Basics of Chemistry

CONCEPT NOTES

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Basics of Chemistry

Section - 1	THE FOUNDATIONS OF CHEMISTRY

Introduction

Chemistry is an integrated subject i.e. it has various facets and dimensions and these various facets and dimensions of chemistry cannot be studied in isolation from one another. For example, one cannot study chemical equilibrium without a sound foundation of thermodynamics. And similarly, studying organic chemistry without studying reaction mechanisms (which are an integral part of chemical kinetics) would not take us too far in our endeavour to study the organic chemistry.

So, the key to studying and developing a sound understanding of chemistry lies in realising this integrated nature of the various aspects of chemistry and planning our study of chemistry accordingly. We can compare chemistry to a necklace of beads in this regard. Just as a thread connects all the beads of a necklace together, certain fundamentals of chemistry make the various aspects of chemistry an integrated subject. Our approach in the **Locus Study Material** for chemistry has been to give this integrated approach a paramount importance and this unit on basic chemistry is a part of our endeavour in that direction.

We intend to understand some basic tenets of chemistry in this unit which will help us understand various aspects of chemistry in an integrated way. Although, the basics covered here are not exhaustive in content, they are quite comprehensive to set the ball rolling on our way to understanding chemistry in an integrated way.

Thereafter, as we move on from unit to unit, this integrated approach would remain as an undercurrent and would be further strengthened.

Content

Although the contents of this unit act as a prelude to the chemistry as a whole, they are particularly helpful in enhancing our understanding of *Physical and Inorganic Chemistry*. But as we have pointed out above, this approach of "connectedness" will be maintained throughout and as a part of this approach, we will provide you with units on ***General Inorganic Chemistry*** and ***General Organic Chemistry*** covering the basic fundamentals that will be used to enhance our understanding of the respective topics of chemistry in the above mentioned "integrated way".

The tree of Chemistry

Chemistry touches almost every aspect of our lives, our culture and our environment. Its scope encompasses the air we breathe, the food we eat, our clothing, our dwellings, transportation, fuel supplies and much more.

In brief, we can say that chemistry is the science that describes matter-its chemical and physical properties, the chemical and physical changes it undergoes, and the energy changes that accompany those processes. (We will describe all these in brief as we move on.)

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Having such a wide scope, quite obviously chemistry is a vast subject. And to simplify the study of such a vast subject a rough division of chemistry is generally followed. This division is explained in this figure:

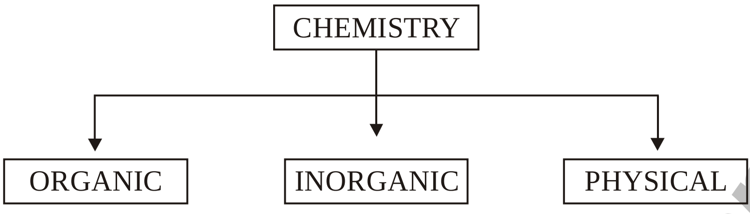

Figure 1: The tree of chemistry

These divisions should not be seen as absolute divisions but as being complementary to each other and overlapping many a times. The basis on which this division is made are as follows:

- Carbon is very versatile in its bonding and behaviour and is a key element in many substances that are essential to life. All living matter contains carbon combined with hydrogen. The number of these compounds is so huge and so versatile, that it becomes imperative to study these as an integrated whole. This integrated whole is what has been termed as **Organic Chemistry**, which is nothing but the chemistry of compounds of carbon and hydrogen.
- The study of compounds that do not contain carbon combined with hydrogen is called **Inorganic Chemistry**.
- **Physical Chemistry** applies the mathematical theories and methods of physics to the properties of matter and to the study of chemical processes and the accompanying energy changes.

Besides these there are various sub-divisions like analytical chemistry and bio chemistry. But for now, the understanding of the above is sufficient and these two terms will be dealt with in respective chapters.

Finally, it should be noted that in the early days of chemistry, its was believed that living matter and inanimate matter were entirely different. The study of the former was termed as organic and of the latter as inorganic. But we now know that many of the compounds found in living matter can be made from non-living or "inorganic" sources. So, now, organic chemistry does not mean study of living matter only. Thus, the terms "organic" and "inorganic" have different meanings than they did originally. Their contemporary meanings are explained above.

Matter and Energy

Lets use this subsection to recapitulate our understanding of Matter and Energy.

Matter is any thing that has mass and occupies space. Mass is a measure of the quantity of matter in sample of any material. The more massive an object is, the more force is required to put it in motion. **Energy** is defined as the capacity to do work or to transfer heat. We are familiar with many forms of energy, including mechanical energy, light energy, electrical energy and heat energy. Light energy from the sun is used by plants as they grow; electrical energy allows us to light a room by flicking a switch; and heat energy cooks our food and warms our homes. Energy can be classified into two principal types: kinetic energy and potential energy.

A detailed discussion of energy will be dealt with in the unit on thermodynamics. As of now it is sufficient to understand what is written above.

Energy is an important concept because all chemical processes are accompanied by energy changes. As some processes occur, energy is released to the surroundings, usually as heat energy. We call such processes **exothermic**. Any combustion (burning) reaction is exothermic.

However, some chemical reactions and physical changes are **endothermic.** i.e. they absorb energy from their surroundings. An example of a physical change that is endothermic is the melting of ice.

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The law of conservation of matter

When we burn a sample of metallic magnesium in the air, the magnesium combines with oxygen from the air to form magnesium oxide, a white powder. This chemical reaction is accompanied by the release of large amounts of heat energy and light energy. When we weigh the product of the reaction, magnesium oxide, we find that it is heavier than the original piece of magnesium. The increase in mass of the solid is due to the combination of oxygen with magnesium to form magnesium oxide. Many experiments have shown that the mass of the magnesium oxide is exactly the sum of the masses of magnesium and oxygen that combined to form it. Similar statements can be made for all chemical reactions. These observations are summarised in the **Law of Conservation of Matter**:

There is no observable change in the quantity of matter during a ***chemical reaction*** *or during a* ***physical change****.*

(please note that a nuclear reaction is not a chemical reaction and the above holds only for chemical reactions and physical changes.)

The law of Conservation of Energy

In exothermic chemical reactions, chemical energy is usually converted into heat energy. Some exothermic processes involve other kinds of energy changes. For example, some liberate light energy without heat, and others produce electrical energy without heat or light. In endothermic reactions, heat energy, light energy, or electrical energy is converted into chemical energy. Although chemical changes always involve energy changes, some energy transformations do not involve chemical changes at all. For example, heat energy may be converted into electrical energy or into mechanical energy without any simultaneous chemical changes. Many experiments have demonstrated that all of the energy involved in any chemical or physical change appears in some form after the change. These observations are summarised in the **Law of Conservation of Energy**:

Energy cannot be created or destroyed in a ***chemical reaction*** *or in a* ***physical change****. It can only be converted from one form to another.*

(We will see some examples of this law in the unit on thermodynamics.)

The law of conservation of matter and energy

Now, note that both the laws above are valid only for a chemical or physical change and not for a nuclear reaction. Actually in nuclear reactions neither matter nor energy are conserved absolutely. With the dawn of the nuclear age in the 1940s, scientists, and then the world, became aware that matter can be converted into energy. In nuclear reactions matter is transformed into energy. The relationship between matter and energy is given by Albert Einstein's now famous equation

$$E = mc^2$$

This equation tells us that the amount of energy released when matter is transformed into energy is the product of the mass of matter transformed and the speed of light squared. Now that the equivalence of matter and energy is recognised, the **Law of Conservation of Matter and Energy** can be stated in a single sentence:

The combined amount of matter and energy in the universe is fixed.

States of Matter

You have studied the three states of matter in earlier classes. Here, we will briefly discuss them. A detailed discussion will be taken up in the units on gaseous state and solid state. Matter can be classified into three states. In the **solid** state, substances are rigid and have definite shapes. Volumes of solids do not vary much with changes

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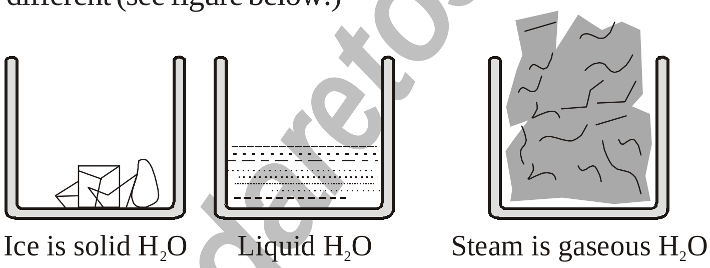
in temperature and pressure. In many solids, called crystalline solids, the individual particles that make up the solid occupy definite positions in the crystal structure. (We will discuss the crystal structure in detail in the solid state.) The strengths of interactions between the individual particles determine how hard and how strong the crystals are. In the **liquid** state, the individual particles are confined to a given volume. A liquid flows and assumes the shape of its container up to the volume of the liquid. Liquids are very hard to compress. **Gases** are much less denser than liquids and solids. They occupy all parts of any vessel in which they are confined. Gases are capable of infinite expansion and are compressed easily. We conclude that they consist primarily of empty space; i.e., the individual particles are quite far apart.

Chemical and physical properties

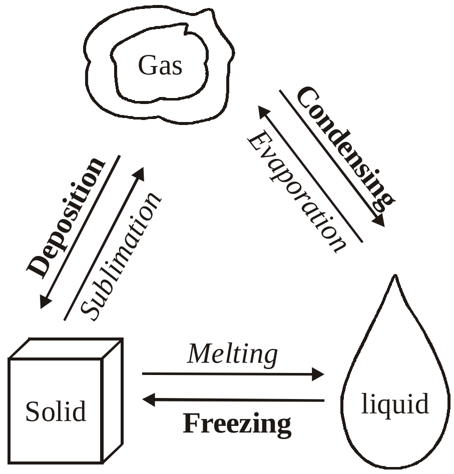
How do we differentiate between two persons so as to who is who? Of course by the height, weight, sex, skin, hair colour and the many subtle features that constitute that person's general appearance. These are the "properties " of that person. Similarly to distinguish among samples of different kinds of matter, we determine and compare their **properties**. We recognise different kinds of matter by their properties, which are broadly classified into chemical properties and physical properties. **Chemical properties** are properties exhibited by matter as it undergoes changes in composition. These properties of substances are related to the kinds of chemical changes that the substances undergo. For instance, we have already described the combination of metallic magnesium with gaseous oxygen to form magnesium oxide, a white powder. A chemical property of magnesium is that it can combine with oxygen, releasing energy in the process. A chemical property of oxygen is that it can combine with magnesium.

All substances also exhibit **physical properties** that can be observed in the absence of any change in composition. Colour, density, hardness, melting point, boiling point, and electrical and thermal conductivities are physical properties. Some physical properties of a substance depend on the conditions, such as temperature and pressure, under which they are measured. For instance, water is a solid(ice) at low temperatures but is a liquid at higher temperatures. At still higher temperatures, it is a gas(steam). As water is converted from one state to another, its composition is constant. Its chemical properties change very little. On the other hand, the physical properties of ice, liquid water, and steam are different (see figure below:)



Property			
Rigidity	Rigid	Flows and assumes shape of container	Fills the container completely
Expansion on heating	Slight	Slight	Expands infinitely
Compressibility	Slight	Slight	Easily compressed

(a)



(b)

Figure: 2 **(a)** A comparison of some properties of the three states of matter (for water) **(b)** Physical changes that occur among the three states of matter. **Sublimation** is the conversion of a solid directly to a gas without passing through the liquid state; the reverse of that process is called **deposition**. The changes shown in italics are endothermic (absorb heat); those shown in bold are exothermic(release heat).

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Properties of matter can be further classified according to whether or not they depend on the amount of substance present. The volume and the mass of a sample depend on, and are directly proportional to, the amount of matter in that sample. Such properties, which depend on the amount of material examined, are called **extensive properties**. By contrast, the colour and the melting point of a substance are the same for a small sample and for a large one. Properties such as these, which are independent of the amount of material examined, are called **intensive properties**. All chemical properties are intensive properties. We will discuss about various intensive and extensive properties in much more detail in the unit on thermodynamics.

Chemical and physical changes

We described the reaction of magnesium as it burns in the oxygen of the air earlier in this chapter. This reaction is a **chemical change**. In any chemical change,(1) one or more substances are used up (at least partially), (2) one or more new substances are formed, and (3) energy is absorbed or released. As substances undergo chemical changes they demonstrate their chemical properties. A **physical change**, on the other hand, occurs with *no change in chemical composition*. Physical properties are usually altered significantly as matter undergoes physical changes (Figure 2b).

Energy is always released or absorbed when chemical or physical changes occur. Energy is required to melt ice, and energy is required to boil water. Conversely, the condensation of steam to form liquid water always liberates energy, as does the freezing of liquid water to form ice. At a pressure of one atmosphere, ice always melts at the same temperature (0ºC) and pure water always boils at the same temperature (100ºC).

(Please note that a pressure of 1 atmosphere is the average atmospheric pressure at sea level.)

Mixtures, substances, compounds, and elements

Mixtures are combinations of two or more pure substances in which each substance retains its own composition and properties. Almost every sample of matter that we ordinarily encounter is a mixture. The most easily recognised type of mixture is one in which different portions of the sample have recongisably different properties. Such a mixture, which is not uniform throughout, is called **heterogeneous**. Examples include mixtures of salt and charcoal (in which two components with different colours can be distinguished readily from one another by sight), foggy air (which includes a suspended mist of water droplets), and vegetable soup. Another kind of mixture has uniform properties throughout; such a mixture is described as a **homogeneous** mixture and is also called a **solution**. Examples include saltwater; some **alloys**, which are homogeneous mixtures of metals in the solid state; and air (free of particulate matter or mists). Air is a mixture of gases. It is mainly nitrogen, oxygen, argon, carbon dioxide, and water vapour. There are only trace amounts of other substances in the atmosphere.

An important characteristic of all mixtures is that they can have variable composition. (For instance, we can make an infinite number of different mixtures of salt and sugar by varying the relative amounts of the two components used). **Consequently, repeating the same experiment on mixtures from different sources may give different results.** Mixtures can be separated by physical means because each component retains its properties (Figure 3). For example, a mixture of salt and water can be separated by evaporating the water and leaving the solid salt behind. To separate a mixture of sand and salt, we could treat it with water to dissolve the salt, collect the sand by filtration, and then evaporate the water to reclaim the solid salt. Very fine iron powder can be mixed with powdered sulphur to give what appears to the naked eye to be homogeneous mixture of the two. However, separation of the components of this mixture is easy. The iron may be removed by a magnet, or the sulfur may be dissolved in carbon disulfide, which does not dissolve iron.

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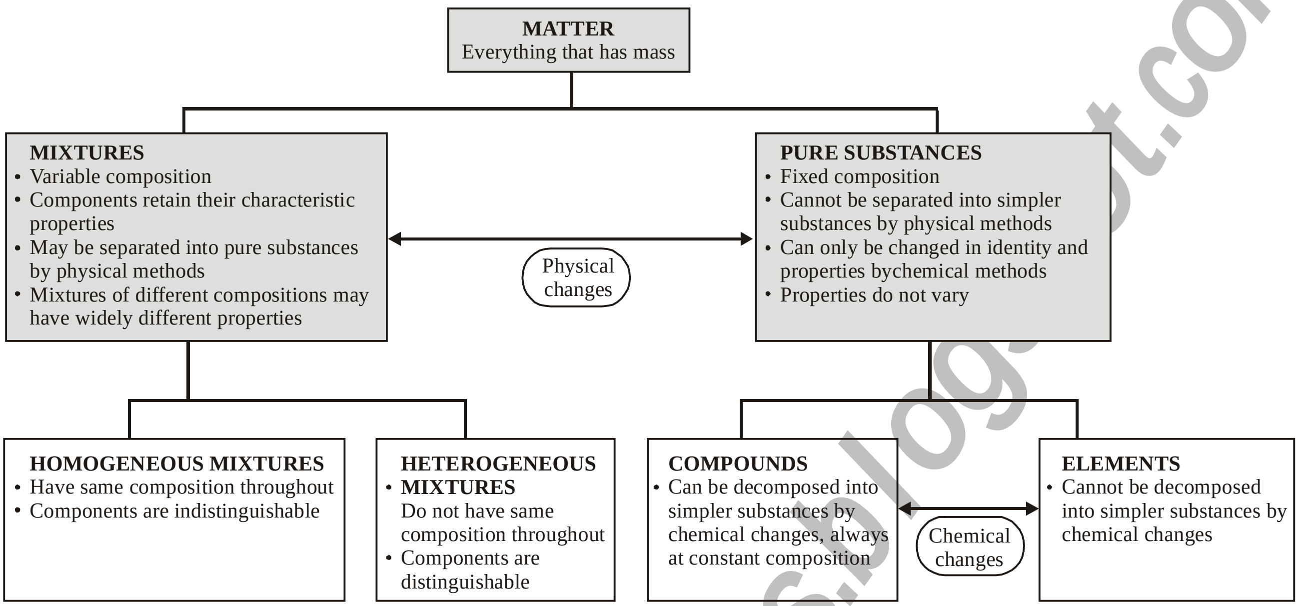

Figure 3: A scheme for classification of matter. Arrows indicate the general means by which matter can be separated

Now, imagine that we have a sample of muddy river water (a heterogeneous mixture). We might first separate the suspended dirt from the liquid by filtration. Then we could remove the dissolved air by warming the water. Dissolved solids might be removed by cooling the sample until some of it freezes, pouring off the liquid, and then melting the ice. Other dissolved components might be separated by distillation or other methods. Eventually, we would obtain a sample of pure water that could not be further separated by any physical separation methods. No matter what the original source of the impure water –the Bay of Bengal, the Ganga river, a can of tomato juice, and so on– water samples obtained by purification all have identical composition and under identical conditions, they all have identical properties. Any such sample is called a substance, or sometimes a pure substance.

*A **substance** cannot be further broken down or purified by physical means.*

Now, suppose we decompose some water by passing electricity through it. (This process of bringing about a decomposition using electricity is called electrolysis.) We find that the water is converted to two simple substances, hydrogen and oxygen, more significantly, hydrogen and oxygen are always present in the same ratio by mass, 11.1 % to 88.9%. These observations allow us to identify water as a compound.

*A **compound** is a substance that can be decomposed by chemical means into simpler substances, always in the same ratio by mass.*

As we continue this process, starting with any substance, we eventually reach a stage at which the new substances formed cannot be further broken down by chemical means. The substances at the end of this chain are called elements.

*An **element** is a substance that cannot be decomposed into simpler substances by chemical changes.*

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For instance, neither of the two gases obtained by the electrolysis of water– hydrogen and oxygen –can be further decomposed; so they are elements.

An another illustration (see Figure 4), pure calcium carbonate (a white solid present in limestone and seashells) can be broken down by heating to give another white solid (call it A) and a gas (call it B) in the mass ratio 56:44. This observation tells us that calcium carbonate is a compound. The white solid A obtained from calcium carbonate can be further broken down into a solid and a gas in a definite ratio by mass, 71.5:28.5. But neither of these can be further decomposed, so they must be elements. The gas is identical to the oxygen obtained from the electrolysis of water; the solid is a metallic element called calcium. Similarly, the gas B, originally obtained from calcium carbonate, can be decomposed into two elements, carbon and oxygen in fixed mass ratio, 27.3:72.7. This sequence illustrates that a compound can be broken apart into simpler substances at fixed mass ratio; those simpler substances may be either elements or simpler compounds.

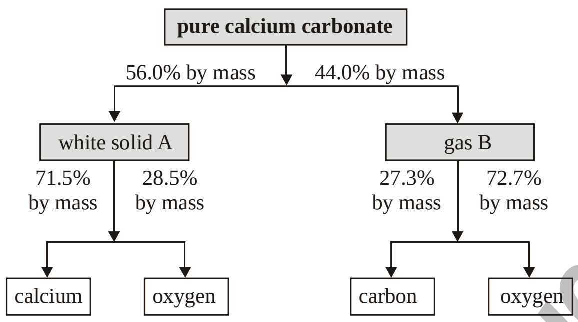

Figure 4: Diagram of the decomposition of calcium carbonate to give a white solid A (56.0% by mass) and a gas B(44.0% by mass). This decomposition into simpler substances at fixed ratio proves that calcium carbonate is a compound. The white solid A further decomposes to give the elements calcium (71.5% by mass) and oxygen (28.5% by mass). This proves that the white solid A is a compound; it is known as calcium oxide. The gas B also can be broken down to give the elements carbon (27.3% by mass) and oxygen (72.7% by mass). This establishes that gas B is a compound; it is known as carbon dioxide.

Further, we may say that *a compound is a pure substance consisting of two or more different elements in a fixed ratio*. Water is 11.1% hydrogen and 88.9% oxygen by mass. Similarly, carbon dioxide is 27.3% carbon and 72.7% oxygen by mass, and calcium oxide (the white solid A above) is 71.5% calcium and 28.5% oxygen by mass. We could also combine the numbers in the previous paragraph to show that calcium carbonate is 40.1% calcium, 12.0% carbon, and 47.9% oxygen by mass. Observations such as these on innumerable pure compounds led to the statement of the **Law of Definite Proportions** (also known as the **Law of Constant Composition**):

Different samples of any pure compound contain the same elements in the same proportions by mass.

The physical and chemical properties of a compound are different from the properties of its constituent elements. Sodium chloride is a white solid that we ordinarily use as table salt. This compound is produced by the combination of the element sodium and the element chlorine.

Recall that elements are substances that cannot be decomposed into simpler substances by chemical changes. Nitrogen, silver, aluminum, copper, gold, and sulphur are other examples of elements.

We use a set of **symbols** to represent the elements. These symbols can be written more quickly than names, and they occupy less space. The symbols for the first 103 elements consist of either a capital letter *or* a capital letter and a lowercase letter, such as C (carbon) or Ca (calcium). Symbols for elements beyond number 103 consist of three letters. A list of the known elements and their symbols is given in the unit on periodicity.

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A short list of symbols of common elements is given in Table 1. Learning this list will be helpful. Many symbols consist of the first one or two letters of the element's English name. Some are derived from the element's Latin name (indicated in parentheses in Table-1 and one, W for tungsten, is from the German *Wolfram*. Names and symbols for additional elements should be learned as they are encountered.

Most of the earth's crust is made up of a relatively small number of elements. Only 10 of the 88 naturally occuring elements make up more than 99% by mass of the earth's crust, oceans, and atmosphere (Table 2). Oxygen accounts for roughly half. Relatively few elements, approximately one fourth of the naturally occurring ones, occur in nature as free elements. The rest are always found chemically combined with other elements.

Only a very small amount of the matter in the earth's crust, oceans, and atmosphere is involved in living matter. The main element in living matter is carbon, but only a tiny fraction of the carbon in the environment occurs in living organisms. More than a quarter of the total mass of the earth's crust, oceans, and atmosphere is made up of silicon, yet it has almost no biological role.

Table 1:
Some Common Elements and Their Symbols

Symbol	Element	Symbol	Element	Symbol	Element
Ag	Silver *(argentum)*	F	Fluorine	Ni	Nickel
Al	Aluminum	Fe	Iron *(ferrum)*	O	Oxygen
Au	Gold *(aurum)*	H	Hydrogen	P	Phosphorus
B	Boron	He	Helium	Pb	Lead *(plumbum)*
Ba	Barium	Hg	Mercury	Pt	Platinum
Bi	Bismuth	I	Iodine	S	Sulphur
Br	Bromine	K	Potassium*(kalium)*	Sb	Antimony*(stibium)*
C	Carbon	Kr	Krypton	Si	Silicon
Ca	Calcium	Li	Lithium	Sn	Tin *(stannum)*
Cd	Cadmium	Mg	Magnesium	Sr	Strontium
Cl	Chlorine	Mn	Manganese	Ti	Titanium
Co	Cobalt	N	Nitrogen	U	Uranium
Cr	Chromium	Na	Sodium *(natrium)*	W	Tungsten (*Wolfrom*)
Cu	Copper *(cuprum)*	Ne	Neon	Zn	Zinc

Table 2:
Abundance of Elements in the Earth's Crust, Oceans, and Atmosphere

Element	Symbol	% by Mass		Element	Symbol	% by Mass	
Oxygen	O	49.5%	99.2%	Chlorine	Cl	0.19%	0.7%
Silicon	Si	25.7		Phosphorus	P	0.12	
Aluminum	Al	7.5		Manganese	Mn	0.09	
Iron	Fe	4.7		Carbon	C	0.08	
Calcium	Ca	3.4		Sulfur	S	0.06	
Sodium	Na	2.6		Barium	Ba	0.04	
Potassium	K	2.4		Chromium	Cr	0.033	
Magnesium	Mg	1.9		Nitrogen	N	0.030	
Hydrogen	H	0.87		Fluorine	F	0.027	
Titanium	Ti	0.58		Zirconium	Zr	0.023	
				All others combined		~0.1%	

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Measurements in Chemistry

Measurements in the scientific world are usually expressed in the units of the metric system or its modernised successor, the international system of units (SI). [The abbreviation SI comes from the French *le systeme International*.] The SI is based on the seven fundamental units listed in the following table:

Table 3:
The seven Fundamental Units of measurement

Physical property	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	S
Electric Current	ampere	A
Temperature	kelvin	K
Luminous Intensity	candela	cd
Amount of substance	mole	mol

Common Prefixes

A lot many measurements in chemistry are very large or very small numbers. In order to simplify their representation, we use certain prefixes. We enlist the common prefixes used in the SI and metric systems in table 4 below. This table should be memorised.

Table-4:
Common Prefixes Used in the SI and metric systems

Prefix	Abbreviation	Measure	Example
Mega-	M	10^6	1 megameter (Mm) = 1×10^6 m
Kilo-	k	10^3	1 kilometer (km) = 1×10^3 m
Deci-	d	10^{-1}	1 decimeter (dm) = 0.1 m
Centi-	c	10^{-2}	1 centimeter (cm) = 0.01 m
Milli-	m	10^{-3}	1 milligram (mg) = 0.001 g
Micro-	μ(mew)	10^{-6}	1 microgram (mg) = 1×10^{-6} g
Nano-	n	10^{-9}	1 nanogram (ng) = 1×10^{-9} g
Pico-	p	10^{-12}	1 picogram (pg) = 1×10^{-12} g

Volume

One common physical quantity the unit of which, though simple, is confusing to students is that of volume. We discuss it here.

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Volumes are measured in liters or milliliters in the metric system. 1 litre (1 L) is one cubic decimeter (1 dm^3) or 1000 cubic centimeters (1000 cm^3). 1 mL is 1 cm^3. What this means is that if we take a cube, the side of which is 1 cm and pile up 1000 such cubes, then the resulting volume of those 1000 cubes taken together is what is called as 1 litre (1 L).

Unit Factor Method For Calculations in Physical Chemistry

Unit factor method is the method of calculation in which one carries along the units for quantities. We will be following this method for most of the solved problems in **Locus study material** for physical chemistry. You are requested to imbibe this approach when you do your calculations while solving a problem, as it has various advantages attached to it.

The key to using unit factor method is the correct use of conversion factors to change one unit into another. A **conversion factor** is a fraction whose numerator and denominator are the **same** quantity expressed in different units. For example, 2.54 cm and 1 inch are the same length i.e. 2.54 cm = 1 inch. This relationship allows us to write two conversion factors :

$$\frac{2.54\,cm}{1\,inch} \text{ and } \frac{1\,inch}{2.54\,cm}$$

The first of these factors is used when we want to convert inches to centimeters. For example, the length in centimters of an object that is 8.50 inches long is given by :

$$\text{Number of centimeters} = (8.50\,\cancel{inches}) \times \frac{2.54\,cm}{1\,\cancel{inch}} = 21.6\,cm$$

desired unit

Given unit

Note that the units of inches in the denominator of the conversion factor cancel the units of inches in the data we were given (8.50 inches). The centimeters in the numerator of the conversion factor become the units of the final answer.

In general the units multiply and divide as follows :

$$\cancel{\text{Given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit.}$$

If the desired units are not obtained in a calculation, then an error must have been made somewhere. A careful inspection of units often reveals the source of error.

We can use more than one conversion factor in the solution of a problem. For example, suppose we want to know the length in inches of an 8 metre rod. However, the relationship between inches and metres is not given in any of the standard tables. But relationship between cms and inches is given and we also know that 100 cms = 1 m. Thus, we can convert first from meters to centimeters, and then from centimeters to inches :

$$\text{Length in inches} = (\text{length in metres}) \times \begin{pmatrix} \text{factor converting} \\ \text{metre} \rightarrow \text{cm} \end{pmatrix} \times \begin{pmatrix} \text{factor converting} \\ \text{cm} \rightarrow \text{inch} \end{pmatrix}.$$

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The relationship between meters and centimeters gives us the first conversion factor. Because we need to cancel meters, we write meters in the denominator :

$$\frac{100\ cm}{1\ m}$$

We then use the relationship 2.54 cm = 1 inch to write the second conversion factor with the **desired units**, inches, in the numerator :

$$\frac{1\ inch}{2.54\ cm}$$

Thus, we have

$$\text{Number of inches} = (8.00\ m)\left(\frac{100\ cm}{1\ \ m}\right)\left(\frac{1\ inch}{2.54\ cm}\right) = 315\ inches$$

The conversion factors above convert from length to length or, more generally, from one unit of a given measure to another unit of the same measure. We also have conversion factors that convert from one measure to a different one. The density of a substance, for example, can be treated as a conversion factor between mass and volume. Suppose that we want to know the mass in grams of two cubic inches (2.00 inch^3) of gold, which has a density of 19.3 g/cm^3. The density gives us the following factors :

$$\frac{19.3\ grams}{1\ cm^3} \text{ and } \frac{1\ cm^3}{19.3\ grams}$$

Because the answer we want is a mass in grams, we can see that we will use the first of these factors, which has mass in grams in the numerator. To use this factor, however we must first relate cubic centimeters to cubic inches.

We know the factor for converting from inches to centimeters, and the cube of this gives us the desired conversion factor :

$$\left(\frac{2.54\ cm}{1\ inch}\right)^3 = \frac{(2.54)^3\ \ cm^3}{(1)^3\ \ inch^3} = \frac{16.39\ cm^3}{1\ inch^3}.$$

Notice that both the numbers and the units are cubed. Applying our conversion factors, we can now solve the problem:

$$\text{Mass in grams} = (2.00\ \ \cancel{inches}^3)\left(\frac{16.39\ \ \cancel{cm}^3}{1\ \cancel{inch}^3}\right)\left(\frac{19.3\ grams}{1\ \cancel{cm}^3}\right) = 633\ grams.$$

Summary of the unit factor method

In using this method to do our calculations, we will always ask three questions :

1. What data are given in the problem ?

2. What quantity do we wish to obtain in the problem ?

3. What conversion factors do we have available to take us from the given quantity to the desired one ?

Answering these questions properly will leave us with all the necessary data to do our calculation easily and efficiently using the unit factor method .

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Advantages of unit factor method

1. It makes the calculations a one step process, no matter how many conversion of units are involved.

2. The units for the answer will come out of the calculations. If you make an error in arranging factors in the calculations (for example, if you use the wrong formula), this will become apparent because the final units will be nonsense.

Specific Gravity

The specific gravity of a substance is the ratio of its density to the density of water, both at the same temperature.

$$\text{Specific gravity} = \frac{D_{\text{substance}}}{D_{\text{water}}} \quad \text{(D here denotes density)}$$

The density of water is 1.000 g/mL at 3.98°C, the temperature at which the density of water is greatest. However, variations in the density of water with changes in temperature are small enough that we may use 1.00 g/mL up to 25°C without introducing significant errors into our calculations.

- The density of table salt is 2.16 g/mL at 20°C. What is its specific gravity?

Solution: We use the definition of specific gravity given above. The numerator and denominator have the same units, so the result is dimensionless.

$$\text{Specific Gravity} = \frac{D_{\text{salt}}}{D_{\text{water}}} = \frac{2.16 \text{ g/mL}}{1.00 \text{ g/mL}} = 2.16$$

This example also demonstrates that the density and specific gravity of a substance are numerically equal near room temperature if density is expressed in g/mL (g, cm^3).

Generally, we are given specific gravities of certain substances and their percentage by mass in a solution. From this informations the amount of substance present in a given volume of solution can be calculated. We illustrate this in Example 8.

Heat and Temperature

Earlier you learned that heat is one form of energy. You have also learned in earlier classes different forms of energy can be interconverted and that in chemical processes, chemical energy is converted to heat energy or vice versa. The amount of heat a process uses (*endothermic*) or gives off (*exothermic*) can tell us a great deal about that process. For this reason it is important for us to be able to measure intensity of heat.

Temperature measures the intensity of heat, the "hotness" or "coldness" of a body. A piece of metal at 100°C feels hot to the touch, while an ice cube at 0°C feels cold. Why? Because the temperature of the metal is higher, and that of the ice cube lower, than body temperature. **Heat** is a form of energy that *always flows spontaneously from a hotter body to a colder body*—never in the reverse direction:

There are three scales on which the temperature can be measured. These are celsius (or centigrade) (°C) Fahrenheit (°F) and kelvin (K) scales. They are related to each other according to the following equations.

(i) $°F = 1.8° C + 32°$ **(ii)** $K = °C + 273.15°$

Please note that any temperature *change* has the same numerical value whether expressed on the Celsius scale or on the Kelvin scale. For example, a change from 25°C to 59°C represents a change of 34 Celsius degrees. Converting these to the Kelving scale, the same change is expressed as (273 + 25) = 298 K to (59 + 273) = 332 K, or a *change* of 34 kelvins.

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Comparing the Fahrenheit and Celsius scales, we find that the intervals between the same reference points are 180 Fahrenheit degrees and 100 Celsius degrees, respectively. Thus a Fahrenheit degree must be smaller than a Celsius degree. It takes 180 Fahrenheit degrees to cover the same temperature *interval* as 100 Celsius degrees. From this information, we can construct the conversion factors for temperature *changes*:

$$\frac{180°F}{100°C} \; or \; \frac{1.8°F}{1.0°C} \text{ and } \frac{100°C}{180°F} \; or \; \frac{1.0°C}{1.8°F}$$

But the starting points of the two scales are different, so we *cannot convert* a temperature on one scale to a temperature on the other just by multiplying by the conversion factor. In converting from °F to °C, we must subtract 32 Fahrenheit degrees to reach the zero point on the celsius scale

$$\underline{?}°F = \left(x°C \times \frac{1.8°F}{1.0°C} \right) + 32°F \text{ and } \underline{?}°C = \frac{1.0°C}{1.8°F}(x°F - 32°F)$$

We illustrate these conversion in Examples 9.

Heat Transfer and the measurement of heat

Chemical reactions and physical changes occur with either the simultaneous evolution of heat (exothermic process) or the absorption of heat (endothermic processes). The amount of heat transferred in a process is usually expressed in joules or in calories.

The SI unit of energy and work is the **joule (J),** which is defined as 1 $kg\,.\,m^2/s^2$. The kinetic energy (KE) of a body of mass m moving at speed v is given by $\frac{1}{2}mv^2$. A 2-*kg* object moving at one meter per second has $KE = \frac{1}{2}(2\,kg)(1\,m/s)^2 = 1\,kg\,.\,m^2/s^2 = 1$ joule. You may find it more convenient to think in terms of the amount of heat required to raise the temperature of one gram of water from 14.5°C to 15.5°C, which is 4.184 Joules.

One **calorie** is defined as exactly 4.184 joules. We shall do most calculations in joules.

The **specific heat** of a substance is the amount of heat required to raise the temperature of one gram of the substance one degree C (also one kelvin) *with no change in phase*. Changes in phase (physical state) absorb or liberate relatively large amounts of energy. The specific heat of each substance, a physical property, is different for the solid, liquid, and gaseous phases of the substance. For example, the specific heat of ice is 2.09 J/g. °C near 0°C: for liquid water it is 4.18 J/g. °C: and for steam it is 2.03 J/g. °C near 100°C. The specific heat for water is quite high.

$$\text{specific heat} = \frac{\text{(amount of heat in J)}}{\text{(mass of substance in g)(temperature change in °C)}}$$

The **heat capacity** of a body is the amount of heat required to raise its temperature by 1°C. The heat capacity of a body is its mass in grams times its specific heat.

Please study Examples 10 and 11 to understand the concept of specific heat more comprehensively.

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A NOTE ON SOLVED EXAMPLES IN PHYSICAL CHEMISTRY

To bring yourself to the level of IIT-JEE it is imperative that you start thinking with concepts on your own. One of the best ways to develop this kind of thinking is to approach a solved example by only reading and understanding the question first. Then one should think about the approach he/she would take to solve that particular problem. Then attempt the problem by your method. Only then see the solution to the problem and compare it to your own approach. This way you will be entirely involved in your learning process and will be able to see clearly (if you were going wrong), so as to where exactly you were going wrong. It may even be that your solution and the solution given in the material are different but both are correct. In that case you would have ended up finding an alternative solution!

Here in your material for physical chemistry we have gone a step beyond to help you in your endeavour to study "critically" as explained above. A ***critical thinking box*** is put up in almost every solved problem as a bridge between the question and its solution. It contains at times detailed step by step *procedure* for approaching the problem and at times only a hint required to solve the problem. However, although varying in content its focus remains the same i.e., to help you think critically. To make use of it comprehensively it is suggested that you think about the approach you would take to solve the problem. Only then study the *critical thinking box* and compare your approach to the approach given. This kind of study will ensure that any errors in your approach are rectified and also help you to develop your own thinking process. You should then try to attempt the problem yourself and only then, see its detailed solution.

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Example – 1

A material is believed to be a compound. Suppose you have several samples of the material obtained from various places around the world. Comment on what you would expect to find upon observing the melting point and colour for each sample. What would you expect to find upon determining the elemental composition for each sample?

Critical thinking
What is a compound? Does property of 2 samples of the same compound vary? If they do vary, under what conditions can they vary? If they do not vary, then consider what conditions? (Hint: the sample may be pure or impure)

Solution: If the material is a pure compound, all samples should have the same melting point, the same colour, and the same elemental composition. If it is a mixture, there should be a difference in these properties depending on the composition.

Example – 2

You are working in the office of a precious metal buyer. A miner brings you a nugget of metal that he claims is gold. You suspect that the metal is a form of "fool's gold" called marcasite, which is composed of iron and sulphur. In the back of your office, you have a chunk of pure gold. What simple experiments could you perform to decide whether or not the miner's nugget is gold?

Critical thinking
Gold is a very unreactive substance, so comparing physical properties is probably your best option. However, color is a physical property that you cannot rely on in this case to get your answer. What are the other physical properties that you can use to make the distinction clearer?

Solution: One experiment that you could perform is to determine the densities of the metal and the chunk of gold. You could measure the mass of the nugget on a balance and the volume of the nugget by water displacement. Using this information, you could calculate the density of the nugget. Repeat the experiment and calculations for the sample of gold. If the nugget is gold, the two densities should be equal and be 19.3 g/cm^3.
Also, you could determine the melting points of the metal and of the chunk of pure gold. The two melting points should be the same (1338K) if the metal is gold.

Example – 3

Consider the following compounds and their densities.

Substance	Density (g/mL)	Substance	Density (g/mL)
isopropyl alcohol	0.785	toluene	0.866
n-butyl alcohol	0.810	ethylene glycol	1.114

You create a column of the liquids in a glass cylinder with the most dense material on the bottom layer and the least dense on the top. You do not allow the liquids to mix.

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(a) First you drop a plastic bead that has a density of 0.24 g/cm^3 into the column. What do you expect to observe?

(b) Next you drop a different plastic bead that has a volume of 0.043 mL and a mass of 3.92 x 10^{-22} g into the column. What would you expect to observe in this case?

(c) You drop another bead into the column and observe that it makes it all the way to the bottom of the column. What can you conclude about the density of this bead?

Critical thinking
A cock floats on water because its density is less than the water. Apply this same fundamental in the various cases of this problem.

Solution: **(a)** Since the bead is less dense than any of the liquids in the container, the bead will float on top of all the liquids.

(b) First, determine the density of the plastic bead. Since density is mass divided by volume, you get

$$d = \frac{m}{V} = \frac{3.92\times10^{-2}\,g}{0.043\,mL} = 0.911\,g/mL = 0.92\,g/mL$$

Thus, the glass bead will pass through the top three layers and float on the ethylene glycol layer, which is more dense.

(c) Since the bead sinks all the way to the bottom, it must be more dense than 1.114 g/mL.

Example – 4

Express 1.47 miles in inches

Critical thinking
First we write down the units of what we wish to know. Then we set it equal to whatever we are given:

no. of inches = 1.47 miles

Then we choose unit factors to convert the given units (miles) to the desired units (inches):

$$miles \rightarrow feet \rightarrow inches$$

Solution: $$\text{no. of inches} = 1.47\ \cancel{\text{miles}} \times \frac{5280\ \cancel{\text{ft}}}{1\ \cancel{\text{mile}}} \times \frac{12\,\text{inches}}{1\,\text{ft}} = 9.31\times10^4\ \text{inches}$$

Note that both miles and feet cancel, leaving only inches, the desired unit. Thus, there is no ambiguity as to how the unit factors should be written.

(Note: Conversion factor is also called unit factor)

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Example – 5

One liter is exactly 1000 cubic centimeters. How many cubic inches are there in 1000 cubic centimeters?

Critical thinking

We would multiply by the unit factor $\frac{1\,inch}{2.54\,cm}$ *to convert cm to inch. Here we require the cube of this unit factor.*

Solution: $no.\ of\ (inch)^3 = 1000\ cm^3 \times \left(\frac{1\,inch}{2.54\,cm}\right)^3 = 1000\ cm^3 \times \frac{1\,(inch)^3}{16.4\,cm^3} = 61.0\,(inch)^3$

Example – 6

A common unit of energy is the erg. Convert 3.74×10^{-2} erg to the SI units of energy, joules and kilojoules. One erg is exactly 1×10^{-7} joule.

Critical thinking

The definition that relates ergs and joules is used to generate the needed unit factor. The second conversion uses a unit factor that is based on the definition of the prefix kilo-.

Solution: $no.\ of\ J = 3.74\times10^{-2}\ erg \times \frac{1\times10^{-7}\ J}{1\,erg} = 3.74\times10^{-9}\ J$

$no.\ of\ kJ = 3.74\times10^{-9}\ j \times \frac{1\ kJ}{1000\ J} = 3.74\times10^{-12}\,kJ$

Example – 7

A 47.3 ml sample of ethyl alcohol(ethanol) has mass of 37.32 g. What is its density? If 103 g of ethanol is needed for a chemical reaction, what volume of liquid would you use?

Critical thinking

We use the definition of density i.e. $Density = \frac{mass}{volume}$*. Thereafter we are given the mass, m, of a sample of ethanol. So, we know values of D and m in the relationship*

$$D = \frac{m}{v}.$$

We rearrange this relationship to solve for V, put in the known values, and carry out the calculation. Alternatively, we can use the unit factor method to solve the problem.

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Solution: $D = \frac{m}{v} = \frac{37.32\,g}{47.3\,ml} = 0.789 \;\; g/mL \cdot$

The density of ethanol is 0.789 g/mL

$$\text{now,} \quad D = \frac{m}{V}$$

$$\Rightarrow \quad V = \frac{m}{D} = \frac{103\,g}{0.789\,g/mL} = 130 \;\; mL$$

Alternatively,

$$\text{no. of mL} = 103 \times \frac{1\,mL}{0.789\,g} = 130\,mL.$$

Example – 8

Battery acid is 40.0 % sulfuric acid H_2SO_4, and 60.0% water by mass. Its specific gravity is 1.31. Calculate the mass of pure H_2SO_4 in 100.0 mL of battery acid.

Critical thinking

The percentages are given on a mass basis, so we must first convert the 100.0 mL of acid solution to mass. To do this, we need a value for the density. We have demonstrated that density and specific gravity are numerically equal at 20ºC because the density of water is 1.00 g/mL. We can use the density as a unit factor to convert the given volume of solution to mass of solution. Then we use the percentage by mass to convert the mass of solution to mass of acid.

Solution: From the given value for specific gravity, we may write

density = 1.31 g/mL

The solution is 40.0% H_2SO_4 and 60.0 % H_2O by mass. From this information we may construct the desired unit factor:

$$\frac{40.0\,g\;H_2SO_4}{100\,g\;\text{soln}} \longrightarrow \begin{array}{l}\text{because } 100\,g \text{ of solution} \\ \text{contains } 40.0\,g \text{ of } H_2SO_4\end{array}$$

We can now solve the problem:

$$\text{grams of } H_2SO_4 = 100.0 \text{ mL soln} \times \frac{1.31\,g\,\text{soln}}{1\,ml\,\text{soln}} \times \frac{40.0g\;H_2SO_4}{100\,g\,\text{soln}} = 52.4 \text{ g } H_2SO_4$$

Example – 9

When the absolute temperature is 400 K, what is the Fahrenheit temperature?

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Critical thinking
We first use the relationship ºC = K – 273º to convert from kelvins to degrees Celsius; then we carry out the further conversion from degrees Celsius to degrees Fahrenheit.

Solution: $\text{no. of } ^\circ\text{C} = (400\text{K} - 273\text{K})\dfrac{1.0^\circ\text{C}}{1.0\,\text{K}} = 127^\circ\text{C}$

$$\text{no. of } ^\circ\text{F} = \left(127^\circ\text{C} \times \frac{1.8^\circ\text{F}}{1.0^\circ\text{C}}\right) + 32^\circ\text{F} = 261^\circ\text{F}$$

Example – 10

How much heat, in joules, is required to raise the temperature of 205 g of water from 21.2ºC to 91.4º C?

Critical thinking
The specific heat of a substance is the amount of heat required to raise the temperature of 1 g of substance by 1ºC.

$$\textit{specific heat} = \frac{(\textit{amount of heat in J})}{(\textit{mass of substance in g})(\textit{temperature change in}^\circ C)}$$

We can rearrange the equation so that

(amount of heat) = (mass of substance) (specific heat) (temperature change)

Alternatively, we can use the unit factor approach.

Solution: amount of heat = (205 g) (4.18 J/g •ºC) (70.2ºC) = 6.02 × 10^4 J
By the unit factor approach.

$$\text{amount of heat} = (205\,\text{g})\left(\frac{4.18\,\text{J}}{1\text{g}\cdot{}^\circ\text{C}}\right)(70.2^\circ\text{C}) = 6.02\times10^4\,\text{J} \;\text{ or }\; 60.2\text{ kJ}$$

All units except joules cancel. To cool 205 g of water from 91.4ºC to 21.2ºC, it would be necessary to remove exactly the same amount of heat, 60.2 kJ.

Example – 11

How much heat, in calories, kilocalories, joules, and kilojoules, is required to raise the temperature of 205 g of iron from 294.2 K to 364.4K? The specific heat of iron is 0.106 cal/g.ºC, or 0.444 J/g . K.

Critical thinking
First we recall that a temperature ***change*** *expressed in kelvins has the same numerical value expressed in degrees Celsius. Remembering that specific heat is in terms of temperature change, we can write the specific heat of iron as 0.106 cal/g.K or 0.444 J/g.K. Then we can solve this problem with the temperature change expressed in kelvins, and avoid the work of converting temperature to ºC.*

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Solution: temperature change = 364.4 K – 294.2 K = 70.2 K

no. of cal = (205 g) (0.106 cal/g•K) (70.2 K) = 1.52×10^3 cal or 1.52 kcal

no. of J = (205 g) (0.444 J/g•K) (70.2 K) = 6.39×10^3 J or 6.39 kJ

Note: The specific heat of iron is much smaller than the specific heat of water:

$$\frac{\text{specific heat of iron}}{\text{specific heat of water}} = \frac{0.444\ \text{J/g}\cdot\text{K}}{4.18\ \text{J/g}\cdot\text{K}} = 0.106$$

As a result, the amount of heat required to raise the temperature of 205 g of iron by 70.2 K (70.2ºC) is less than that required to do the same for 205 g of water by the same ratio.

$$\frac{\text{amount heat for iron}}{\text{amount heat for water}} = \frac{6.39\,\text{kJ}}{60.2\,\text{kJ}} = 0.106$$

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TRY YOURSELF - I

Q. 1 When H_2 and O_2 is mixed in proportion 2 :1 by volume at constant temperature and pressure, what do we get, a compound or mixture, if

(a) The final composition shows properties of O_2.

(b) The final composition doesn't show properties of O_2.

Q. 2 The specific gravity of silver is 10.5 (a) What is the volume, in cm^3, of an ingot of silver with mass 0.765 kg? (b) If this sample of silver is a cube, how long is each edge in cm? (c) How long is the edge of this cube in inches?

Q. 3 Vinegar has density of 1.0056 g/cm^3. What is the mass of two liters of vinegar?

You are requested to attempt the following basic problems (*Q. 4 to Q. 6*) *to master the unit factor method:*

Q. 4 The density of air at ordinary atmospheric pressure and 25°C is 1.19 grams/litre. What is the mass, in kilograms, of the air in a room that measures $12.5 \times 15.5 \times 8.0$ feet ?

Q. 5 A certain printed page has an average of 25 words per square inch of paper. The average length of the words is 5.3 letters. What is the average number of letters per square centimeter of paper ?

Q. 6 The total amount of fresh water on earth is estimated to be $3.73 \times 10^8\ km^3$. What is this volume is cubic meters ? In liters ?

Q. 7 The specific heat of aluminium is 0.895 J/g.ºC. Calculate the amount of heat required to raise the temperature of 22.1 grams of aluminum from 27.0ºC to 44.3ºC.

Q. 8 The lethal dose of potassium cyanide(KCN) taken orally is 1.6 milligrams per kilogram of body weight. Calculate the lethal dose of potassium cyanide taken orally by a 165-pound person.

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TRY YOURSELF - I

ANSWERS

Ans. 1	(a) mixture	(b) compound	
Ans. 2	(a) 72.9 cm^3	(b) 4.18 cm	(c) 1.65 inches
Ans. 3	2.0112 kg		
Ans. 4	52 kilograms.		
Ans. 5	21 letters/cm^2		
Ans. 6	$3.73\times10^{17}\ m^3$; 3.73×10^{20} *liters.*		
Ans. 7	342 Joules		
Ans. 8	120 mg KCN		

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