Chemical Reactions
and
Equations
1
CHAPTER
C
onsider the following situations of daily life and think what happens
when –
nmilk is left at room temperature during summers.
nan iron tawa/pan/nail is left exposed to humid atmosphere.
ngrapes get fermented.
nfood is cooked.
nfood gets digested in our body.
nwe respire.
In all the above situations, the nature and the identity of the initial
substance have somewhat changed. We have already learnt about physical
and chemical changes of matter in our previous classes. Whenever a chemical
change occurs, we can say that a chemical reaction has taken place.
You may perhaps be wondering as to what is actually meant by a
chemical reaction. How do we come to know that a chemical reaction
has taken place? Let us perform some activities to find the answer to
these questions.
Figure 1.1
Burning of a magnesium ribbon in air and collection of magnesium
oxide in a watch-glass
Activity 1.1Activity 1.1
Activity 1.1Activity 1.1
Activity 1.1
CAUTION: This Activity needs
the teacher’s assistance. It
would be better if students
wear suitable eyeglasses.
nClean a magnesium ribbon
about 3-4 cm long by rubbing
it with sandpaper.
nHold it with a pair of tongs.
Burn it using a spirit lamp or
burner and collect the ash so
formed in a watch-glass as
shown in Fig. 1.1. Burn the
magnesium ribbon keeping it
away as far as possible from
your eyes.
nWhat do you observe?
“Facts are not science — as the dictionary is not literature.”
Martin H. Fischer
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Activity 1.2Activity 1.2
Activity 1.2Activity 1.2
Activity 1.2
Figure 1.2Figure 1.2
Figure 1.2Figure 1.2
Figure 1.2
Formation of hydrogen
gas by the action of
dilute sulphuric acid on
zinc
From the above three activities, we can say that any of
the following observations helps us to determine whether
a chemical reaction has taken place –
nchange in state
nchange in colour
nevolution of a gas
nchange in temperature.
As we observe the changes around us, we can see
that there is a large variety of chemical reactions taking
place around us. We will study about the various types
of chemical reactions and their symbolic representation
in this Chapter.
Activity 1.3Activity 1.3
Activity 1.3Activity 1.3
Activity 1.3
nTake a few zinc granules in a conical flask or a test tube.
nAdd dilute hydrochloric acid or sulphuric acid to this
(Fig. 1.2).
CAUTION: Handle the acid with care.
nDo you observe anything happening around the zinc
granules?
nTouch the conical flask or test tube. Is there any change in
its temperature?
nTake lead nitrate
solution in a test
tube.
nAdd potassium
iodide solution
to this.
nWhat do you
observe?
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Activity 1.1 can be described as – when a magnesium ribbon is burnt in
oxygen, it gets converted to magnesium oxide. This description of a
chemical reaction in a sentence form is quite long. It can be written in a
shorter form. The simplest way to do this is to write it in the form of a
word-equation.
The word-equation for the above reaction would be –
Magnesium + Oxygen→Magnesium oxide(1.1)
(Reactants)(Product)
The substances that undergo chemical change in the reaction (1.1),
magnesium and oxygen, are the reactants. The new substance is
magnesium oxide, formed during the reaction, as a product.
A word-equation shows change of reactants to products through an
arrow placed between them. The reactants are written on the left-hand
side (LHS) with a plus sign (+) between them. Similarly, products are
written on the right-hand side (RHS) with a plus sign (+) between them.
The arrowhead points towards the products, and shows the direction of
the reaction.
You must have observed that magnesium ribbon burns with a
dazzling white flame and changes into a white powder. This powder is
magnesium oxide. It is formed due to the reaction between magnesium
and oxygen present in the air.
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Chemical Reactions and Equations
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1.1.1 Writing a Chemical Equation
Is there any other shorter way for representing chemical equations?
Chemical equations can be made more concise and useful if we use
chemical formulae instead of words. A chemical equation represents a
chemical reaction. If you recall formulae of magnesium, oxygen and
magnesium oxide, the above word-equation can be written as –
Mg + O
2
→ MgO(1.2)
Count and compare the number of atoms of each element on the
LHS and RHS of the arrow. Is the number of atoms of each element the
same on both the sides? If yes, then the equation is balanced. If not,
then the equation is unbalanced because the mass is not the same on
both sides of the equation. Such a chemical equation is a skeletal
chemical equation for a reaction. Equation (1.2) is a skeletal chemical
equation for the burning of magnesium in air.
1.1.2 Balanced Chemical Equations
Recall the law of conservation of mass that you studied in Class IX; mass
can neither be created nor destroyed in a chemical reaction. That is, the
total mass of the elements present in the products of a chemical reaction
has to be equal to the total mass of the elements present in the reactants.
In other words, the number of atoms of each element remains the
same, before and after a chemical reaction. Hence, we need to balance a
skeletal chemical equation. Is the chemical Eq. (1.2) balanced? Let us
learn about balancing a chemical equation step by step.
The word-equation for Activity 1.3 may be represented as –
Zinc + Sulphuric acid → Zinc sulphate + Hydrogen
The above word-equation may be represented by the following
chemical equation –
Zn + H
2
SO
4
→ ZnSO
4
+ H
2
(1.3)
Let us examine the number of atoms of different elements on both
sides of the arrow.
ElementNumber of atoms inNumber of atoms
reactants (LHS)in products (RHS)
Zn11
H22
S11
O44
As the number of atoms of each element is the same on both sides of
the arrow, Eq. (1.3) is a balanced chemical equation.
Let us try to balance the following chemical equation –
Fe + H
2
O → Fe
3
O
4
+ H
2
(1.4)
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Step I: To balance a chemical equation, first draw boxes around each
formula. Do not change anything inside the boxes while balancing the
equation.
Fe + H
2
O → Fe
3
O
4
+ H
2
(1.5)
Step II: List the number of atoms of different elements present in the
unbalanced equation (1.5).
ElementNumber of atomsNumber of atoms
in reactants (LHS)in products (RHS)
Fe13
H22
O14
To equalise the number of atoms, it must be remembered that we
cannot alter the formulae of the compounds or elements involved in the
reactions. For example, to balance oxygen atoms we can put coefficient
‘4’ as 4 H
2
O and not H
2
O
4
or (H
2
O)
4
. Now the partly balanced equation
becomes–
Fe + 4 H
2
O → Fe
3
O
4
+ H
2
Step IV: Fe and H atoms are still not balanced. Pick any of these elements
to proceed further. Let us balance hydrogen atoms in the partly balanced
equation.
To equalise the number of H atoms, make the number of molecules
of hydrogen as four on the RHS.
Step III: It is often convenient to start balancing with the compound
that contains the maximum number of atoms. It may be a reactant or a
product. In that compound, select the element which has the maximum
number of atoms. Using these criteria, we select Fe
3
O
4
and the element
oxygen in it. There are four oxygen atoms on the RHS and only one on
the LHS.
To balance the oxygen atoms –
The equation would be –
Fe + 4 H
2
O → Fe
3
O
4
+ 4 H
2
Atoms ofIn reactantsIn products
oxygen
(i)Initial1 (in H
2
O)4 (in Fe
3
O
4
)
(ii)To balance1×44
Atoms ofIn reactantsIn products
hydrogen
(i)Initial8 (in 4 H
2
O)2 (in H
2
)
(ii)To balance82 × 4
(1.6)
(partly balanced equation)
(1.7)
(partly balanced equation)
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To equalise Fe, we take three atoms of Fe on the LHS.
3 Fe + 4 H
2
O → Fe
3
O
4
+ 4 H
2
(1.8)
Step VI: Finally, to check the correctness of the balanced equation, we
count atoms of each element on both sides of the equation.
3Fe + 4H
2
O → Fe
3
O
4
+ 4H
2
The numbers of atoms of elements on both sides of Eq. (1.9) are
equal. This equation is now balanced. This method of balancing chemical
equations is called hit-and-trial method as we make trials to balance
the equation by using the smallest whole number coefficient.
Step VII:
Writing Symbols of Physical StatesWriting Symbols of Physical States
Writing Symbols of Physical StatesWriting Symbols of Physical States
Writing Symbols of Physical States Carefully examine
the above balanced Eq. (1.9). Does this equation tell us anything about
the physical state of each reactant and product? No information has
been given in this equation about their physical states.
To make a chemical equation more informative, the physical states
of the reactants and products are mentioned along with their chemical
formulae. The gaseous, liquid, aqueous and solid states of reactants
and products are represented by the notations (g), (l), (aq) and (s),
respectively. The word aqueous (aq) is written if the reactant or product
is present as a solution in water.
The balanced Eq. (1.9) becomes
3Fe(s) + 4H
2
O(g) → Fe
3
O
4
(s) + 4H
2
(g)(1.10)
Note that the symbol (g) is used with H
2
O to indicate that in this
reaction water is used in the form of steam.
Usually physical states are not included in a chemical equation unless
it is necessary to specify them.
Sometimes the reaction conditions, such as temperature, pressure,
catalyst, etc., for the reaction are indicated above and/or below the arrow
in the equation. For example –
CO(g) + 2H (g)
2
340
atm
CH OH(l)
3
→
(1.11)
6CO (aq) 12H O(l)C H O (aq) 6O
226 12 6
++
Sunlight
Chlorophyll
→
222
(aq) 6H O(l)+
(1.12)
(Glucose)
Using these steps, can you balance Eq. (1.2) given in the text earlier?
Step V: Examine the above equation and pick up the third element
which is not balanced. You find that only one element is left to be
balanced, that is, iron.
Atoms ofIn reactantsIn products
iron
(i)Initial1 (in Fe)3 (in Fe
3
O
4
)
(ii)To balance1×33
(1.9)
(balanced equation)
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1.2 TYPES OF CHEMIC1.2 TYPES OF CHEMIC
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We have learnt in Class IX that during a chemical reaction atoms of one
element do not change into those of another element. Nor do atoms
disappear from the mixture or appear from elsewhere. Actually, chemical
reactions involve the breaking and making of bonds between atoms to
produce new substances. You will study about types of bonds formed
between atoms in Chapters 3 and 4.
1.2.1 Combination Reaction
Activity 1.4Activity 1.4
Activity 1.4Activity 1.4
Activity 1.4
nTake a small amount of calcium oxide
or quick lime in a beaker.
nSlowly add water to this.
nTouch the beaker as shown in Fig. 1.3.
nDo you feel any change in temperature?
Figure 1.3Figure 1.3
Figure 1.3Figure 1.3
Figure 1.3
Formation of slaked
lime by the reaction of
calcium oxide with
water
Calcium oxide reacts vigorously with water to produce slaked lime
(calcium hydroxide) releasing a large amount of heat.
CaO(s)+H
2
O(l) → Ca(OH)
2
(aq) + Heat(1.13)
(Quick lime) (Slaked lime)
In this reaction, calcium oxide and water combine to form a single
product, calcium hydroxide. Such a reaction in which a single product
is formed from two or more reactants is known as a combination reaction.
QUESTIONS
1.Why should a magnesium ribbon be cleaned before burning in air?
2.Write the balanced equation for the following chemical reactions.
(i)Hydrogen + Chlorine → Hydrogen chloride
(ii)Barium chloride + Aluminium sulphate → Barium sulphate +
Aluminium chloride
(iii)Sodium + Water → Sodium hydroxide + Hydrogen
3.Write a balanced chemical equation with state symbols for the
following reactions.
(i)Solutions of barium chloride and sodium sulphate in water react
to give insoluble barium sulphate and the solution of sodium
chloride.
(ii)Sodium hydroxide solution (in water) reacts with hydrochloric
acid solution (in water) to produce sodium chloride solution and
water.
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Let us discuss some more examples of combination reactions.
(i)Burning of coal
C(s) + O
2
(g) → CO
2
(g)(1.15)
(ii)Formation of water from H
2
(g) and O
2
(g)
2H
2
(g) + O
2
(g) → 2H
2
O(l)(1.16)
In simple language we can say that when two or more substances
(elements or compounds) combine to form a single product, the reactions
are called combination reactions.
In Activity 1.4, we also observed that a large amount of heat is evolved.
This makes the reaction mixture warm. Reactions in which heat is
released along with the formation of products are called exothermic
chemical reactions.
Other examples of exothermic reactions are –
(i)Burning of natural gas
CH
4
(g) + 2O
2
(g) → CO
2
(g) + 2H
2
O (g)(1.17)
(ii)Do you know that respiration is an exothermic process?
We all know that we need energy to stay alive. We get this energy
from the food we eat. During digestion, food is broken down into simpler
substances. For example, rice, potatoes and bread contain
carbohydrates. These carbohydrates are broken down to form glucose.
This glucose combines with oxygen in the cells of our body and provides
energy. The special name of this reaction is respiration, the process of
which you will study in Chapter 6.
C
6
H
12
O
6
(aq) + 6O
2
(aq) → 6CO
2
(aq) + 6H
2
O(l) + energy(1.18)
(Glucose)
(iii)The decomposition of vegetable matter into compost is also an
example of an exothermic reaction.
Identify the type of the reaction taking place in Activity 1.1, where
heat is given out along with the formation of a single product.
Do Y
ou Know?
A solution of slaked lime produced by the reaction 1.13 is used for whitewashing
walls. Calcium hydroxide reacts slowly with the carbon dioxide in air to form a thin
layer of calcium carbonate on the walls. Calcium carbonate is formed after two to
three days of whitewashing and gives a shiny finish to the walls. It is interesting to note
that the chemical formula for marble is also CaCO
3
.
Ca(OH)
2
(aq) + CO
2
(g)→CaCO
3
(s) + H
2
O(l)(1.14)
(Calcium(Calcium
hydroxide)carbonate)
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Figure 1.5
Heating of lead nitrate and
emission of nitrogen dioxide
Figure 1.4
Correct way of heating
the boiling tube
containing crystals
of ferrous sulphate
and of smelling the
odour
Activity 1.6Activity 1.6
Activity 1.6Activity 1.6
Activity 1.6
nTake about 2 g lead nitrate powder in a boiling
tube.
nHold the boiling tube with a pair of tongs and
heat it over a flame, as shown in Fig. 1.5.
nWhat do you observe? Note down the change,
if any.
You will observe the emission of brown fumes.
These fumes are of nitrogen dioxide (NO
2
). The
reaction that takes place is –
Activity 1.5Activity 1.5
Activity 1.5Activity 1.5
Activity 1.5
nTake about 2 g ferrous sulphate crystals
in a dry boiling tube.
nNote the colour of the ferrous sulphate
crystals.
nHeat the boiling tube over the flame of
a burner or spirit lamp as shown in
Fig. 1.4.
nObserve the colour of the crystals after
heating.
Have you noticed that the green colour of the ferrous sulphate crystals
has changed? You can also smell the characteristic odour of burning
sulphur.
2FeSO
4
(s)
Heat
→
Fe
2
O
3
(s) + SO
2
(g) + SO
3
(g)(1.19)
(Ferrous sulphate) (Ferric oxide)
In this reaction you can observe that a single reactant breaks down
to give simpler products. This is a decomposition reaction. Ferrous
sulphate crystals (FeSO
4
. 7H
2
O) lose water when heated and the colour
of the crystals changes. It then decomposes to ferric oxide (Fe
2
O
3
),
sulphur dioxide (SO
2
) and sulphur trioxide (SO
3
). Ferric oxide is a solid,
while SO
2
and SO
3
are gases.
Decomposition of calcium carbonate to calcium oxide and carbon
dioxide on heating is an important decomposition reaction used in
various industries. Calcium oxide is called lime or quick lime. It has
many uses – one is in the manufacture of cement. When a decomposition
reaction is carried out by heating, it is called thermal decomposition.
CaCO
3
(s)
Heat
→ 
CaO(s) + CO
2
(g)(1.20)
(Limestone) (Quick lime)
Another example of a thermal decomposition reaction is given
in Activity 1.6.
1.2.2 Decomposition Reaction
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Activity 1.7Activity 1.7
Activity 1.7Activity 1.7
Activity 1.7
Activity 1.8Activity 1.8
Activity 1.8Activity 1.8
Activity 1.8
nTake about 2 g silver chloride in a china dish.
nWhat is its colour?
nPlace this china dish in sunlight for some time
(Fig. 1.7).
nObserve the colour of the silver chloride after some
time.
Figure 1.7
Silver chloride turns grey
in sunlight to form silver
metal
You will see that white silver chloride turns grey in sunlight. This is
due to the decomposition of silver chloride into silver and chlorine by
light.
2AgCl(s)
Sunlight
→
2Ag(s) + Cl
2
(g)(1.22)
nTake a plastic mug. Drill two holes at its
base and fit rubber stoppers in these holes.
Insert carbon electrodes in these rubber
stoppers as shown in Fig. 1.6.
nConnect these electrodes to a 6 volt
battery.
nFill the mug with water such that the
electrodes are immersed. Add a few drops
of dilute sulphuric acid to the water.
nTake two test tubes filled with water and
invert them over the two carbon electrodes.
nSwitch on the current and leave the
apparatus undisturbed for some time.
nYou will observe the formation of bubbles
at both the electrodes. These bubbles displace water in the
test tubes.
nIs the volume of the gas collected the same in both the test tubes?
nOnce the test tubes are filled with the respective gases, remove
them carefully.
nTest these gases one by one by bringing a burning candle close
to the mouth of the test tubes.
CAUTION: This step must be performed carefully by the teacher.
nWhat happens in each case?
nWhich gas is present in each test tube?
Figure 1.6
Electrolysis of water
2Pb(NO
3
)
2
(s)
Heat
→
2PbO(s)+4NO
2
(g) + O
2
(g)(1.21)
(Lead nitrate) (Lead oxide)(Nitrogen (Oxygen)
dioxide)
Let us perform some more decomposition reactions as given in
Activities 1.7 and 1.8.
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Figure 1.8
(a) Iron nails dipped in copper sulphate solution
?
QUESTIONS
1.A solution of a substance ‘X’ is used for whitewashing.
(i)Name the substance ‘X’ and write its formula.
(ii)Write the reaction of the substance ‘X’ named in (i) above with
water.
2.Why is the amount of gas collected in one of the test tubes in Activity
1.7 double of the amount collected in the other? Name this gas.
1.2.3 Displacement Reaction
Activity 1.9Activity 1.9
Activity 1.9Activity 1.9
Activity 1.9
nTake three iron nails and clean them by
rubbing with sand paper.
nTake two test tubes marked as (A) and
(B). In each test tube, take about 10 mL
copper sulphate solution.
nTie two iron nails with a thread and
immerse them carefully in the copper
sulphate solution in test tube B for
about 20 minutes [Fig. 1.8 (a)]. Keep one
iron nail aside for comparison.
nAfter 20 minutes, take out the iron nails
from the copper sulphate solution.
nCompare the intensity of the blue colour
of copper sulphate solutions in test tubes
(A) and (B) [Fig. 1.8 (b)].
nAlso, compare the colour of the iron nails
dipped in the copper sulphate solution
with the one kept aside [Fig. 1.8 (b)].
Take about 2 g barium hydroxide in a test tube. Add 1 g of ammonium chloride and mix
with the help of a glass rod. Touch the bottom of the test tube with your palm. What do you
feel? Is this an exothermic or endothermic reaction?
Carry out the following Activity
Silver bromide also behaves in the same way.
2AgBr(s)
Sunlight
→ 
2Ag(s) + Br
2
(g)(1.23)
The above reactions are used in black and white photography.
What form of energy is causing these decomposition reactions?
We have seen that the decomposition reactions require energy either
in the form of heat, light or electricity for breaking down the reactants.
Reactions in which energy is absorbed are known as endothermic
reactions.
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Why does the iron nail become brownish in colour and the blue colour
of copper sulphate solution fades?
The following chemical reaction takes place in this Activity–
Fe(s) + CuSO
4
(aq)→FeSO
4
(aq) + Cu(s)(1.24)
(Copper sulphate)(Iron sulphate)
In this reaction, iron has displaced or removed another element,
copper, from copper sulphate solution. This reaction is known as
displacement reaction.
Other examples of displacement reactions are
Zn(s) + CuSO
4
(aq) →ZnSO
4
(aq) + Cu(s)(1.25)
(Copper sulphate)(Zinc sulphate)
Pb(s) + CuCl
2
(aq) →PbCl
2
(aq) + Cu(s)(1.26)
(Copper chloride)(Lead chloride)
Zinc and lead are more reactive elements than copper. They displace
copper from its compounds.
1.2.4 Double Displacement Reaction
Activity 1.10Activity 1.10
Activity 1.10Activity 1.10
Activity 1.10
nTake about 3 mL of sodium sulphate
solution in a test tube.
nIn another test tube, take about 3 mL of
barium chloride solution.
nMix the two solutions (Fig. 1.9).
nWhat do you observe?
Figure 1.9Figure 1.9
Figure 1.9Figure 1.9
Figure 1.9
Formation of barium
sulphate and sodium
chloride
You will observe that a white substance, which is
insoluble in water, is formed. This insoluble substance
formed is known as a precipitate. Any reaction that
produces a precipitate can be called a precipitation reaction.
Na
2
SO
4
(aq) + BaCl
2
(aq) → BaSO
4
(s) + 2NaCl(aq)(1.27)
(Sodium(Barium(Barium(Sodium
sulphate)chloride)sulphate)chloride)
Figure 1.8 Figure 1.8
Figure 1.8 Figure 1.8
Figure 1.8 (b) Iron nails and copper sulphate solutions compared before and after the experiment
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1.2.5 Oxidation and Reduction
Activity 1.11Activity 1.11
Activity 1.11Activity 1.11
Activity 1.11
nHeat a china dish containing about 1 g
copper powder (Fig. 1.10).
nWhat do you observe?
Figure 1.10Figure 1.10
Figure 1.10Figure 1.10
Figure 1.10
Oxidation of copper to
copper oxide
The surface of copper powder becomes coated with
black copper(II) oxide. Why has this black
substance formed?
This is because oxygen is added to copper and
copper oxide is formed.
2Cu + O
2
Heat
→
2CuO(1.28)
If hydrogen gas is passed over this heated material (CuO), the black
coating on the surface turns brown as the reverse reaction takes place
and copper is obtained.
CuO + HCu + H O
22
Heat
→
(1.29)
If a substance gains oxygen during a reaction, it is said to be oxidised.
If a substance loses oxygen during a reaction, it is said to be reduced.
During this reaction (1.29), the copper(II) oxide is losing oxygen and
is being reduced. The hydrogen is gaining oxygen and is being oxidised.
In other words, one reactant gets oxidised while the other gets reduced
during a reaction. Such reactions are called oxidation-reduction reactions
or redox reactions.
(1.30)
Some other examples of redox reactions are:
ZnO + C →+Zn CO
(1.31)
MnOHClMnClH O Cl
2222
42+→++
(1.32)
Recall Activity 1.2Recall Activity 1.2
Recall Activity 1.2Recall Activity 1.2
Recall Activity 1.2, where you have mixed the solutions of lead(II) nitrate
and potassium iodide.
(i)What was the colour of the precipitate formed? Can you name the compound
precipitated?
(ii)Write the balanced chemical equation for this reaction.
(iii)Is this also a double displacement reaction?
What causes this? The white precipitate of BaSO
4
is formed by the
reaction of
2–
4
SO
and Ba
2+
. The other product formed is sodium chloride
which remains in the solution. Such reactions in which there is an
exchange of ions between the reactants are called double displacement
reactions.
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In reaction (1.31) carbon is oxidised to CO and ZnO is reduced to Zn.
In reaction (1.32) HCl is oxidised to Cl
2
whereas MnO
2
is reduced to MnCl
2
.
From the above examples we can say that if a substance gains oxygen
or loses hydrogen during a reaction, it is oxidised. If a substance loses
oxygen or gains hydrogen during a reaction, it is reduced.
QUESTIONS
?
1.Why does the colour of copper sulphate solution change when
an iron nail is dipped in it?
2.Give an example of a double displacement reaction other than
the one given in Activity 1.10.
3.Identify the substances that are oxidised and the substances
that are reduced in the following reactions.
(i)4Na(s) + O
2
(g) → 2Na
2
O(s)
(ii)CuO(s) + H
2
(g) → Cu(s) + H
2
O(l)
1.3
1.3
1.31.3
1.3
HAHA
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XIDXID
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1.3.1 Corrosion
You must have observed that iron articles are shiny when new, but get
coated with a reddish brown powder when left for some time. This process
is commonly known as rusting of iron. Some other metals also get
tarnished in this manner. Have you noticed the colour of the coating
formed on copper and silver? When a metal is attacked by substances
around it such as moisture, acids, etc., it is said to corrode and this
process is called corrosion. The black coating on silver and the green
coating on copper are other examples of corrosion.
Corrosion causes damage to car bodies, bridges, iron railings, ships
and to all objects made of metals, specially those of iron. Corrosion of
iron is a serious problem. Every year an enormous amount of money is
spent to replace damaged iron. You will learn more about corrosion in
Chapter 3.
1.3.2 Rancidity
Have you ever tasted or smelt the fat/oil containing food materials left
for a long time?
When fats and oils are oxidised, they become rancid and their smell
and taste change. Usually substances which prevent oxidation
(antioxidants) are added to foods containing fats and oil. Keeping food
in air tight containers helps to slow down oxidation. Do you know that
chips manufacturers usually flush bags of chips with gas such as
nitrogen to prevent the chips from getting oxidised ?
Recall Activity 1.1Recall Activity 1.1
Recall Activity 1.1Recall Activity 1.1
Recall Activity 1.1,
where a magnesium ribbon burns with a dazzling flame in air (oxygen)
and changes into a white substance, magnesium oxide. Is magnesium being oxidised or
reduced in this reaction?
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What you have learnt
nA complete chemical equation represents the reactants, products and their physical
states symbolically.
nA chemical equation is balanced so that the numbers of atoms of each type involved
in a chemical reaction are the same on the reactant and product sides of the
equation. Equations must always be balanced.
nIn a combination reaction two or more substances combine to form a new single
substance.
nDecomposition reactions are opposite to combination reactions. In a decomposition
reaction, a single substance decomposes to give two or more substances.
nReactions in which heat is given out along with the products are called exothermic
reactions.
nReactions in which energy is absorbed are known as endothermic reactions.
nWhen an element displaces another element from its compound, a displacement
reaction occurs.
nTwo different atoms or groups of atoms (ions) are exchanged in double displacement
reactions.
nPrecipitation reactions produce insoluble salts.
nReactions also involve the gain or loss of oxygen or hydrogen by substances.
Oxidation is the gain of oxygen or loss of hydrogen. Reduction is the loss of oxygen
or gain of hydrogen.
EXERCISES
1.Which of the statements about the reaction below are incorrect?
2PbO(s) + C(s) → 2Pb(s) + CO
2
(g)
(a)Lead is getting reduced.
(b)Carbon dioxide is getting oxidised.
(c)Carbon is getting oxidised.
(d)Lead oxide is getting reduced.
(i)(a) and (b)
(ii)(a) and (c)
(iii)(a), (b) and (c)
(iv)all
2.Fe
2
O
3
+ 2Al → Al
2
O
3
+ 2Fe
The above reaction is an example of a
(a)combination reaction.
(b)double displacement reaction.
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(c)decomposition reaction.
(d)displacement reaction.
3.What happens when dilute hydrochloric acid is added to iron fillings? Tick the
correct answer.
(a)Hydrogen gas and iron chloride are produced.
(b)Chlorine gas and iron hydroxide are produced.
(c)No reaction takes place.
(d)Iron salt and water are produced.
4.What is a balanced chemical equation? Why should chemical equations be
balanced?
5.Translate the following statements into chemical equations and then balance them.
(a)Hydrogen gas combines with nitrogen to form ammonia.
(b)Hydrogen sulphide gas burns in air to give water and sulpur dioxide.
(c)Barium chloride reacts with aluminium sulphate to give aluminium chloride
and a precipitate of barium sulphate.
(d)Potassium metal reacts with water to give potassium hydroxide and hydrogen
gas.
6.Balance the following chemical equations.
(a)HNO
3
+
Ca(OH)
2
→ Ca(NO
3
)
2
+ H
2
O
(b)NaOH + H
2
SO
4
→ Na
2
SO
4
+ H
2
O
(c)NaCl + AgNO
3
→ AgCl + NaNO
3
(d)BaCl
2
+ H
2
SO
4
→ BaSO
4
+ HCl
7.Write the balanced chemical equations for the following reactions.
(a)Calcium hydroxide + Carbon dioxide
→ Calcium carbonate + Water
(b)Zinc + Silver nitrate → Zinc nitrate + Silver
(c)Aluminium + Copper chloride → Aluminium chloride + Copper
(d)Barium chloride + Potassium sulphate → Barium sulphate + Potassium chloride
8.Write the balanced chemical equation for the following and identify the type of
reaction in each case.
(a)Potassium bromide(aq) + Barium iodide(aq) → Potassium iodide(aq) +
Barium bromide(s)
(b)Zinc carbonate(s) → Zinc oxide(s) + Carbon dioxide(g)
(c)Hydrogen(g) + Chlorine(g) → Hydrogen chloride(g)
(d)Magnesium(s) + Hydrochloric acid(aq) → Magnesium chloride(aq) + Hydrogen(g)
9.What does one mean by exothermic and endothermic reactions? Give examples.
10.Why is respiration considered an exothermic reaction? Explain.
11.Why are decomposition reactions called the opposite of combination reactions?
Write equations for these reactions.
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Group Activity
Perform the following activity.
nTake four beakers and label them as A, B, C and D.
nPut 25 mL of water in A, B and C beakers and copper sulphate solution in beaker D.
nMeasure and record the temperature of each liquid contained in the beakers above.
nAdd two spatulas of potassium sulphate, ammonium nitrate, anhydrous copper
sulphate and fine iron fillings to beakers A, B, C and D respectively and stir.
nFinally measure and record the temperature of each of the mixture above.
Find out which reactions are exothermic and which ones are endothermic in nature.
12.Write one equation each for decomposition reactions where energy is supplied in
the form of heat, light or electricity.
13.What is the difference between displacement and double displacement reactions?
Write equations for these reactions.
14. In the r
efining of silver, the recovery of silver from silver nitrate solution involved
displacement by copper metal. Write down the reaction involved.
15.What do you mean by a precipitation reaction? Explain by giving examples.
16.Explain the following in terms of gain or loss of oxygen with two examples each.
(a)Oxidation
(b)Reduction
17.A shiny brown coloured element ‘X’ on heating in air becomes black in colour.
Name the element ‘X’ and the black coloured compound formed.
18.Why do we apply paint on iron articles?
19.Oil and fat containing food items are flushed with nitrogen. Why?
20.Explain the following terms with one example each.
(a)Corrosion
(b)Rancidity
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