c and e can be expressed by inexact numbers.

19. Four students measured the mass of a piece of gold whose true mass is 18.0g. Their results are: 18.19, 17.4g, 19.0g and 17.8g.
    a. What is the best estimate for the mass of a piece of gold.
    b. Explain whether the results are precise or accurate.

Explanation:

a. The best estimate value is the average value of the four students.

$$= \frac{18.1+17.4+19.0+17.8}{4} = 18.075 \approx 18.1\text{g}$$

b. The average value is close to the true value. Thus, they have made accurate measurement.

20. Write two causes for uncertainty in measurement.

Answer:

Observational and personal errors.

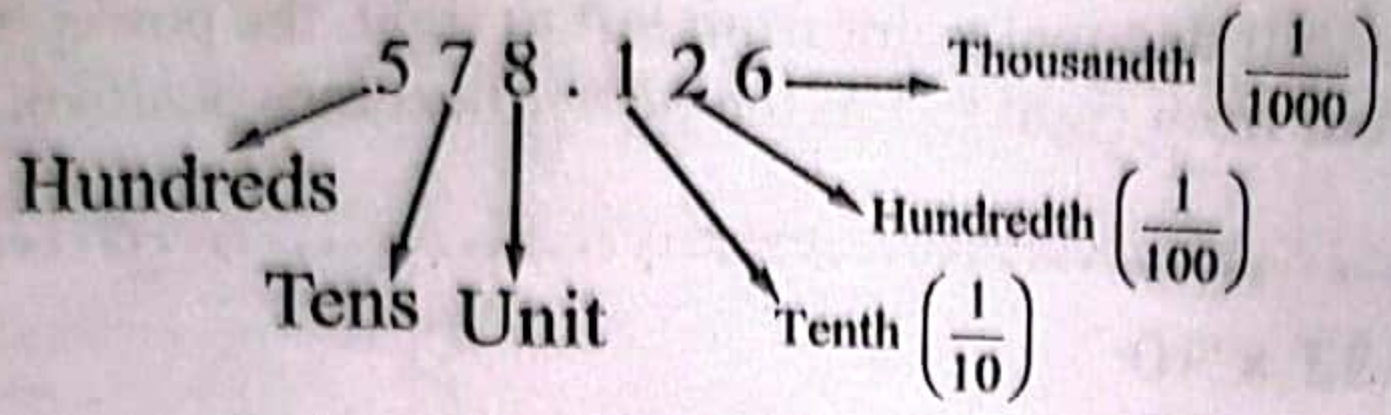
1.4 Decimal place and Rounding off

A decimal place is the number of digital to the right of the decimal point.

A decimal point (.) is used to separate the fraction part from the whole

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number part.



Rounding off is finding a number that is closest to a given number but with fewer digits.

Rules for rounding of a number

1. If the digit to be removed is from 0 to 4, the preceding digit remains same.

Example

Round off 6.7314 to two decimal place. Ans: 6.73

2. If the digit to be removed is from 6 to 9, add one to the preceding digit.

Example

Round of 7.8374 to two decimal place. Answer: 7.84

3. If the digit to be removed is 5, the preceding digit is unchanged if it is even and increases by one if it is odd.

Example

Round of 8.7635 to 3 decimal place.

$\Rightarrow$ 8.764

Round of 8.7625 to 3 decimal place.

$\Rightarrow$ 8.762

Note

1. If 5 is followed by zeros only, the left most digit is unchanged.
2. If 5 is followed by non - zero digits, the left most digit is increased by one.

Example

21.6500$\Rightarrow$ 21.6 but 21.6534 $\Rightarrow$ 21.7

1.5 Scientific Notation

➥ It is a way of expressing large or small numbers as a factor of the powers of 10.

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- It can be expressed by the form $y \times 10^x$, where, y is 1 to 10 excluding 10 and x is any integer.
- When we shift decimal point from left to right, the power is negative but from right to left the power becomes positive.

Example

$0.000000132 = 1.32 \times 10^{-7}$

$624000000000 = 6.24 \times 10^{12}$

Significant figures

- Significant figures are the meaningful digits in a reported number. The last digit in a measured number is referred to as **uncertain digits** and the uncertainty associated with a measured number is generally considered as ± 1 in the place of the last digit.
- When we report the width of a card to be 2.5cm, we are implying that its width is 2.5 ± 0.1cm.
- If the width of the card is 2.45cm, this implies its width is 2.45 ± 0.01cm.

Rules to determine number of significant figures (sf)

1. Any digit that is not zero is significant. Eg. 156.3 has four sf
2. Zeros located between non - zero digits are significant. Eg.406 has three sf
3. Zeros to the left of non - zero digit are not significant. Eg. 0.00231 has three sf
4. Zeros to the right of the last non - zero digit are significant if the number contains a decimal point. Eg. 3.100 has four sf
5. Zeros to the right of the last non - zero digit in a number without decimal point may or may not be significant. To avoid ambiguity, use scientific notation.

Example

500 may have different significant figures.

$\Rightarrow 5 \times 10^2 = 1$ sf

$\Rightarrow 5.0 \times 10^2 = 2$ sf

$\Rightarrow 5.00 \times 10^2 = 3$ sf

Calculations with significant figures

1. **In addition and subtraction**, the answer cannot have more digits to

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the right of the decimal point than any of the original numbers.

Example

103.50 ← two digits after the decimal place

+0.316 ← three digits after the decimal place

102.816 ← round to 102.82

2. **In multiplication and division,** the number of significant figures in the product or quotient is determined by the original number that has the smallest number of significant figures.

Example

$1.4 \times 8.011 = 11.2154 \leftarrow$ round to 11

3. **Exact numbers** can be considered to have an infinite number of significant numbers and do not limit the number of significant figures in a calculated result.

Illustrative Example

21. How many significant figures are there in 0.0500? (EUEE)

A. One B. Three C. Two D. Four

Explanation: Zeros before 5 are not significant, but after 5 are significant due to the presence of decimal point.

Answer: B

22. A measuring cylinder has a precision of ±0.01mL. A liquid sample that has a volume of 10mL is measured by measuring cylinder. How many significant figures should be reported for the measurement? (EUEE)

A. 4 B. 5 C. 6 D. 7

Explanation: A precision ±0.01mL implies there are two decimal places after the decimal place of the measured volume.

⇒10.00mL

Answer: A

23. Determine the number of significant figures.

a. 1.056 c. 62415

b. 0.000691 d. 600

Solution:

a. 4 b. 3 c. 5 d. ambiguous

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24. Solve the following and write your answer to the correct significant figures.

a. 317.5mL + 0.675mL

b. 47.80L - 2.075L

c. 13.5g ÷ 45.18L

d. 6.25cm × 1.175cm

Solution:

a. 318.2mL

b. 45.73L

c. 0.299 g/L

d. 7.34cm²

25. Write the following in the given number of significant figures by using scientific notation.

a. 3467893412 (2 sf)

b. 46789.8763 (3 sf)

c. 0.00000000069998 (4 sf)

d. 613423 (1 sf)

Solution:

a. 3.5×10^{9}

b. 4.68×10^{4}

c. 7.000×10^{-10}

d. 6×10^{5}

1.6 The Scientific Method

Scientific method is logical, careful and systematic approach to solve scientific problems.

Steps in Scientific Method

1. **Observation** is looking through carefully the nature of an event.
   - ➦ There are two types of observations, qualitative and quantitative observations.

Qualitative observations have no numbers with them.

Example

The sky is blue. Honey is sweet.
Water is liquid.

Quantitative Observations have numbers with them and are more useful.

Example

A book weighs 10 grams. Water boils at 100°C.

2. **Hypothesis** is tentative explanation for an observation.
   - ➦ Hypothesis is useful if it can be used to make predictions.
3. **Experiment or testing hypothesis** is examining facts about hypothesis.

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4. **Theory** is tested hypothesis.
   - ⇀ It always serves as aid to new experiments and is constantly being tested.
   - ⇀ It can be rejected or modified when it is proven incorrect.
5. **Scientific law** is a general statement that summarizes facts that come from many experiments.

Example
The law of gravitational force.
The law of conservation of mass.

Illustrative Examples

26. What is the first step of the scientific method? (EUEE ... 2008 E.C)
   A. Making observations
   B. Forming a hypothesis
   C. Performing an experiment
   D. Predicting the result of an experiment.

Answer: A

27. Which of the following represents a tentative explanation of a certain scientific law?
   A. Hypothesis
   B. Observation
   C. Experiment
   D. Theory

Answer: D

28. In order to advance to the level of a theory, a hypothesis should (EUEE... 2008)
   A. be obviously accepted by most people.
   B. be repeatedly confirmed by experimentation.
   C. be a fully functional experiment.
   D. report the past experience.

Answer: B

29. Write the difference between quantitative and qualitative observation.

Explanation: Quantitative observations are numerical data while qualitative observations means non - numerical data.

30. Classify the following as quantitative or qualitative observations.
   a. The leaves of most trees are green in summer.
   b. The apple consists of over 95% water.

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c. Chemistry is an easy subject.
d. I got 90% in my last chemistry examination.
e. My hair is black.

Solution:

➥ **a, c and e** are qualitative observations.
➥ **b and d** are quantitative observations.

31. State the limitations of the scientific method.

Explanations: Scientific method is too rigid and doesn't permit conclusions to be reached from limited informations.

32. Write the steps in the scientific method.

Solution:

Observation ⟶ Hypothesis ⟶ Experiment ⟶ Theory ⟶ Law

1.7 Some Experimental Skills in Chemistry

Chemistry is an experimental science that relates theoretical study with practical activities.

Chemistry laboratory apparatus

➥ Laboratory equipments are classified into three based on their uses. These are:

1. **Reaction vessels** are used for the occurrence of a reaction.

Example

Beakers, flasks, test tubes.

2. **Measuring equipments** are used to measure the quantities of substances.

Example

Measuring cylinder, pipet, buret, balance, thermometer, volumetric falsks.

3. **Supporting and heating devices**

Example

Stand, clamp, test tube holder, tong, Bunsen burner, splint lamp.

Chemistry Laboratory Safety rules

➥ Follow the following laboratory safety rules when you perform an

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activity in the laboratory.

1. Wear gloves, goggles and laboratory coat.
2. Wear shoes that cover your feet.
3. Tie back lose hair.
4. Do not taste any chemical.
5. Do not suck the solutions in the pipet by mouth, rather use a rubber suction bulb or other device to fill a pipette.
6. To dilute concentrated acids, add acid to water (AW!) but not water to acid (WA!).
7. Never look directly into a test tube.
8. Point test tubes that are being heated away from you and others.
9. Carry out only allowed experiments.
10. Read the labels on the reagent bottles carefully to make sure that you are using the right chemicals.
11. When lighting Bunsen burner, first light the match stick then turn on the gas.
12. If corrosive bases are swallowed, weak acids such as acetic acid or lemon juice is administered.
13. If corrosive acids are swallowed, administer weak bases such as magnesium hydroxide ($Mg(OH)_2$), aluminum hydroxide ($Al(OH)_3$).
14. If an acid or a base enters your eye, wash with water repeatedly and then consult doctor.
15. If a concentrated acid is spilled or splashed on your body, first wash the affected part with running water and then wash with 10% of Na_2CO_3 solution.
16. Before leaving the laboratory, clean your apparatus, work surface and your hands well.

Writing a laboratory Report

- A laboratory report is a written composition of the results of an experiment.
- To write a laboratory report include the following.

1. Title _____
2. Objective _____
3. Materials and Chemicals _____
_____

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4. Procedure _____

5. Observation _____

6. Result and discussion _____

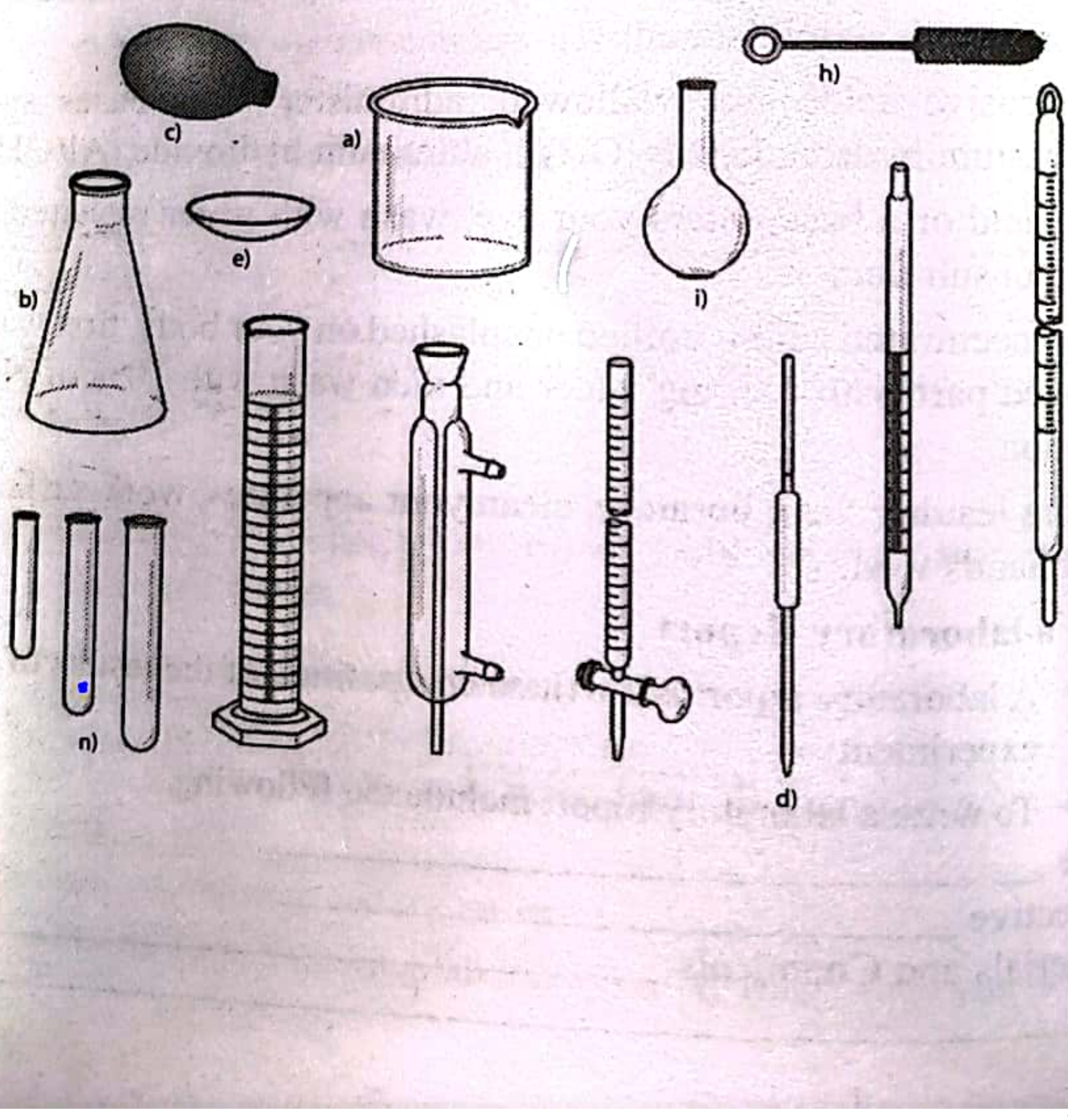
7. Conclusion _____

Laboratory report answers the following questions.

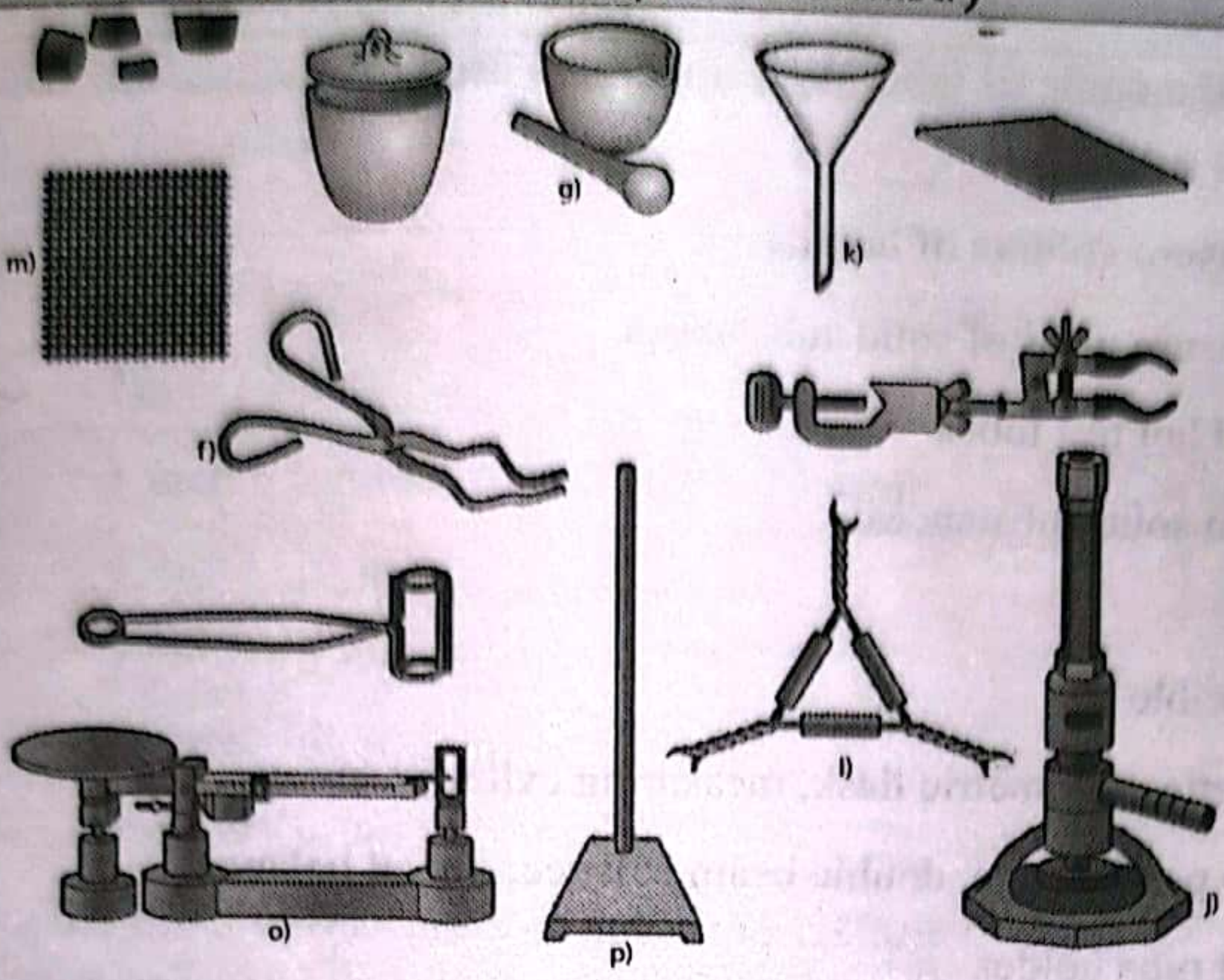
- ➦ What did you do?
- ➦ Why did you do it?
- ➦ How did you do it?
- ➦ What happened?

Illustrative Examples

33. Write the name of the following laboratory apparatus.



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Solution:

a. Beaker
b. Conical flasks
c. Glass or plastic funnel
d. Dropping pipette or teat pipette
e. Evaporating dish
f. Tong
g. Crucible and lid
h. Plastic or nickel spatuala
i. Tripod
j. Bunsen burner
k. Funnel
l. Round botton flask
m. Wire gauze
n. Test tubes
o. Balance

34. State some laboratory rules that are not stated in this book.

Solution:

- Never remove chemicals from the laboratory.
- Never return extra chemicals into the reagent bottles.
- Add boiling chips to liquids to be boiled.
- Never place chemicals directly on the pan balances, use filter paper of other equipments.
- Eating and drinking in the laboratory are not allowed.

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35. Write the name of laboratory apparatus used to

a. Melt solids.

b. Measure volume of liquids.

c. Measure mass of solid substances.

d. Told hot test tubes.

e. Burn solid substances.

Solution:

a. Crucible

b. Pipette, volumetric flask, measuring cylinder, buret.

c. Top pan balance, double beam balance, digital balance.

d. Test tube holder.

e. Tongs

36. Write a short laboratory report of the experiment of burning magnesium ribbon to produce basic oxide.

Explanation:

Title: Burning of magnesium ribbon.

Objective: To prepare basic oxide from burning of magnesium ribbon.

Materials and chemicals: Magnesium ribbon, Bunsen - burner, test tube or crucible, water, litmus paper, tong.

Procedure: Turning on the Bunsen burner, holding the magnesium ribbon with tong and burning it.

Observation: A bright light was observed when magnesium burnt in air. The white ash formed was collected in the crucible.

Result and discussion: At the end of burning, white ash was formed and dissolved in water. When the solution was checked with indicators it turned the litmus paper to blue.

Conclusion: The white ash formed was known to be basic oxide since it changed red litmus paper to blue.

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Solved Problems

37. Relate the different units of length.

Solution:

$1km = 10^3m = 10^5cm = 10^6mm = 0.62$ mile

$1m = 1.094$ yard

1 inch $= 2.54cm = 0.0254m = 25.4mm$

1 foot = 12 inch

1 yard = 3 feet

1 mile = 5280 feet

1 cm = 10mm

38. Relate the different units of mass and volume.

Solution:

- 1 kg $= 1000g = 10^6mg$
- 1 pound = 0.4536kg
- 1 ounce = 0.02835 kg
- 1 metric tone = 1000kg

Volume Units

- $1m^3 = 10^3dm^3 = 10^3L = 10^6cm^3 = 10^9mm^3$
- $1dm^3 = 1L$, $1cm^3 = 1mL$
- 1L = 1.06 quarts.

39. Show that kg/L and g/mL are equivalent units of volume.

Explanation: 1 kg = 1000g and 1L = 1000mL

$$\Rightarrow \frac{1kg}{1L} = \frac{1000g}{1000mL} \Rightarrow 1kg/L = 1g/mL$$

40. Bromine is a reddish - brown liquid. Calculate its density if 586g of the substance occupies 188mL.

Solution:

$$\rho = \frac{m}{v} = \frac{586g}{188mL} = 3.12g/mL$$

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41. Determine the number of significant figures in each of the followi measurements.

a. 0.005L
b. 60,156 tons
c. 6 kg
d. 60m
e. 9.50×10^{4}cm
f. $10.001 m^3$

Solution:

a. 1
b. 5
c. 1
d. ambigous
e. 3
f. 5

42. Carryout the following operations in the correct number of significan figures.

a. 7.310 km ÷ 5.70km
b. $(3.26 \times 10^{-3}mg) - (7.88 \times 10^{-5}mg)$
c. $(4.02 \times 10^{6}\ dm) + (7.74 \times 10^{7}\ dm)$
d. $(1.0 \times 10^{4}\ m) \times (9.9 \times 10^{6}m)$
e. $(7.95 \times 10^{4}L) \div (2.5 \times 10^{2}L)$

Solution:

a. 1.28
b. $3.18 \times 10^{-3}mg$
c. $8.14 \times 10^{7}dm$
d. $9.9 \times 10^{10}\ m^2$
e. 3.2×10^{2}

43. At what temperature does the numerical reading on a Celcius thermometer equal that on Fahrenheit thermometer?

Solution:

$$°C = °F$$

$$\frac{5}{9}(°F - 32) = \frac{9}{5}°C + 32 \text{ but } °F = °C$$

$$\frac{5}{9}°C - \frac{5}{9} \times 32 = \frac{9}{5}°C + 32$$

$$\frac{5}{9}°C - \frac{9}{5}°C = 32 + \frac{5}{9} \times 32$$

$$-\frac{56}{45}°C = 32 \times \frac{14}{9}$$

$$°C = -40$$

$\therefore$ At -40, °C and °F are equal.

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44. Chalcopyrite, the principal ore of copper contains 35% copper by mass. How many grams of copper can be obtained from $5.11 \times 10^3 kg$ of the ore?

Solution:

Percentage means parts per hundred $\left(\frac{1}{100}\right)$

$\Rightarrow$ Mass of copper = 35% × mass of ore

$$= \frac{35}{100} \times 5.11 \times 10^3 \text{ kg}$$

$$= 1.8 \times 10^3 \text{kg}$$

45. How many seconds are in a solar year (365.24 days)?

Solution:

1 day = 24 hours = 1440 minutes = 86,400 seconds

$\Rightarrow$365.24 days = 86,400 sec × 365.24 = 3.156×10^7s

46. Classify the following as chemical or physical property.
    a. The normal color of bromine is red - orange.
    b. Iron rusts in the presence of air and moisture.
    c. Dynamite can explode when it interacts with oxygen.
    d. The density of uranium metal is 19.07 g/mL.
    e. Aluminum metal melts at 660°C.

Solution:

- a, d and e are physical properties.
- b and c are chemical properties.

47. Oxygen freezes to a solid at -218°C. What is this temperature in °F? in K?

Solution:

$$°F = \frac{9}{5}.°C + 32 = \frac{9}{5} \times -218 + 32 = -360.4°F$$

$$K = °C + 273 = -218 + 273 = 55K$$

48. Calculate the number of moles of 1.2×10^{24} HCl molecules.

Solution:

$$n = \frac{\text{given number of molecules}}{6.02 \times 10^{23} \text{molecule/mol}}$$

$$n = \frac{1.2 \times 10^{24} \text{ molecules}}{6.02 \times 10^{23} \text{ molecules/mol}} = 2\text{mol}$$

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49. Calculate the number of hydrogen atoms in 8.5g of ammonia. ($M_{NH_3} = 17g/mol$)

Solution:

$$n_{NH_3} = \frac{m_{NH_3}}{M_{NH_3}} = \frac{8.5g}{17g/mol} = 0.5mol$$

$$\Rightarrow n_H = n_{NH_3} \times \text{subscript of H} = 0.5mol \times 3 = 1.5mol$$

$\therefore$ Number of atoms of hydrogen $= n \times 6.02 \times 10^{23}$ atoms/mol

$= 1.5 \text{ mol} \times 6.02 \times 10^{23}$ atoms/mol

$= 9.03 \times 10^{23}$ atoms

50. Convert the following

a. 0.0712 cm^3 to dm^3

b. $6A°$ to μm

c. 6.72×10^{-7}m to nm

Solution:

a. 71.2dm^3

b. $6 \times 10^{-4}\mu m$

c. 672 nm

51. Hexane, a solvent used for rubber cement, has a specific gravity of 0.668. What is the density of hexane in g/cm^3 and in kg/m^3? Water has a density of $1g/cm^3$.

Solution:

$$\text{Specific gravity} = \text{specific density} = \text{sp.gr} = \frac{\rho_{\text{substancees}}}{\rho_{\text{water}}}$$

$$\text{sp.gr} = \frac{\rho_{\text{hexane}}}{\rho_{\text{water}}}$$

$$\rho_{\text{hexane}} = \text{sp. gr. } \rho_{\text{water}}$$

$$= 0.668 \times 1 \text{ g/cm}^3 = 0.668 \text{ g/cm}^3$$

$$1 \text{ g/cm}^3 = 1000 \text{ kg/m}^3$$

$$\therefore 0.668 \text{ g/cm}^3 = 668 \text{ kg/m}^3$$

52. What is scientific method?

Explanation: Scientific method is a set of guidelines for the practice of science to ensure a certain scientific method.

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Supplementary Exercise

I. Match the quantities stated on the left with their SI unit stated on the right.

	A	B
1	Length	A. square meter (m^2)
2	mass	B. Newton (N)
3	pressure	C. Kelvin (K)
4	Force	D. Pascal (Pa)
5	Temperature	E. Kilogram (kg)
6	Time	F. Ampere (A)
7	Current	G. Cubic meter (m^3)
8	Volume	H. Second (s)
9	Density	I. Kilogram per cubic meter (kg/m^3)
10	Area	J. Meter (m)
		K. Mole (mol)
		L. Unit less

II. Choose the correct answer

11. Which of the following has no unit?
    A. Amount of substance
    B. Luminuous intensity
    C. specific gravity
    D. weight

12. All of the following are chemical changes. EXCEPT
    A. Heating of sugar
    B. decaying of food
    C. reaction of potassium with water
    D. sublimation of iodine to vapour

13. Which of the following is not intensive properties?
    A. colour
    B. Area
    C. Odour
    D. Conductivity

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14. Which of the following branches of chemistry involves separating, identifying and determing the relative amount of a sample of matter?
    A. Analytical chemistry
    B. Inorganic chemistry
    C. Organic chemistry
    D. Physical chemistry
15. Which one is used as a unit of time for long time interval?
    A. second
    B. microsecond
    C. millisecond
    D. hour
16. A fraction 10^{12} represents a prexix
    A. giga
    B. mega
    C. tera
    D. pico
17. Which one of the following is a cause for uncertainty?
    A. The person doing the experiment
    B. The measuring equipment
    C. The environment where the measurement is being made.
    D. All of the above
18. Which of the following measurement can be expressed by exact numbers?
    A. The volume of water in a cylinder
    B. The area of a rectangle
    C. The mass of a piece of calcium metal
    D. The number of microseconds in a minute
19. A concise verbal statement or mathematical equation that summarizes a broad variety of observation is
    A. theory
    B. law
    C. hypothesis
    D. observation
20. Which one of the following is measuring equipment?
    A. Spirit burner
    B. test tube
    C. thermometer
    D. boiling tube
21. Which of the following is not allowed in the chemistry laboratory?
    A. Wearing any type of cloth and shoe
    B. Knowing the hazards of the materials being used.
    C. Reading the labels on the reagent bottles carefully
    D. Using glycerin when inserting glass tubing into rubber stoppers.

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22. Which of the following has three significant figures?

A. 530
C. 0.003
B. 0.7130
D. 4.01

23. A blood sample analysis showed a total of 0.00467g lead. What is the result if it is expressed in micrograms?

A. 0.467μg
C. $4.67 \times 10^6 \mu g$
B. $4.67 \times 10^3 \mu g$
D. $4.67 \times 10^4 \mu g$

24. Which of the following is true?

A. A scientific notation for 0.00021 is 2.1×10^{-4}
B. A substance with density 2.7g/ml floats on water
C. one meter has about the same length as a yard.
D. The prefix kilo means multiplied by 0.001.

25. The number 5.68 ± 0.0001 has _____ significant figures

A. 4
C. 6
B. 5
D. 3

26. Which of the following is derived quantity?

A. force
C. temperature
B. mass
D. time

27. Which of the following has the lowest uncertainty?

A. 6.0
C. 6.000
B. 6.00
D. 6.0000

28. The knowledge of chemistry is not important in one of the following?

A. in agriculture
C. in medicine
B. in geology
D. none

29. Which of the following is not correct?

A. Density has no units.
B. Every measurement has a unit tied to it.
C. Thee Kelvin degree is larger than the Celsius degree
D. Changes of state are physical properties.

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30. The density of carbon in the form of diamond is 3.51g/cm^3. If you have a small diamond with a volume of 0.027cm^3, What is its mass in gram?
    A. 0.095
    B. 2.1×10^{23}
    C. 0.0078
    D. 0.027

31. One atmosphere is not equivalent with
    A. 760mmHg
    B. 760torr
    C. 760cmHg
    D. 101.3kpa

32. Which of the following is caused by poor calibration of instruments?
    A. Random error
    B. Relative average deviation
    C. Average deviation
    D. systematic error

33. The number of significant figures in 2.35×10^{-5}A is
    A. 5
    B. 4
    C. 3
    D. 6

34. One mole is not equal to
    A. 6.02×10^{23} particles
    B. A gram atom
    C. A gram molecule
    D. A formula unit

35. Which one of the following has the largest number of atoms?
    A. 1gH_2
    B. 1gO_2
    C. 1gN_2
    D. 1gF_2

III. Give short answer

36. Calculate the number of mole of hydrogen in 171g of sucrose ($C_{12}H_{22}O_{11}$).

37. State the number of significant figures in the following quantities.
    a. 6.00g of gold
    b. 0.004sec
    c. 100 test tubes
    d. 6.001×10^6mL of water

38. Round off the following numbers to the indicated number of significant figures (sf.)
    a. 8.348 (2sf)
    b. 8.348 (3sf)
    c. 8.345 (3sf)
    d. 8.375 (3sf)
    e. 2.99298 (2 sf)

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39. A volumetric flask has a volume of 250mL and weighs 100g. If it weighs 227g when filled with pure ether, what is the density of ether?

40. Perform the following calculations and express the result to the proper number of significant figures

    a. 21.235 + 3.24 b. 135.24 – 48.1 c. (3.60×273)/1372

41. Convert the following temperatures as indicated

    a. -127°C to K b. 3K to °C c. 212°F to K

42. Find the mass of 1.5mol of calciumhydroxide ($Ca(OH)_2$)

43. The ____ of measurement describes how well repeated measurements of a value agree with one another while ____ is determined by the agreement between the measured quantity and the correct value.

44. A student determined his body temperature to be 38°C. What is the error involved in his measurement, if the normal body temperature is 37°C?

45. Calculate the number of seconds in 10min.

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