**CHAPTER**

Acids and bases

This chapter builds on the acid–base chemistry you studied in Year 11. You will examine the nature of acids and bases from the point of view of proton exchange. The role of water in acid–base chemistry will be discussed, as will the concept of pH as a measure of acidity.

You will learn to carry out calculations involving pH. The ideas about equilibrium systems that you explored in Chapter 2 will be applied to your understanding of acid–base chemistry.

**Science as a human endeavour**

Models and theories are contested and refined or replaced when new evidence challenges them, or when a new model or theory has greater explanatory scope. Davy initially proposed that acids were substances that contained replaceable hydrogen (hydrogen that could be partly or totally replaced by metals) and bases were substances that reacted with acids to form salts and water. The Arrhenius model, which includes only soluble acids and bases, identified acids as substances which produce hydrogen ions in solution and bases as substances which produce hydroxide ions in solution. Subsequently, the Brønsted–Lowry model describes acid–base behaviour in terms of proton donors and proton acceptors. This approach includes a wider range of substances and can be more broadly applied. Science understanding

• observable changes in chemical reactions and physical changes can be described and explained at an atomic and molecular level

• acids are substances that can act as proton (hydrogen ion) donors and can be classified as monoprotic or polyprotic, depending on the number of protons available for donation

• the strength of acids is explained by the degree of ionisation at equilibrium in aqueous solution, which can be represented by chemical equations and acidity constants (K

a

e

P

r

o

o

• P

• • • the explained equations pairs the equations; and a relationship hydrolysis neutral using that the nature g ) of Brønsted–Lowry illustrate between the salts of Brønsted–Lowry salts of acids and bases in equilibrium systems can be

model and represented using chemical the transfer of protons between conjugate acid-base

weak acids and weak bases can be represented using

model can be applied to explain the acidic, basic derived from bases and monoprotic and polyprotic acids water is a weak electrolyte; the self-ionisation of water is represented by K

w

= [H+] [OH−] where K

w

= 1.0 × 10−14 at 25°C K

w

can be used to calculate the concentration of hydrogen ions or hydroxide ions in solutions of strong acids or bases

• the pH scale is a logarithmic scale and the pH of a solution can be calculated from the concentration of hydrogen ions using the relationship pH = −log

10

[H+]

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f

s

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CHEMFILE

Saving the Nobel Prize gold medal George de Hevesy (1885–1966) worked for the Niels Bohr Institute in Denmark during World War II. The Institute was looking after a number of valuable gold medals that had been awarded to recipients of the Nobel Prize. The Nobel Prize is awarded for outstanding contributions to the fields of physics, chemistry, literature, peace, physiology or medicine, and economics. When Germany invaded Denmark at the beginning of World War II, de Hevesy was concerned that the Germans would confiscate the gold medals. He dissolved the gold medals in aqua regia, which is a mixture of concentrated hydrochloric and nitric acids. Aqua regia is strong enough to dissolve gold, a metal known for its lack of chemical reactivity. He hid the bottle containing the gold solution among the hundreds of other bottles on his laboratory shelves. The bottle was never found by the German occupiers and after the war de Hevesy retrieved the precious bottle and precipitated the gold out of solution. The gold was sent to the Nobel foundation where the medals were recast into duplicates of the originals and returned to FIGURE their 4.1.2 rightful P Recipients owners (Figure of the a Nobel

4.1.2).

Prize are given a sum of money and a gold medal weighing about 175 g.

AREA OF STUDY 2 | ACIDS AND BASES

g

**4.1 Introducing acids and bases**

Acids and bases make up some of the most commonly used products in our daily lives (Figure 4.1.1). They are important industrial and agricultural chemicals and are significant in a biological and environmental context. Acids and bases are found in many household products. Bases, in particular sodium hydroxide (caustic soda) and sodium carbonate (soda ash), are used in many cleaning products for their ability to react with oils and grease.

In this section, you will be introduced to a theory that explains the chemical properties of acids and bases, helping you to understand their usefulness within the home and industry as well as their role in the environment. You will also look at how to measure the acidity of a solution, therefore defining the level of acidity or basicity of a solution.

**Common FIGURE 4.1.1**

household products that are acids, bases or salts

ACIDS AND BASE Acids are used in our homes, in agriculture and in industry. They also have an important role in our bodies. Table 4.1.1 gives the names, chemical formulae and uses of some common acids.

**TABLE 4.1.1**

Common acids and their everyday uses

**Name Formula Uses**

Hydrochloric acid HCl Present in stomach acid to help break down

proteins. Used as a cleaning agent for brickwork

Sulfuric acid H

2

SO

4

One of the most common chemicals manufactured. Used in car batteries and in the manufacture of fertilisers and detergents

Nitric acid HNO

3

Used in the manufacture of fertilisers, dyes and explosives

Ethanoic (acetic) acid CH

3

COOH Found in vinegar. Used as a preservative

Carbonic acid H

2

CO

3

Found in carbonated soft drinks and beer

Phosphoric acid H

3

Used in some soft drinks and in the manufacture of fertilisers

Citric acid C

6

PO

4

H

8

O

7

Found in citrus fruits

Ascorbic acid C

6

H

8

O

6

Found in citrus fruits (vitamin C)

e

P

r

o

o

f

s

Many cleaning agents used in the home, such as washing powders and oven cleaners, contain bases. Solutions of ammonia are used as floor cleaners, and sodium hydroxide is the major active ingredient in oven cleaner. Bases are effective cleaners because they react with fats and oils to produce water-soluble soaps. A soluble base is referred to as an alkali.

Table 4.1.2 gives the names, chemical formulae and uses of some common bases.

**TABLE 4.1.2**

Common bases and their uses

**Name Formula Uses**

Sodium hydroxide (caustic soda)

NaOH Used in drain and oven cleaners, and in soap

making

Ammonia NH

3

Used in household cleaners, fertilisers and explosives

Calcium hydroxide Ca(OH)

2

Found in cement and mortar. Used in garden lime to adjust soil pH

PROPERTIES All The TABLE CHANGING AND Over acids effect indicators the chemist, While Magnesium Sodium Properties Turn Tend Taste React Solutions Solutions In acids basis properties and 4.1.3 the litmus on to sour the this with BASES be carbonate P indicators of years, have thought bases have conduct Properties in late bases explanation corrosive of indicator hydroxide the Chapter acids a some 18th were of there nature relatively IDEAS an acids OF that of and red electric defined century, a properties acids have 6.) ACIDS of applied and reactions acidic low Mg(OH) Na ABOUT and their been 2 current CO pH bases in bases attempts 3 terms properties constituent 2 to in AND many g with are sulfuric common. THE of summarised other attempts BASES their were Key of Used and Properties Turn Are Taste React Solutions Solutions were magnesia, acid NATURE elements. substances. ingredient glass observed e Bases made in caustic litmus bitter the with (H to due in 2 define have conduct manufacture also acids to used of and Table indicator properties, to Antoine in bases define OF a (You some have feel the to relatively acids an 4.1.3. overcome ACIDS slippery presence blue will antacids, electric common acids Lavoisier, of and P such washing learn high bases. current indigestion and as such pH more of properties. their powder bases a as oxygen. At French r

about milk

taste, first,

on

o

o

f

s

Oven cleaner removes fatty deposits by turning them into soap. Note the safety instructions.

phosphoric SO

4

), nitric acid (HNO

3

) and

an acid.

acid (H

3

PO

4

), it did not explain why hydrochloric acid (HCl) was

In about 1810, Humphrey Davy suggested that the acid properties of substances were associated with hydrogen and not oxygen. He came to this conclusion after producing hydrogen gas by reacting acids with metals. Davey also suggested that acids react with bases to form salts and water.

Aqueous solutions of bases are called alkalis. Calcium carbonate will react with acids. However, it is not considered an alkali because it is insoluble in water.

CHEMFILE

Handle strong bases with care Bases feel slippery to the touch because they react with fats in your skin to produce soap. Strong bases should be handled with care. Oven cleaners contain about 4% of the strong base sodium hydroxide. A common name of sodium hydroxide is caustic soda. Figure 4.1.3 shows the safety instructions on a can of oven cleaner.

**FIGURE 4.1.3**

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CHEMISTRY IN ACTION

**Naturally occurring acids**

Rainwater has a pH of about 6. Rainwater is naturally acidic because of the presence of carbon dioxide in the atmosphere. Carbon dioxide dissolves in water in the atmosphere and oceans forming carbonic acid:

CO

2

(g) + H

2

O(l) → H

2

CO

3

(aq)

The pH of rain can drop below 6 when other gases, particularly oxides of non-metals such as sulfur and nitrogen, combine chemically with water and oxygen in the atmosphere and oceans, producing acids.

Sulfur dioxide and sulfur trioxide are both released in large quantities by volcanoes and combine with water in the atmosphere to produce sulfuric and sulfurous acids. Nitrogen oxides are produced when nitrogen combines explosively with oxygen during lightning strikes. The gases dissolve in atmospheric water to produce nitrous and nitric acid.

These gases and corresponding acids are an important part of the natural cycles of the Earth. However, their production from the combustion of fossil fuels and other industry is causing an imbalance that is contributing to large-scale problems such as acid rain and ocean acidification. Acid rain affects food chains at the decomposer level by disrupting the recycling of nutrients and killing aquatic plants and animals. It also affects the surface of plant leaves, reducing their ability to prevent water loss, which leads to deforestation. Increased acidity in the oceans, due mainly to increased atmospheric carbon dioxide levels, reduces the capacity of marine organisms such as shellfish and coral reef organisms to remove calcium carbonate from the water for shell and reef building. Table 4.1.4 shows the major gaseous non- metal oxides produced from the combustion of fossil fuels and the corresponding acids that they form.

AREA OF STUDY 2 | ACIDS AND BASES

**f TABLE 4.1.4**

Non-metal oxides and the acids that they form

**Oxide Acid produced in the environment**

CO

2

(carbonic acid)

SO

2

H

CO

3

, SO

3

NO, NO

2 acid Figure P rain.

4.1.4 shows r H H

NO HNO the 2 2

2

SO SO

2

reacts 3 3

4 o damage (nitric (sulfurous (sulfuric to acid

form: caused acid) acid) o HNO to 2

(nitrous forests acid)

by

P

a

g 1887, • • chemist This Acids water. Bases In e

Forest with dead trees destroyed by pollution and acid rain

theory was further developed by the Swedish scientist Svante Arrhenius. In he defined acids and bases as follows.

are substances that dissociate (break apart) and ionise (form ions) in They produce hydrogen ions (H+). dissociate in water to produce hydroxide ions (OH−). 1923, Danish physical chemist Johannes Nicolaus Brønsted and English Thomas Martin Lowry were working independently. They each came up with the theory that now bears both of their names. The Brønsted–Lowry theory is more general than the one proposed by Arrhenius and provides an explanation for some observed acid–base behaviours that cannot be explained by the earlier theories.

**FIGURE 4.1.4**

s

BRØNSTED−LOWRY THEORY OF ACIDS AND BASES According to the Brønsted−Lowry theory, a substance behaves as an acid when it donates a proton, i.e. H+, to a base. A substance behaves as a base when it accepts a proton from an acid.

In summary:

• acids are proton donors

• bases are proton acceptors

• an acid–base reaction involves an exchange of protons from an acid to a base. For example, hydrogen chloride (HCl) is a gaseous molecular compound that is very soluble in water. The molecules ionise in water according to the following reaction:

HCl(g) + H

2

O(l) → H

3

O+(aq) + Cl−(aq) In an aqueous solution of hydrogen chloride, nearly all the hydrogen chloride is present as ions—virtually no molecules of hydrogen chloride remain. This solution is known as hydrochloric acid.

forming theory, the the or with H+(aq) has chapter. in Advantages Acid–base gases be • • reactions gas Limitations Like Figure a formed In The A The solutions gaseous The Figure HCl reaction and occurred. dissociation all water, hydrolysis can this HCl hydronium structural in theories, ammonia molecule, Brønsted–Lowry P the because 4.1.6. be reaction, 4.1.7 reactions by a hydrogen either of has of dissociation an hydronium a Therefore, HCl(g) hydrochloric reaction acid–base reaction shows acted reaction: the of of in formulae gas. accepting so HCl(aq) ion each are each HCl(g) a Brønsted–Lowry the the the chloride as with can the not between: occurs HCl water the an HCl(g) case equation reaction. model Brønsted–Lowry HCl(g) gaseous be Brønsted–Lowry ion of restricted water + acid. acid + or hydronium represented molecule the NH and the H has donating when (H 2

g and + O(l) The hydronium can classifies acid 3 gaseous → acted (aq) 3 NH makes acid–base For O+(aq)). to ammonia: H+(aq) be a model water → donates 3 has → molecule example, (g) aqueous written ion as H a as it NH ammonia: 3 proton. a both donated → O+(aq) molecule either is e harder ion has base. + According reaction NH 4 usually a Cl−(aq) +(aq) as and proton solutions. of some or the 4 This H either model Cl(s) Writing + model ion these a to 3 the + salt O+(aq) Cl−(aq) has proton written between Cl−(aq) limitations. see is in to water a ammonium to outlined accepted reactions aqueous A the hydrolysis that the the or reaction to base. as molecule H+(aq). P hydrogen hydronium a a Brønsted–Lowry H water proton in For 3 solution a O+(aq) as Figure between proton chloride example, reaction: Therefore,

acid−base are molecule,

chloride transfer

shown r in ion reacts 4.1.5. from

two this

can as

it cannot be applied to the reaction between acidic and basic oxides. Under certain conditions solid calcium oxide, a basic oxide, reacts with gaseous carbon dioxide, an acidic oxide, to produce the salt calcium carbonate:

CaO(s) + CO

2

(g) → CaCO

3

(s)

o

O+(aq) + Cl–(aq) acid base

**FIGURE 4.1.5**

The reaction between hydrochloric acid and water is an example of an acid–base reaction, which involves a transfer of protons from an acid to a base.

O

H H

+

+

+ + +

H

H+

**FIGURE 4.1.6**

O

H H

Formation of the hydronium ion

Remember that H+(aq) is often used interchangeably with H donating 3

O+(aq) when protons.

discussing acids

**A FIGURE 4.1.7**

gaseous acid–base reaction between hydrogen chloride gas and ammonia gas. Fumes of solid ammonium chloride are formed.

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H+ HCl(g) + o H

2

O(l) H

3

f

s

**90**

Acids donate a proton to form:

HCl H

2

− H+

Cl− SO

4

HSO

4 HNO

3 H

3

−

NO

3

−

O+

H

2 CH

3

O COOH

CH

3 H

2

COO− CO

3

HCO

3 H

3

−

PO

4

H

2 NH

4

PO

4

−

+

NH

3 HCO

3

−

CO

3 H

2

2−

O

OH−

Bases accept a proton to form:

OH− NH

3 CO

3

+ H+

H

2

O NH

4

+

2−

HCO

3 PO

4

−

3−

HPO

4 CH

3

2−

COO–

CH

3 O2−

COOH OH−

**FIGURE 4.1.10**

Some common acid–base conjugate pairs

When hydronium produced.P produced. water, acids hydroxide When (H react 3 O+) bases (OH−) with ions a water, are ions react are

with

AREA OF STUDY 2 | ACIDS AND BASES

CONJUGATE ACID−BASE PAIRS A solution of hydrochloric acid is produced when hydrogen chloride ionises in water:

HCl(aq) + H

2

O(l) → H

3

O+(aq) + Cl−(aq) Because Cl− can be formed from HCl by the loss of a single proton, it is called the conjugate base of HCl. Similarly, HCl is described as the conjugate acid of Cl−. are a HCl conjugate and Cl− pair.

are called a conjugate acid−base pair. Similarly, H

3

O+ and H

2

O

A conjugate acid–base pair is two species that differ by a proton, i.e. H+. There are two conjugate acid–base pairs in an acid–base reaction. For the reaction between HCl and H

2

O, the conjugate pairs are HCl/Cl− and H

3

O+/H

2

O. The relationship between acid–base conjugate pairs is represented in Figure 4.1.8. In the reaction between NH acid–base corresponding pairs base are by NH

one 4

+/NH

proton. 3

3 and and H

H

2

O/OH− 2

O (shown because in Figure each 4.1.9), acid the conjugate differs from its

HCl(g) + H

2 O(l) acid base FIGURE 4.1.8 Conjugate formed when an acid Figure 4.1.10 conjugates.

**FIGURE NH base o 3 (aq) 4.1.9 + H**

acid 2 O(l) o NH

acid 4 +(aq) f

+ OH–(aq)

base s

AMPHIPROTIC Some substances g with. to as acid: be a For However, These base: Therefore, amphiprotic. e example, reactions in they in donates can shows the acid–base SUBSTANCES H donate can 3 following O+(aq) acid a the P proton behave pairs formulae or + to accept Cl–(aq) are base

a reaction, as base. either r of protons HCl(g) + H 2 O(l) → the reaction below water NH

3

(aq) + H

2 O(l)  are represented in Figure +

The reaction between ammonia and water, showing the conjugate acid–base pairs

some common acids and bases and depending on what they are reacting acids or bases. Such substances are water gains a proton from HCl and Cl−(aq) + donates a NH

4 +(aq) 4.1.11.

their

said

acts

H

**3**

O+(aq) proton to NH

3

and acts as an

+ OH−(aq)

H

+

−

H H

−

+ +

**FIGURE 4.1.11**

Cl H

O O

+ Cl H H

HCl H

2

Cl−

H

H H

O+

NH

3

O

H

3

O H H

H

+

N

H

HN

O H

H

H

2

O

NH

4

+ OH− The amphiprotic nature of water is demonstrated by its reactions with HCl and NH

3

.

It is evident that water can act as either an acid or a base, depending on the solute present.

If the solute is a stronger acid than water, then water will react as a base. If the solute is a stronger base than water, then water will react as an acid.

Some common amphiprotic substances are listed in Figure 4.1.12.

Amphiprotic substance

Accepts a proton to form:

− H+ + H+ P a

g

e Donates a proton to form:

OH− CO

3

H

3

O+ H

2

CO

3 H

3

PO

4 H

2

PO

4

−

H

2

SO

4

P

r

o

FIGURE Each molecule H H

**4.1.13 H**

o

f

s

O

O – H + acidic proton

The structure of ethanoic acid. can donate only one proton to a water molecule when ethanoic acid reacts with water. H

2

O

2−

HCO

3 HPO

4

−

2−

H

2

PO

4 PO

4

−

3−

HPO

4 SO

4

2−

2−

HSO

4

−

**FIGURE 4.1.12**

Substances that are amphiprotic

When an amphiprotic substance is placed in water, it reacts as both an acid and a and base. a base For according example, the to the hydrogen equations:

carbonate ion (HCO

3

−) can act as both an acid

HCO

acid 3

−(aq) + H

base

2

O(l)  CO

3

2−(aq) + H

3

O+(aq)

HCO

base 3

−(aq) + H

acid

2

O(l)  H

2

CO

3

(aq) + OH−(aq)

Although both reactions are possible for all amphiprotic substances in water, generally one reaction dominates. The dominant reaction can be identified by measuring the pH (a measure of the amount of hydronium ion in solution) of the solution. (You will look more closely at pH in section 4.3.)

MONOPROTIC ACIDS Monoprotic acids can donate only one proton per molecule. These acids include hydrochloric acid (HCl), hydrofluoric (acetic) While acid ethanoic (CH 3

COOH). acid contains four acid (HF), nitric acid (HNO 3

) and ethanoic

hydrogen atoms, each molecule can donate only one monoprotic proton to produce an acid. Only the hydrogen ethanoate that is ion part (CH

of the 3

COO−) highly and polar it is therefore a O−H bond is donated. This hydrogen atom is called the acidic proton (Figure 4.1.13).

POLYPROTIC ACIDS Some acids can donate more than one proton from each molecule and are said to be polyprotic acids. The number of hydrogen ions an acid can donate depends on the structure of the acid. Polyprotic acids do not donate all of their protons at once, but do so in steps when reacting with a base. Diprotic acids Diprotic acids can donate (H

2

CO

3

) are diprotic acids.

two protons. Sulfuric acid (H

2

SO

4

) and carbonic acid

A diprotic acid, such as sulfuric acid, ionises in two stages: Stage 1: H

2

SO

4

(l) + H

2

O(l) → HSO

4

−(aq) + H

3

O+(aq) Stage 2: HSO

4

−(aq) + H

2

O(l)  SO

4

2−(aq) + H

3

O+(aq) Sulfuric acid is described as a strong acid in water because it readily donates a proton are found and in so an this aqueous stage occurs solution.

almost to completion. Virtually no H

2

SO

4

molecules

small The proportion HSO

4

− ion formed can also act as an acid. In a 1.0 mol L−1 solution, only a

double HSO

(reversible) 4

− of those ions reacts further to produce is described as a weak acid because it is H

only 3

O+ ions partially and SO

ionised. 4

2− ions. A arrow () indicates that an incomplete reaction occurs and an equilibrium is formed. (You will learn more about strong and weak acids later in this chapter.) Therefore, a solution of sulfuric acid contains hydrogen ions, hydrogen sulfate ions and sulfate ions.

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C C

**92**

) and boric acid (H

3

BO

3

) are triprotic acids. A triprotic acid, such as phosphoric acid, ionises in three stages: Stage 1: H

3

PO

4

(aq) + H

2

O(l)  H

2

PO

4

−(aq) + H

3

O+(aq) Stage 2: H

2

PO

4

−(aq) + H

2

O(l)  HPO

4

2−(aq) + H

3

O+(aq)

s

**4.1 Review**

Stage 3: HPO

4

2−(aq) + H

2 O(l) Phosphoric acid is a weak of phosphoric acid, only a ionisation stage. to stage 3.

 PO

4

3−(aq) + H

3

O+(aq) acid in water. Therefore, small proportion of the protons in a 1.0 is donated mol f

L−1 solution at each

SUMMARY

• The Brønsted–Lowry model describes acid–base properties in terms of proton transfer. A Brønsted– Lowry acid is a proton donor and a base is a proton acceptor.

• When an acid donates a proton, it forms its conjugate base. When a base accepts a proton, it forms its conjugate acid.

The extent of the ionisation r

o decreases progressively o

from stage 1

• Conjugate acid–base pairs are molecules or ions that differ from each other by an H+ ion.

• A proton or hydrogen ion in solution can be represented by the hydronium ion, H

3

KEY QUESTIONS

1 An acidic solution is formed when hydrogen bromide

gas (HBr) is mixed with water (H

2

• Amphiprotic substances can act as either acids or bases, depending on the substance with which

simply as H+(aq).

P

g

O+(aq), e or

• • P they A proton The a triprotic least greater polyprotic first are extent.

to acid, reacting. ionisation extent a base.

acid the than can third of the donate a ionisation diprotic second more acid ionisation. occurs than occurs one

to the In to

a

2 3 for In conjugate H For base a 2 SO the HF(aq) this each 4 that (aq) following reaction. equation, is + pairs? + OH−(aq) written HNO

reaction, 3 a (aq) give in → bold. → H the what 2 HSO

conjugate 4

are − aq) O). the + Write of two H 2 the NO

acid–base an 3 acid +(aq)

equation

or

5 Show, using an equation, that the reaction between

a solution of sodium hydroxide and a solution of hydrochloric acid is a Brønsted–Lowry acid–base reaction. 6 Write an equation to show each of the following

species acting as an acid and as a base with water. a HCO

3

−

b HPO

4

2−

O(l) + F−(aq)

c HSO

4

−

b HCOOH(aq) + H

**2**

**O(l) → H**

3

O+(aq) + HCOO−(aq)

**d H**

2

O c CH

3

NH

2

(aq) + HCl(aq) → CH

**3**

**NH**

**3**

+(aq) + Cl−(aq) 4 Write the formulae of the conjugate acids of the

following bases. a NH

3 b CH

3

COO− c HPO

4

2−

d CO

3

2−

AREA OF STUDY 2 | ACIDS AND BASES

Triprotic acids Triprotic acids can donate three protons. Phosphoric acid (H

3

PO

4

**4.2 Strength of acids and bases**

The acid solutions in the two beakers shown in Figure 4.2.1 are of equal concentration, yet the acid in the beaker on the left reacts more vigorously with zinc than the acid on the right. The acid on the left is described as a stronger acid than the acid on the right.

r FIGURE acid proton the • • In strong strong differences solutions section 4.2.1

donors and and Zinc are weak weak 4.1, of between: reacts and equal you bases acids bases. more concentration learnt vigorously as proton that and with the acceptors. volume.

a strong Brønsted−Lowry acid In (left) this than section, with theory a weak you P

defines acid will (right). investigate acids The

as

ACID AND BASE STRENGTH Experiments show that different acid solutions of the same concentration do not have the same concentrations of hydronium ions.

Some acids can donate a proton more readily than others. The Brønsted–Lowry theory describes the strength of an acid as its ability to donate hydrogen ions to a base. The strength of a base is a measure of its ability to accept hydrogen ions from an acid.

Because aqueous solutions of acids and bases are most commonly used, it is convenient to use an acid’s tendency to donate a proton to water, or a base’s tendency to accept a proton from water, as a measure of its strength.

Table 4.2.1 gives the names and chemical formulae of some strong and weak acids and bases.

**TABLE 4.2.1**

a

g

e

Strong Hydrochloric HCl

acids P Examples acid,

of common strong and weak acids and bases

**Weak acids Strong bases Weak bases**

Ammonia, NH

3

Sulfuric acid, H

2

Ethanoic acid,

Sodium hydroxide, CH

3

COOH

NaOH

Potassium hydroxide, KOH

Nitric acid, HNO

3

Carbonic acid, SO

4

H

2

CO

3

Phosphoric acid,

Calcium hydroxide, H

3

PO

4

Ca(OH)

2

o

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o

f

s

**94**

a CHEMFILE Differences Acids different current. immersed hydrochloric electric the hydrochloric 1.0 glows mol globe of faintly. L−1 circuit When equal abilities glows in ethanoic P

acid acid a This concentrations containing copper solution brightly. in to and is is conductivity can have conduct an electric electrodes are of 1.0 mol L−1 connected to an a light globe, When the substituted with

acid, the light only further evidence that hydrochloric acid is a stronger acid than ethanoic acid. The brighter light indicates that there are more ions present in a solution of hydrochloric acid than in ethanoic acid.

AREA OF STUDY 2 | ACIDS AND BASES

Strong acids As you saw previously, when hydrogen chloride gas (HCl) is bubbled through water, it ionises completely—virtually no HCl molecules remain in the solution (Figure compounds, reaction protons unreacted are Weak Vinegar molecular (acetate) proportion The Acids the most single easily. is acids is 4.2.2a). ions. that acid compound complete. of a common which solution ethanoic reaction readily In Therefore, molecules a H HNO Similarly, 1.0 also 2 HCl(g) SO donate that strong mol arrow of acid ionise 4 3 (l) (l) remaining. solutions ionises ethanoic L−1 + + + molecules → a acids. pure H H H completely solution proton 2 2

2 in O(l) O(l) O(l) in each HNO of acid. water Hydrochloric → → → are strong of equation H H H 3 ethanoic in Pure to called 3 3

3 O+(aq) O+(aq) O+(aq) and water: produce acids ethanoic above strong H acid, acid + + + 2 contain SO Cl−(aq) HSO NO hydrogen indicates 4 (CH sulfuric acids. 3 acid o −(aq) are 4 −(aq)

3 ions, Strong covalent is acid ions that a with polar and and the f acids virtually nitric molecular ionisation

ethanoate covalent donate s acid no

A molecules 1.0 H (double) 3

1.0 O+ The mol mol are L−1 partial arrows: only L−1 and solution CH

solution about ionisation 3 only COOH(aq) of 0.004 some of ethanoic P

ethanoic of mol hydronium a + r

are ionised o

at any one COOH), only a small time (Figure 4.2.2b). acid contains ions a high proportion and ethanoate ions. At of CH

25°C, 3

COOH in a

L−1.

acid, the concentrations of CH

3

COO−(aq) and

weak acid is shown in an equation using reversible

g

reversible are a The Therefore, present e

double and in solution ethanoic the reaction ethanoic and acid arrow the acid is described reaction H

H H

 2

3 partially O(l) 2

O(l) in  H

3

O+(aq) + CH

3

COO−(aq) the equation above indicates that reaction is ionises. All reactant and product species is in a state of equilibrium.

as a weak acid in water (Figure 4.2.2).

**b**

O(l)

CH

3

**FIGURE 4.2.2**

H

2

O+(aq)

Cl–(aq)

COO–(aq)

H

3

O+(aq)

CH

3

COOH(aq)

(a) In a 1.0 mol L−1 solution of hydrochloric acid, the acid molecules are completely ionised in water. (b) However, in a 1.0 mol L−1 solution of ethanoic acid, only a small proportion of the ethanoic acid molecules are ionised.

Strong bases The ions ionic compound (Na+) and oxide sodium ions (O2−). oxide The (Na

oxide 2

O) dissociates ions react completely in water, releasing with the sodium water, accepting a proton to form hydroxide ions (OH−):

O2−(aq) base + H

acid

2

O(l) → OH−(aq) + OH−(aq)

The oxide ion is an example of a strong base. Strong bases accept protons easily.

Sodium hydroxide (NaOH) is often referred to as a strong base. However, according to the Brønsted–Lowry definition of acids and bases, it is more correct to say that sodium hydroxide is an ionic compound that is a source of the strong base OH−. That is, it is the OH− that is acting as the base; the sodium ion is not involved in the acid–base reaction. Weak bases Ammonia is a covalent molecular compound that ionises in water by accepting a proton. This ionisation can be represented by the equation:

NH

3

(aq) + H

2

O(l)  NH

4

+(aq) + OH−(aq) Ammonia behaves as a base because it gains a proton. Water has donated a proton and so it behaves as an acid.

Only a small proportion of ammonia molecules are ionised at any instant, so a 1.0 mol L−1 solution of ammonia contains mostly ammonia molecules together with a smaller number of ammonium ions and hydroxide ions. This is shown by the double arrow in the equation. Ammonia is a weak base in water.

Relative strength of conjugate You will recall from section 4.1 (H+). The ionisation of HF can HF(aq) + be that OH−(aq) represented conjugate  H

acid−base acids 2 by the and equation:

bases pairs

differ by one proton r

The relative strength of some conjugate acid–base pairs

P

O(l) + F−(aq) F− pair a acid–base Acid The of base equilibrium are HF The in strength a is, this is ionisation stronger conjugate the the pairs reaction.

conjugate weaker when of is an an shown acid–base acid the is acid constants its acid substance is, in conjugate or Figure the of pair. a F− weaker base donates and H 4.2.3.

2 acid. O can OH− is and its The be a hydrogen is conjugate OH− described the relative are conjugate another ion base. strength in to, terms Similarly, or base conjugate accepts of P of some of H the the 2 a O. acid–base conjugate

hydrogen stronger

position HF and

CHEMFILE

Super acids Fluorosulfuric acid (HSO

3

ion from, water. This type of reaction with water is called a hydrolysis reaction.

Ethanoic acid (CH

3

g n

Acid

HCl

n

Base

o r t s

H

2

SO

4

h t g n e r t s d i c a g n

HNO

3 H HSO 3

O+

4

−

k

H

3

PO

4 HF

i s a e r c n I

a e w o e l b i g i l g e

CH H H H H OH− o 2

2

2

2

CO S PO O 3

COOH

4 3

−

Cl− HSO NO H SO H F− CH HCO HS− HPO OH− O2 f 2

2

O

PO

−

4

3 3

2−

COO− −

4

4 3

−

4

2− −

− s h t g n e r t s e s a b g n i s a e r c n I

**FIGURE 4.2.3**

F) is one of the strongest acids known. It has a similar geometry as the sulfuric acid molecule

constant CH

In this 3

COOH(aq) for equation, the hydrolysis + ethanoic H

COOH) 2 O(l) reaction  acid is H

g classified is 3 O+(aq) is behaving small:

+ as CH e as a weak 3

a COO−(aq) Brønsted–Lowry acid because K a the = acid. 1.8 equilibrium × K

10−5

a

(Figure 4.2.4). The highly electronegative fluorine atom causes the oxygen– hydrogen bond in fluorosulfuric acid to be more polarised than the oxygen– hydrogen bond in sulfuric acid. The acidic proton is easily transferred to a base.

O

O

symbol equilibrium ethanoic given acid, constant to the the acidity equilibrium a is called constant K

a the constant acidity expression for constant. a would hydrolysis be For written the reaction hydrolysis as:

is the

HO

S

O

F

S

O

of an acid. This

HO

HO reaction of

**FIGURE 4.2.4**

Structure of sulfuric acid (left) and fluorosulfuric acid (right) molecules

=

[H

3

[CH O

+

][CH

3

COOH] 3

COO

−

]

Fluorosulfuric acid is classified as a super acid. Super acids are acids that have acidity greater than the acidity of Note that the concentration of water does not appear in the denominator of

pure sulfuric acid. the expression for virtually constant.’

*K*

a

. This is because water is the solvent and its concentration is

Super acids such as fluorosulfuric acid and triflic acid (CF

3

Square brackets [ ] are often used to represent molar concentration.

An equilibrium mixture of ethanoic acid and water contains very few hydronium ions or ethanoate ions compared to the number of ethanoic acid molecules. For example, in a 0.1 mol L−1 ethanoic acid solution, approximately 99% of the ethanoic acid molecules are un-ionised.

Ammonia is classified as a weak base because the equilibrium constant for its hydrolysis reaction is small:

NH

3

1000 times stronger SO

than 3

H) are sulfuric

about

acid. a million Carborane times stronger acid (H(CHB

than 11

sulfuric

)) is

acid. The strongest known super acid is fluoroantimonic acid (H

2

Cl

11

which is 1016 times stronger FSbF

than 6 ),

100% sulfuric acid. Super acids are used in the production of plastics and high-octane petrol, in (aq) + H

2

O(l)  NH

4

+(aq) + OH−(aq) K

b

= 1.8 × 10−5

coal gasification and in research.

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**96**

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P

a

In this equation, ammonia is behaving as symbol given to the equilibrium constant for a a Brønsted–Lowry hydrolysis reaction base. of a base. K

b

is the This equilibrium constant is called the basicity constant. (It is only a ammonia are the coincidence that same. These values the indicate values that for the K a

extent for ethanoic acid of hydrolysis and in solution

*K*

b

for

is approximately the same for both compounds.)

Table 4.2.2 lists the hydrolysis reactions for several weak acids and bases, together with the equilibrium constant values for these reactions.

**TABLE 4.2.2**

Some weak acids and bases and their aqueous solution equilibria. The size of the equilibrium constant indicates the extent of the reaction; strong acids and bases have larger equilibrium constants that weak acids and base

**Weak acid or base**

**Hydrolysis equation Equilibrium**

**constant**

Phosphoric acid H

3

PO

4

(aq) + H

2

O(l)  H

3

O+(aq) + H

2

PO

4

−(aq) K

a

= 7.5 × 10−3

Carbonic acid H

2

CO

3

(aq) + H

2

O(l)  H

3

O+(aq) + HCO

3

−(aq) K

a

= 4.3 × 10−7

Ammonium ion NH

4

(aq) K

a

= 5.6 × 10−10

Carbonate ion CO

3 Ethanoate Strength When +(aq) + H 2

O(l)  H

3 O+(aq) o + NH

3

o

f

s

‘strong’ the amount acid is referring a and strong ion ‘weak’ of versus acid to acid solutions with CH or because 3 2−(aq) base ‘concentrated’ concentration of + dissolved it H acids 2 readily O(l) and  r and in HCO donates bases, a ‘dilute’. 3 given −(aq) COOH(aq) it protons. is + volume important OH−(aq) Concentrated + A OH−(aq) of concentrated not solution. to and confuse K

*K*

b

b dilute = = Hydrochloric 1.8 5.6 × 10−4 × 10−10

the terms describe

solution of hydrochloric acid can be prepared by bubbling a large amount of hydrogen chloride into a given volume of water. By using only a small amount of hydrogen chloride, a dilute solution of hydrochloric acid would be produced.

However, in both cases, the hydrogen chloride is completely ionised—it is a strong acid. Similarly, a solution of ethanoic acid may be concentrated or dilute. However, because it is partially ionised, it is a weak acid (Figure 4.2.5).

weak, concentrated ethanoic acid

COO−(aq) + H

2

O(l)  CH

3

g

strong, dilute hydrochloric acid

acid (CH

3

weak, dilute

strong, concentrated ethanoic acid

hydrochloric acid

COO–(aq), Cl–(aq))

**FIGURE 4.2.5**

e

P

COOH, HCl) conjugate base (CH

3 H+(aq) The concentration of ions in an acid solution depends on both the concentration and strength of the acid.

Terms such as ‘weak acids’, ‘strong acids’, ‘dilute solutions’ and ‘concentrated solutions’ are qualitative (or descriptive) terms. You can describe a solution accurately and quantitatively by stating concentrations in mol L−1 or g L−1.

**4.2 Review**

SUMMARY

• A concentrated acid or base contains more moles of solute per litre than a dilute acid or base.

• In the context of acids and bases, the terms ‘strong’ and ‘weak’ refer to the relative tendency to accept or donate protons. - A strong acid donates a proton more readily

than a weak acid. - A strong base accepts a proton more readily

than a weak base.

KEY QUESTIONS

1 Write balanced equations to show that, in water:

a HClO

4

) indicates the strength of the acid.

P

is a strong acid

• • 4 Perchloric Given o

s

b HCN is a weak acid c CH

3

P 5 expect why. Write the a b c d NH

2

is a weak base. 2 Write balanced equations for the three ionisation

stages of arsenic acid (H

3

AsO

4

). 3 Which one of the following equations represents the

reaction of a strong acid with water?

The base. conjugate The ionisation Hydrofluoric Hypochlorous following stronger value expressions 1.0 to The r mol be acid stronger of of acid. a weak L−1 the an an better is acid o acid solutions, acid a acid, equilibrium for acids. stronger a (HF) conductor the (HClO) in base, the aqueous equilibrium weaker which the acid constant weaker of than solution acid is electricity? its constant ethanoic would for is conjugate f its

(K the

a

you

Explain

*(K*

acid.

a

) for

A HNO

3

(aq) + H

2

O(l) → H

3

O+(aq) + NO

3

−(aq) B Hydrogen sulfide (H

2

S) Ammonium ion (NH

4

+) C D HF(aq) LiOH(s) NH

3 (aq) + → + H

H Li+ 2 a O(l) 2 O(l) (aq)   + H

OH−(aq) NH 3 O+(aq) g 4 +(aq) + + F−(aq)

OH−(aq)

e

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**98**

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P

a

g

**4.3 Acidity of solutions**

In previous sections, you were introduced to the Brønsted−Lowry theory that defines an acid as a proton donor and a base as a proton acceptor. You also learnt that acids and bases can be classified as strong or weak, depending on how well they donate or accept protons.

In this section, you will learn about the pH scale, which is a measure of acidity. You will also learn about the relationship between the concentration of hydronium and hydroxide ions in different solutions.

IONIC PRODUCT OF WATER Water molecules can react with each other to form the equilibrium represented by the equation:

The production H

2

O(l) + H

2

O(l)  H

3

O+(aq) + OH−(aq)

in Figure 4.3.1.

of the H

3

O+ ion and OH− ion within this reaction can be seen

+ +

H

2

O H

2

O

H

3

O+ OH–

**The FIGURE 4.3.1**

ionisation of water molecules

Pure water undergoes this self-ionisation to a very small extent. In this reaction, water behaves as both a very weak acid and a very weak base, producing one hydronium amphiprotic at present and 25°C The Experimental OH− e in the concentration ions, a ion (H 3

O+) for every one hydroxide ion (OH−). Water is displaying properties.

of hydronium and hydroxide ions is very low. In pure water H glass 3

O+ of and water, OH− there concentrations are 560 million are each H 2

O 10−7 molecules!

mol L−1. For each H

3

O+ ion

evidence shows that all aqueous solutions contain and that the both H

3

O+ is always 1.00 product of their molar concentrations, [H

3

O+][OH−], increases, then × the 10−14 concentration at 25°C. If of either the other [H 3

O+] must or decrease [OH−] in an aqueous solution

proportionally. Remember that [OH−] represents the [H

concentration 3 O+] represents the concentration of of hydroxide ions. The expression is known as ionic product (or ionisation constant) of water hydrogen ions and

and is [H

represented

3

O+][OH−]

by the symbol K

w

:

*K*

w

= [H

3

O+][OH−] = 1.00 × 10−14 at 25°C

Acidic and basic solutions In water, be constant, solutions greater as well the than of as concentration acidic from 10−7 mol substances, self-ionisation L−1 of at OH− 25°C. H

3 of O+ ions water. Because ions in an are So acidic the formed the product concentration solution by reaction [H

3 at O+][OH−] of this of H

the temperature 3 O+ acid ions remains with will

must be less than 10−7 mol L−1.

The opposite is true for basic solutions. The concentration of OH− ions in a basic solution is greater than 10−7 mol L−1 and that of H

3

O+ ions is less than 10−7 mol L−1. In summary, at 25°C:

• pure water and neutral solutions: [H

3

O+] = [OH−] = 10−7 mol L−1

• acidic solutions: [H

3

O+] > 10−7 mol L−1 and [OH−] < 10−7 mol L−1

• The basic higher solutions: [H

3

O+] < 10−7 mol L−1 and [OH−] > 10−7 mol L−1

solution is.

the concentration of H

3

O+ ions in a solution, the more acidic the

P

r

o

o

f

s

**Calculating concentration of H**

**3**

O+ in aqueous solutions The and hydroxide expression for K

w

can be used to determine the concentrations of hydronium at 25°C is 1.00 ions × 10−14.

in solution, because we know that the value of K

w

in solutions

**Worked example 4.3.1**

**CALCULATING CONCENTRATION OF HYDRONIUM AND HYDROXIDE IONS IN AN AQUEOUS SOLUTION**

For a 0.10 mol L−1 HCl solution at 25°C, calculate [H

3

O+] and [OH−].

**Thinking Working**

Find the concentration of hydronium (H

3

HCl is a strong acid, so it will ionise O+) ions.

completely in solution. Each molecule of HCl donates one proton to water to form one H

3

O+ ion: HCl(aq) + H

2

O(l) → H

3 O+(aq) Because water, solution ions i.e. [H

Use the expression for the ionisation constant of water to calculate the

*K*

w = concentration of OH− ions. [OH −

of [H

]

3

= 0.10 O+] 3 0.10 1.00 O+][OH−] with HCl mol mol is a concentration L−1 completely L−1:

HCl will produce ionised + of Cl−(aq)

H

3

O+ in r a

Worked CALCULATING AQUEOUS pH: Definition The called proposed levels For Alternatively, range a A of the 5.6 CONVENIENT acidity. P example pH by × SOLUTION of 10−6 the H scale, CONCENTRATION of The 3 O+ mol this Danish pH concentrations 4.3.1: pH a has L−1 expression HNO of been scientist a Try WAY solution 3 developed solution pH yourself can g [H Sören OF TO in = 3 be O+] −log is HYDRONIUM solutions at rearranged MEASURE defined Sörenson Since to = 25°C, 10 10−pH [H measure e

3

[H

O+] calculate [OH−] as: is 3 = 0.10 mol L−1

= 1.00 ×

10

−

14

[H 3

O +

]

O+] = 0.10 mol = 1.00 0.10

×10−14

= 1.0 × 10−13 AND HYDROXIDE [H 3 O+] ACIDITY

so great that acidity. The in 1909 as a to give:

P L−1 × mol 10−14 L−1

**IONS IN AN**

and [OH−].

a convenient scale, pH scale was first way of expressing

The pH scale eliminates the need to write cumbersome powers of 10 when you describe hydrogen ion concentration. The use of pH greatly simplifies the measurement and calculation of acidity. Since the scale is based upon the negative logarithm of the hydrogen ion concentration, the pH of a solution decreases as the concentration of hydrogen ions increases.

o

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o

f

s

pH of acidic and basic solutions Figure 4.3.2 shows a pH meter, which is used to accurately measure the pH of a solution. Acidic, basic and neutral solutions can be defined in terms of their pH at 25°C.

• Neutral solutions have a pH equal to 7.

• Acidic solutions have a pH less than 7.

• Basic solutions have a pH greater than 7. On the pH scale, the most acidic solutions have pH values slightly less than 0 and the most basic solutions have values of about 14. The pH values of some common substances are provided in Table 4.3.1.

**TABLE 4.3.1**

pH values of some common substances at 25°C

**FIGURE 4.3.2**

A pH meter is used to measure

**Solution pH [H**

**3**

**O+] (mol L−1) [OH−] (mol L−1) [H**

**3**

O+] × [OH−] the acidity of a solution.

1.0 mol L−1 HCl 0.0 1.00 1.00 × 10−14 1.00 × 10−14

Lemon juice 3.0 1.00 × 10−3 1.00 × 10−11 1.00 × 10−14

Vinegar 4.0 1.00 × 10−4 1.00 × 10−10 1.00 × 10−14

Tomatoes 5.0 1.00 × 10−5 1.00 × 10−9 1.00 × 10−14

Rain water 6.0 1.00 × 10−6 1.00 × 10−8 1.00 × 10−14

Pure water 7.0 1.00 × 10−7 1.00 × 10−7 1.00 × 10−14

Seawater 8.0 1.00 × 10−8 1.00 × 10−6 1.00 × 10−14

Soap 9.0 1.00 × 10−9 1.00 × 10−5 1.00 × 10−14

Oven cleaner 13.0 1.00 × 10−13 1.00 × 10−1 1.00 × 10−14

A solution with pH 2 has

1.0 mol L−1 NaOH 14.0 1.00 × 10−14 1.00 1.00 × 10−14

10 times the concentration of hydronium ions as one of pH 3. The solution of pH 2 also has one-tenth the concentration of hydroxide ions as the one of pH 3.

The acidities of solutions are important in a large variety of everyday applications. The high acidity of gastric juices is essential for protein digestion in the stomach. There is a complex system of pH control in your blood because even small deviations from the normal pH range of 7.35–7.45 for any length of time can lead to serious illness and death.

CHEMFILE

pH of toothpaste The pH in your mouth varies, but a healthy pH is above 7. When you eat food containing sugar, it gets consumed by bacteria, which produce acids that lower the pH and cause tooth decay. The following reaction shows what happens when tooth enamel reacts with acid:

Ca

5 Saliva acid demineralisation the The any regular (PO enamel. acids average remains is 4 toothpaste ) 3 naturally OH(s) remaining P pH unaffected + of can H alkaline regular are 3 O+(aq) at begin abrasives the and a and toothpastes time  at lowers a neutralises 5Ca2+(aq) pH of such use. of the 6.5 as is g The pH silica, 8, + some in 3PO below active which the and 4 acid root 3−(aq) 7. neutralises ingredients fluoride, Tooth but and e + excess 2H 5.5 2

which O(l)

at

in

is basic and raises the pH in the mouth. In contrast, whitening toothpastes (Figure 4.3.3) usually contain peroxide bleaching agents in the form of hydrogen peroxide or carbamide peroxide. These compounds are acidic and lower the pH in the mouth. The average pH of whitening products is

**100**

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**FIGURE 4.3.3**

around 6.8 (slightly acidic) but can be as low as 3.7 (highly acidic), depending on the peroxide concentration in the toothpaste. Scanning electron microscope studies have shown that a concentration as low as 10% peroxide in the toothpaste causes surface dissolution of tooth enamel, exposing a porous surface beneath to further decay. Concentrations of up to 35% are found in some whitening toothpastes.

P

Whitening toothpaste containing peroxide

r

o

o

f

s

Indicators One of the characteristic properties of acids and bases is their ability to change the colour of certain plant extracts. These extracts are called indicators. Indicators are used to estimate the pH range of a solution. You will learn more about indicators in Chapter 6. Calculations involving pH pH has can a logarithm be calculated function from that the will formula simplify pH pH = −log

calculations.

10

[H

3

O+]. Your scientific calculator

**Worked example 4.3.2**

**CALCULATING pH OF AN AQUEOUS SOLUTION FROM [H**

**3**

**O+]**

What Express is the your pH answer of a solution to two decimal in which places.

the concentration of [H

3

O+] is 0.14 mol L−1?

**Thinking Working**

Write ions in down the solution.

the concentration of [H

3

O+]

[H

3 O+] Substitute the value of [H

3

O+] into: pH = −log

10

[H

3

O+] Use the logarithm calculator to determine Worked example CALCULATING pH What is the pH of 6.0 × 10−9 mol L−1? Worked example function on the answer. 4.3.2: Try OF AN AQUEOUS a solution in Express your 4.3.3

pH = = your = yourself SOLUTION which the concentration answer −log

−log 0.85 FROM = 0.14 10 10

[H (0.14) [H of mol 3 O+]

**3**

**O+]**

[H

L−1

(use 3 O+] P

your calculator)

r

CALCULATING What Thinking Write Ba(OH) is down 2 the pH the pH of reaction IN a 0.005 a A SOLUTION in mol which

L−1 g solution OF A to BASE

of two e Ba(OH) significant 2

at 25°C?

figures.

is

**Working**

dissociates.

In dissociates water, each to release mole of 2 Ba(OH)

moles 2

of completely OH− ions. Ba(OH)

2

(aq) → Ba2+(aq) + 2OH−(aq)

Determine [OH−] Determine solution ions.

P by the concentration of

[OH−] = 2 × [Ba(OH)

2

] = 2 × 0.005 mol L−1 = 0.01 mol L−1

the substituting [H 3

O+] in the the [OH−]

diluted

into the ionic product of water: K

w

*K*

w

= [H

3

O+][OH−] = 1.00 × 10−14

= [H

3

O+][OH−] = 1.00 × 10−14

[H 3

O +

]

=

[OH K w

−

]

=

1.00 0.01 × 10

− 14

= 1 × 10−12 mol L−1

Substitute the value of [H

3

O+] into: pH = −log

10

pH = −log

10

[H

3

O+] [H

3

O+]

= −log

10 Use the logarithm function on your calculator to determine the answer.

(1 × 10−12) (use your calculator) = 12.0

o

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o

f

s

**102**

AREA OF STUDY 2 | ACIDS AND BASES

P

a

**Worked example 4.3.3: Try yourself**

**CALCULATING pH IN A SOLUTION OF A BASE**

What is the pH of a 0.01 mol L−1 solution of Ba(OH)

2

at 25°C?

**Worked example CALCULATING pH IS NOT GIVEN**

What is the pH of solution?

Thinking Determine the number NaOH. IN a 4.3.4 solution A SOLUTION of moles at 25°C of

WHEN that contains Working

**SOLUTE 1.0 CONCENTRATION**

g NaOH in 100 mL f of

Write the equation m

o NaOH. Determine the number of OH− based on the n(NaOH) n(NaOH) o

=

=

*M*

1.0

equation.

Use the formula for concentration given and c Determine =

*V n*

volume: e the [H

3 for determining number dissociation dissociation of P moles

of moles

of r

NaOH(aq) NaOH water. n(OH−) n V = = is = = 0.025 0.100 40.0

= 0.025 mol

→ Na+(aq) + OH− (aq) completely dissociated in

n(NaOH) 0.025 mol

mol L

*c =*

*V n*

=

0.025 0.100

= 0.25 mol L−1

O+] in the diluted O+][OH−] solution by substituting the [OH−] into

*K*

w

= [H

3

= 1.00 × 10−14 g the K w

ionic product of water: = [H

3

O+][OH−] = 1.00 × 10−14

[H 3

O +

]

=

[OH K w

−

]

−14 =

0.25 = 4.0 × 10−14 mol L−1

Substitute the value of [H

3

1.00 ×10

O+] into:

pH = −log

10

[H

3

O+] pH = −log

10

= −log

calculator) 10

(4.0 × 10−14) (use your

= 13.4

**Worked example 4.3.4: Try yourself**

**CALCULATING pH IN A SOLUTION WHEN SOLUTE CONCENTRATION IS NOT GIVEN**

What is the pH of a solution at 25°C that contains 0.50 g KOH in 500 mL of solution?

[H

3

O+] Use the logarithm function on your calculator to determine the answer.

s

**Worked example 4.3.5 CALCULATING [H**

**3**

**O+] IN A SOLUTION OF A GIVEN pH**

Calculate [H

3

O+] in a solution of pH 5.0 at 25°C.

**Thinking Working**

Decide be used:

which form of the relationship between pH and [H

3

O+] should

pH = −log

10 or [H

3 Substitute calculator [H 3

O+]

O+] = 10−pH

the value of pH into the relationship expression and use a to determine the answer.

**Worked example 4.3.5: Try yourself CALCULATING [H**

**3**

As you have the pH and are calculating [H

3

O+] = 10−pH

O+] = = = 10−5.0 1.0 10−pH

(or × 0.000010 10−5 mol L−1

o mol L−1)

f [H

3 O+], s use:

[H

3

**O+] IN A SOLUTION OF A GIVEN pH**

Calculate [H

3

O+] in a solution of pH 10.4 at 25°C.

Worked example 4.3.6 CALCULATING pH OF A SOLUTION AFTER 50.0 mL of 0.200 mol L−1 NaOH is added Thinking Identify the type of reaction and write an Identify the values for the concentrations reactants.

Calculate Identify (This Identify Calculate Calculate will P the which always the the the relationship number be 1:1.)

of between moles of g

the H+ and number to equation and OH−. PARTIAL 100.0 e

volumes mL for NEUTRALISATION of the of 0.150 the

reaction mol P

L−1 HCl. Working

acid HCl(aq) HCl c V H+ n of moles of H+ and OH−. Calculate r

o

= = = = = the pH of the final solution. + base → salt + water

+ NaOH(aq) → NaCl(aq) + H

2

O(l)

NaOH 0.150 mol L−1 c = 0.200 mol L−1 100.0 mL V = 50.0 mL

OH− number of H+ a

and of c × V 0.150 × 0.100 0.0150 mol n = c × V = 0.200 × 0.050 = 0.0100 mol

OH− mol is of in excess. reactant in excess. 1 mol of H+ reacts with 1 mol of OH−

concentration of the excess H+ (H

3

HCl is in excess.

Number of mole of H+ in excess = 0.015 – 0.010 mol 0.0050 mol

O+). The volume of the final solution will equal the sum of the volumes of

*c =*

*V n*

the solutions mixed together.

=

0.0050 0.150 = 0.333 [H

3

O+] = 0.033 mol L−1

Calculate pH using: pH = −log

10

pH = −log

10

[H

3

O+] [H

3

O+]

= −log

10

0.033 Use the logarithm function on your calculator to determine pH.

= 1.48

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**Worked example 4.3.6: Try yourself**

**CALCULATING THE pH OF A SOLUTION AFTER PARTIAL NEUTRALISATION**

43.0 mL of 0.200 mol L−1 nitric acid is added to 15.0 mL hydroxide (at 25°C). Calculate the pH of the final solution.

Effect of temperature on pH Earlier, the ionisation constant of water was defined as:

*K*

w

= [H

3 O+][OH−] = 1.00 × 10−14 at You can use this relationship to calculate either [H different solutions. have in equilibrium the concentrations TABLE Temperature Table If The pH the shown 4.3.2 ΔH is pH 15 25 35 temperature 0 5 4.3.2. not + The H of reaction: that (°C) equal 2 effect O(l) pure of This the H

of + to water 3 1.14 1.85 4.51 1.00 2.09 O+ value temperature is H is 7.00, not 2 and O(l) K because × × × × × P is w 10−15 10−15 10−15 10−14 10−14 of 25°C, pure only OH−  K w

H 7.00 water increases then ions the 3 O+(aq) r at self-ionisation the are can 25°C. value equal. + as still OH−(aq) o the At be of temperature other K described w will of 3 temperatures, of water O+] 25°C be o ΔH 0.300 different. as or = increases, neutral is [OH−] +56 mol an kJ L−1 Experiments f endothermic

even mol−1 because at barium as 25°C though shown s

the in

on pH of pure water

**pH**

7.47

7.37

P

a

g

e 45 55 7.17 7.00

6.83

4.01 × 10−14 6.70

7.29 × 10−14 6.57

**4.3 Review**

SUMMARY

• Water self-ionises according to the equation: H

2

O(l) + H

2

O(l)  H

3

O+(aq) + OH−(aq)

• The ionic product for water is: K

w

= [H

3

O+][OH−] = 1.00 × 10−14 at 25°C

• The concentration of [H

3

O+] can be measured using the pH scale: pH = −log

10

[H

3

O+]

KEY QUESTIONS

1 Calculate [OH−] at 25°C in an aqueous solution with

[H

3

O+] = 0.001 mol L−1. 2 What is [OH−] in a solution at 25°C with

[H

3

O+] = 5.70 × 10−19 mol L−1? 3 Calculate which [OH−] [H

3 = O+] 1.0 at × 25°C 10−5 mol in an L−1.

aqueous solution in

4 What is the pH of a solution in which

[H

3

O+] = 0.01 mol L−1?

P

a

g

e

5 Calculate the pH of a 0.001 mol L−1 solution of nitric

acid (HNO

3

). 6 The pH of water in a lake is 6.0. Calculate [H

3

O+] in the lake. P 7 8 Determine the pH of a 200 mL solution that contains

0.365 g of dissolved HCl. Determine the pH of the final solution when 100.0 mL of a solution HCl with a pH of 2 is mixed with 200.0 mL of a solution NaOH with a pH of 11.

• At 25°C, the pH of a neutral solution is 7.0. The pH of an acidic solution is less than 7.0 and the pH of a basic solution is greater than 7.0.

r

o

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o

f

s

**106**

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P

a

**4.4 Dilution of acids and bases**

Although acids are frequently purchased to be diluted. For example, a bricklayer to remove mortar splashes from bricks solution is prepared by diluting concentrated In this section, you will learn how to and bases once they have been diluted.

CONCENTRATION OF ACIDS The concentration of acids and bases is is also referred to as molar concentration. The molar concentration of a solution, c

calculate AND = usually used as uses in concentrated mol hydrochloric to a BASES

expressed 10% build the L−1, concentration solution is a given house. solutions, in acid units of by The by hydrochloric the a of and they brick-cleaning factor expression: mol f pH often L−1. of of s ten. acids need

This acid

where c is the molar is the volume of the The most convenient n V

concentration (mol L−1), solution (L).

way of preparing a n solution is the amount of a o

dilute of solute acid is (mol) by mixing and V

concentrated acid with water, as shown in Figure 4.4.1. This is known as a dilution.

g

Measure out 10.0 10.0 mL mol of

L–1 HCl with a pipette.

Transfer to a 100.0 mL volumetric flask partly filled with water. e

P

r

o

Fill the flask with distilled water to the calibration mark and shake thoroughly.

**FIGURE 4.4.1**

Preparing a 1.00 mol L−1 HCl solution by diluting a 10.0 mol L−1 solution. Heat is released when a concentrated acid is added to water, so the volumetric flask is partly filled with water before the acid is added. (Extra safety precautions would be required for diluting concentrated sulfuric acid.)

The amount of solute (in moles) in a solution does not change when a solution is diluted. As the number of moles is given by the following relationship:

n = cV if we change the volume, the number of moles will stay the same, or mathematically:

*c*

1

*V*

1

= n and c

2

*V*

2

= n where concentration c

1

and are V

and 1

are the initial concentration volume after dilution. (So as and the volume volume, of and the solution c

2

and V

increases,

2

the

the concentration decreases.)

This can be simplified to give the relationship:

*c*

1

*V*

1

= c

2

*V*

2

which can be rearranged to:

c

2

= c V

1 V

2

1

to calculate the new concentration of the diluted solution.

To calculate the concentration of a dilute acid you need to know the:

• volume of the concentrated solution (V

1

)

• • concentration total The molar volume concentration of (molarity) water added. of of the some (V

2 concentrated = concentrated initial volume solution acids + volume (c are 1 )

shown of water in Table added)

4.4.1.

**TABLE 4.4.1**

Molar concentrations of some concentrated acids

**Concentrated acid (% by mass) Formula Concentration (mol L−1)**

Ethanoic acid (99.5%) CH

3

COOH 17

Hydrochloric acid (36%) HCl 12

Nitric acid (70%) HNO

3

Phosphoric acid (85%) H

3

PO

4

Sulfuric acid (98%) H

2

SO

4

In the laboratory, you can prepare solutions of a concentration by:

• diluting a more concentrated solution

• dissolving a weighed amount of the base in a measured shown in Figure 4.4.2.

16

15

18

r

**N FIGURE 4.4.2 of water.**

(a) (b) (c) (d) (e) g P Ensure Accurately Add Transfer Dissolver Preparing a C water F complete the the T to weight a base make base a solution transfer to out in the a water.

a volumetric by solution mass b of dissolving the g of up base the flask.

to base. a by the weighed washing calibration c e amount of with water. mark d base and base volume P of of a required water, as

e in a measured volume

shake thoroughly.

o

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o

f

s

**108**

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P

a

**Worked example 4.4.1**

**CALCULATING MOLAR CONCENTRATION AFTER DILUTION**

Calculate the molar concentration of hydrochloric acid when 10.0 mL of water is added to 5.0 mL of 1.2 mol L−1 HCl.

**Thinking Working**

The number of moles of solute does not change during a dilution. So c

1

*c*

1

*V*

1

*= c*

2

*V*

2

concentration V 1

*= c*

2

*V*

2

, in where mol c is the

L−1 and V is the volume of the solution. (Each of the volume units must be the same, litres or millilitres.)

Identify given values for concentrations 10.0 and volumes before and after dilution. volume g Worked CALCULATING Worked CALCULATING Identify Transpose the find Calculate to How solution Thinking 20.0 known the e much mL the required example example to the the values 1.00 of water unknown. molar 5.00 MOLAR THE equation mol value. must into mol 4.4.1: 4.4.2 VOLUME L−1? concentration the CONCENTRATION L−1 be and P equation Try substitute yourself

to r c V V You concentration c 2 1

1

2 = = = = o 1.2 are 5.0 15.0 mL c 1 V × mol 2 mL is required was V mL

1 15.0 L−1

added after mL. o to dilution. calculate to 5.0 mL, f the final

s

*c*

2

, the

= 1.2

× 15.0

5.0

= 0.40 mol L−1

**AFTER DILUTION**

of nitric acid when 80.0 mL of water is added HNO

3

.

**OF WATER TO BE ADDED IN A DILUTION**

added to 30.0 mL of 2.50 mol L−1 HCl to dilute the

**Working**

The number of moles of solute does not change during a dilution.

*c*

1

*V*

1

*= c*

2

*V*

2

So concentration c

1

*V*

1

*= c*

2

*V*

2 , in where mol Identify given values for concentrations and volumes before and after dilution. Identify the unknown.

c is the L−1 and V is the volume of the solution. (Each of the volume units must be the same, although not necessarily litres.)

*c*

1

= 2.50 mol L−1 V

1

= 30.0 mL c

2

= 1.00 mol L−1 You are required volume of the to diluted calculate solution. V

2

, the

Transpose the equation and substitute the known values into the equation to find the required value.

*V*

=

c 1 c ×

2

*V*

1 2

=

2.50 1.00

× 30.0

= 75.0 mL

Calculate the volume of water to be added.

Volume of dilute solution = 75.0 mL Initial volume of acid = 30.0 mL So 75.0 – 30.0 = 45.0 mL of water must be added.

**Worked example 4.4.2: Try yourself**

**CALCULATING THE VOLUME OF WATER TO BE ADDED IN A DILUTION**

How much water must be added to 15.0 mL of 10.0 mol L−1 NaOH to dilute the solution to 2.00 mol L−1?

EFFECT Calculation Consider concentration the 9.0 pH to to bases weak Worked CALCULATING 5.0 diluted Thinking Identify and Calculate concentration transposing c 1 3.0. decrease If Similarly, You V mL pH increases 1 mL this acids volumes = after However, of c of will of 2 solution? P given example V 1.0 a this water, 2 dilution. 0.010 c ON and 0.10 until 2 now , the mL to the which of solution before values pH of bases 2.0. pH mol H progressive when it mol the formula: look of solution H 3 OF reaches O+ 3 (A 4.4.3 A O+ is L−1 OF a concentration L−1 pH and is for A at further ions acids the is after discussion complex DILUTED HNO solution how 1.0. DILUTION concentrations after is of very in diluted are dilution, dilution 3 this an to dilution is dilution. repeatedly close calculate and g diluted of solution of ACID acid of by hydrochloric of the by beyond to H by a a OF 7. 3 to effect factor 0.10 O+ after a is the Working

*c V V c*

c diluted, 100.0 factor 2 1

2 1

2 STRONG 0.10 ions = = = = = = e the mL pH of 0.010 ? 5.0 100.0 c of 1 mol V scope mL. decreases dilution dilution acid. × of NaOH 10 2 of mL the V 1 10 mol mL L−1. to solutions What × pH HCl of to 10.0 ACIDS L−1 solution Since on this 100 cannot is to is mL the the 0.01 course.) a mL of pH strong P pH by pH will increase strong AND will mol = the of of −log cause increase L−1 solutions acid, the addition acids BASES 10 above [H and the r so 3

O+],

and pH

pH the

the of

of 7.

0.010 5.0 100.0 = 0.00050 mol L−1

Calculate pH using: pH = −log

10

pH = −log

10

[H

3

O+] [H

3

O+]

= −log

10 Use the logarithm function on your calculator to determine pH.

(0.00050) = 3.30

o

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o

f

s

**110**

Concentration of OH– in original solution

Concentration of OH– in diluted solution

Concentration of H

3

solution O+ pH solution of in diluted P pH K diluted

w = = = – 1.00 [H log 3

O+][OH–] = 10 × c a 2 10–14 V 2

mol2 L–2

**FIGURE 4.4.4**

Calculating the pH of a diluted solution of a base

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Use c

1

*V*

1

[H

3

O+]

**Worked example 4.4.3: Try yourself**

**CALCULATING pH OF A DILUTED ACID**

10.0 mL of 0.1 mol L−1 HCl is diluted to 30.0 mL. Calculate the pH of the diluted solution.

CHEMFILE

Transporting sulfuric acid Transportation of sulfuric acid can be very dangerous because the acid is extremely corrosive. Rather than transport it as diluted acid, which would seem safer, it is transported in the concentrated form and diluted at its destination. Concentrated sulfuric acid is 98% pure acid and is almost completely molecular (Figure 4.4.3). It is not ionised and so does not react with iron or steel in the way that it would in the dilute (aqueous) form.

2H+(aq) + Fe(s) → Fe2+(aq) + H

2

s

Therefore, is stronger or concentrated FIGURE preferable transporting e 4.4.3 than sulfuric The to acid glass the transport structure is acid acid, or then plastic can care the of diluted P be molecular acid and must stored at therefore in be its HO the H in taken destination.

2 r steel concentrated HO safer. O

S to containers avoid o This O

form also and (g)

means in transported this o way that, because when in tankers. f storing

steel It is

contamination with water. The

SO

4

When diluting a concentrated solution of sulfuric acid, the acid is always added to water, not water to the acid. When molecular sulfuric acid reacts with water, it hydrolyses in two stages. The first hydrolysis reaction is in equilibrium but the value of K is very high, g

so exothermic. the that dilution protonation process:

These two of water H factors 2

is strongly favoured. The hydrolysis reaction is also highly mean that large quantities of heat are produced during

SO

4

(l) + H

2

O(l) → HSO

4

−(aq) + H

3

O+(aq) HSO

4

O+(aq) If you add water to the acid, you form an extremely concentrated solution of hydrolysed acid initially and enough heat can be released that the solution may boil violently. Addition of a small quantity of the concentrated acid to water ensures the acid is the limiting reagent and the small quantity of heat released is dissipated through the solution. This is a good general rule to apply to any dilution procedure.

Calculation of pH of a base after dilution The following steps and flow chart shown in Figure 4.4.4 show the sequence used to calculate the pH of a base after it has been diluted. Remember that pH is a measure of the hydronium ion concentration. 1. Calculate [OH−] in the diluted solution. 2. Calculate [H

3

−(aq) + H

2

O(aq)  SO

4

2−(aq) + H

3

O+], using the expression for the ionic product of water: K

w

= [H

3

O+][OH−] = 1.00 × 10−14 3. Calculate the pH of the solution, using pH = −log

10

[H

3

O+] (Figure 4.4.4).

**Worked example 4.4.4**

**CALCULATING pH OF A DILUTED BASE**

10.0 mL of 0.1 mol L−1 NaOH is diluted to 100.0 mL. Calculate the pH of the diluted solution.

**Thinking Working**

Identify given values for concentrations and volumes before and after dilution.

*c*

1

= 0.1 mol L−1 V

1

= 10.0 mL V

2

= 100.0 mL c

2

= ?

Calculate dilution, by c

2

, transposing which is [OH−] after

the formula: c

1

*c*

2

=

c 1 V ×

2

*V*

1

*V*

1

*= c*

2

*V*

2

=

0.1 100.0

× 10.0

= 0.01 mol L−1

Determine solution by [H

substituting 3

O+] in the [OH−] diluted

into the ionic product of water: K

w

*K*

w

= [H

3

O+][OH−] = 1.00 × 10−14

[H 3

O +

]

=

[OH K w

−

] = [H

3

O+][OH−] = 1.00 × 10−14

=

Calculate pH using: pH = −log

10

[H

3

O+] Use the logarithm function on your calculator to determine pH.

1.00 × 10

− 14

0.01

= 1 × 10−12 mol pH = −log 10

[H

3

O+] = −log

10

(1 × 10−12)

Worked CALCULATING 15.0 diluted mL solution. P example of 0.02 pH mol OF 4.4.4: a

L−1 A DILUTED KOH Try is diluted yourself g BASE

to 60.0 e = mL. 12.0 Calculate the P L−1

pH of the

r

o

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o

f

s

**4.4 Review**

SUMMARY

• Amounts of acid or base in solution do not change during a dilution. The volume of the solution increases and its concentration decreases.

• Solutions of acids and bases of a required concentration can be prepared by diluting more concentrated solutions, using the formula:

*c*

1

KEY QUESTIONS

1 1.0 L of water is added to 3.0 L of 0.10 mol L−1 HCl. What is the concentration of the diluted acid? 2 How much water must be added to 10 mL of a

2.0 mol L−1 sulfuric acid solution to dilute it to 0.50 mol L−1? 3 What volume of water must be added to dilute 20.0 mL

of 0.600 mol L−1 HCl to 0.100 mol L−1? 4 Describe the effect on the pH of a strong monoprotic

acid solution of pH 0.1 when it is diluted by a factor of 10. 5 Calculate the pH of the solution at 25°C that is formed

by the dilution of a 20.0 mL solution of 0.100 mol L−1 NaOH to 50.0 mL.

• The pH of a diluted acid can be determined by calculating the concentration of hydronium ions in the diluted solution.

• The pH decreases when a solution of a base is diluted.

• The pH of a diluted base can be determined by calculating the concentration of hydroxide ions in

where c

1

the diluted base, and using the ionic product of water to calculate the concentration of hydronium ions in the diluted base.

P

*V*

1

*= c*

2

*V*

2 and V

1

are the initial volume and concentration, and c

2

and V

2

are the final volume and concentration after dilution.

• The pH increases when a solution of an acid is diluted.

6 P For the: a b c d each i concentration ii concentration iii pH. of the r solutions of of o H OH− 3

O+ a–e ions ions (all o

f

at 25°C), calculate

a

g

e

0.001 mol L−1 HNO

3

(aq) 0.03 mol L−1 HCl(aq) 0.01 mol L−1 NaOH(aq) 1.0 × 10−4.5 mol L−1 HCl(aq) e 0.005 mol L−1 Ba(OH)

2

(aq)

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s

**4.5 pH of salt solutions**

In section 4.2, you were introduced to the concept that some ions act as Brønsted– Lowry acids and bases in hydrolysis reactions. For example, the ammonium ion can donate a proton to water and act as a Brønsted–Lowry acid:

NH

4

+(aq) + H

2

O(l)  H

3

O+(aq) + NH

3

(aq) The ethanoate ion can accept a proton from water and act as a Brønsted–Lowry base:

CH

3

COO− (aq) + H

2

O(l)  CH

3

COOH(aq) + OH− (aq) In this section, you will learn how to predict the acidity of the salt solutions that are formed during neutralisation reactions.

SALTS A salt is an ionic compound that contains a positive ion (a cation) other than a hydrogen ion (H+) or a negative ion (an anion) other than the oxide ion (O2−) or the ammonium base. positive base (H

anions ions. example, with produce a acid. of hydrolysis.) SALT Salts hydrogen 2 sodium Salts For Ammonium A Sodium In In hydroxide CO water potassium This A The are dissociation this a produced 3 salt the SOLUTIONS ions ) solution are the P formed the reacts reaction formation reaction in formed salt carbonate carbonate nitrate formed ion hydrogencarbonate CH are salt the ion hydroxide potassium with carbonate from 3 derived of following COO−(aq) CO by ammonium may reaction (OH−). NH H of are in KOH(aq) the sodium in a 2 a ammonia, 3 CO an of this 2−(aq) 4 water dissociates reaction Na neutralisation then examples +(aq) is carbonate from hydroxide ion FROM 3 2 basic. Sodium type (aq) CO and takes chloride, ((NH

way: react carbonate, to + with + + ethanoate 3 the + H H H the (s) of + form between a HCl(aq) place ion (The 2 2

2 4 of g 2NH O(l) base: O(l) O(l) in ) reaction with water base NEUTRALISATION fluoride, → chloride 2 ion CO salts. water ions and the reactions the 2Na+(aq) when 3    Na+ water acts and 3 (aq) the is ) dissociates potassium an a hydrogencarbonate in → is HCO NH CH according an hydroxide can calcium ions negative carbonate as acid ions a the → the KCl(aq) to e example 3 salt 3 by COOH(aq) be (aq) a (NH 3 + produce from salt −(aq) reaction Brønsted–Lowry and in the thought CO dissolves + carbonate, ions in 4 ions the produced to ) combination ion: 3 a H hydrochloric + 2

2−(aq) + ion of aqueous CO base. the 3 H

OH−(aq) O+(aq) solution may from either a 2 explains of O(l) 3 undergoes following hydrolysis + (aq) in REACTIONS as For OH−(aq) water. ion have the iron(III) when a solution H P compound example, 3 as do of acid. O+ base, acid: why come The the equation: an carbonic not hydrolysis reaction. ions sulfate acid and a conjugate accepting cations from undergo solution or sodium r whose with reacts OH− acid and

For the

or

to a

chloride is formed by the reaction between hydrochloric acid and sodium hydroxide:

HCl(aq) + NaOH(aq) → NaCl(aq) + H

2

O(l)

o

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o

f

s

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P

a

A solution of NaCl is neutral, i.e. its pH is 7. However, some salts, such as ammonium chloride, form acidic solutions when dissolved in water. Other salts, such as sodium ethanoate, produce basic solutions. The pH of the solution produced during reacted. Salts Sodium Sodium strong solution A a neutralisation reaction depends on the strength of the acid and of a strong acid and strong bases

nitrate is the salt of the strong acid, HNO

3 , and the strong HNO

3

(aq) + NaOH(aq) → NaNO

3

(aq) + H

2

O(l) ions do not acid, therefore HNO

3 has , are a salt produced in base s

being

**neutral salt.**

The anion and cation base, f

NaOH:

donating a proton.

Salts of a strong Solutions of salts of a strong acid of the weak base hydrolyses Ammonium chloride base (NH

3

):

very the pH hydrolyse acid (NH of reaction of weak a acid 4 7. neutral and to and bases in form between water. a a salt and weak weak hydronium do do Nitrate base a o not not strong base react are react ions, ions. acidic acid with with the o and because water conjugate water. a base of the Sodium nitrate

strong base is a

by accepting or

the conjugate

An acidic solution The ammonium ion (NH

3 ), undergoes a is HCl(aq) P formed Cl) + when NH

is r the 3 (aq) ammonium salt → of NH

a strong 4

Cl(aq) chloride acid is (HCl) dissolved and the in weak water. (NH hydrolysis 4

+), which reaction:

is the conjugate acid of the weak base, ammonia The e NH

4

NH

4

+(aq) + H

2

O(l)  NH

3

(aq) + H

3

O+(aq) + ion acts as a Brønsted–Lowry acid and donates a proton to water to form the hydronium ion. The ammonium ion is a weak acid and the ionisation reaction with water does not go to completion. The Cl− ion shows no acid–base g properties acidic Salts Solutions hydrolyse the (NaOH): A The weak salt salt. anion of acid produced of to because a form salts or weak ethanoic cation of hydroxide it in does a not hydrolyse in water.

the reaction between a strong acid and a weak base is an

of an acidic salt reacts with water by donating a proton. acid and a strong base weak acid and a strong base are basic because the anions

acid (CH

ions. 3

COOH) Sodium and ethanoate the strong (CH

3 base COONa) sodium is the hydroxide salt of

CH

3

COOH(aq) + NaOH(aq) → CH

3

COONa(aq) + H

2

O(l) Sodium ethanoate dissociates in water Na+ ions do not hydrolyse and show no to form Na+ ions and CH

3

COO− ions. The is a Brønsted–Lowry base and accepts a acid–base proton from properties. a water molecule The CH

3 to COO− produce ion

a hydroxide ion: CH

3

COO−(aq) + H

2

O(l)  CH

3

COOH(aq) + OH−(aq) Consequently the solution is basic. A salt produced in the reaction between a weak acid and a strong base is a basic salt.

The anion or cation of a basic salt reacts with water by accepting a proton.

Salts of a weak acid and a weak base A solution of a salt formed from the reaction of a weak acid and a weak base involves two competing hydrolysis reactions.

• The conjugate base of the weak acid reacts with water to form hydroxide ions.

• The conjugate acid of the weak base reacts with water to form hydronium ions. The pH of a solution of a salt of a weak acid and a weak base will depend upon the extent to which each of these competing reactions occurs. For example, ammonium ammonia (NH 3

) and the weak hypochlorite acid hypochlorous (NH 4

ClO) acid (HClO):

is a salt of the weak base

NH

3

(aq) + HClO(aq)  NH

4

+(aq) + ClO−(aq) Two hydrolysis reactions take place when ammonium hypochlorite is dissolved in water:

NH ClO−(aq) 4

+(aq) + + H H

2 2 O(l) O(l)   HClO(aq) NH

3

(aq) + + H

OH−(aq) 3

O+(aq) K K a

b

= 5.6 = 3.3 × 10−10 × 10−8 An indication of the extent to which each of these reactions occurs is provided by ClO− NH

therefore produced. summarised TABLE Strong Weak the The The 4 ClO 4.5.1 occurs base value base acid–base acidity is depends Acid–base basic to of in Table a the of greater because properties on salts equilibrium properties Strong Neutral Acidic 4.5.1.

the produced extent salts more acid degree salts of of salts than constants, OH− the of derived by the salts hydrolysis ions the hydrolysis from formed are K, reaction neutralisation for produced of these by of the of reactions NH a anion reactions. than weak 4

reactions

+. Therefore, H and of acid 3 O+ P acids The cation ions. and hydrolysis a and solution weak in bases the r base salt of of

is

o

o

f

s

As can be seen from the information the pH of the solution of the salt common ions are summarised in TABLE 4.5.2

to Table be predicted. 4.5.2.

above, Weak Basic Depends e salts acid

it The on is the extent acid–base ions of hydrolysis in the properties salt of that each of enables

ion

some

**Neutral strong Cl− Br− NO**

3

−

**acids/bases) (derived P Acid–base from**

properties a of Acidic weak NH HSO H

2 common PO 4

+

4 g −

4

**bases)**

**− (derived ions**

**from**

**Basic (derived from weak acids)**

CO HCO SO HPO

4 3

2− 2−

3

−

PO

2− 4

F−

3− 4

CHEMFILE

Urinary alkalinisers Urinary alkalinisers form part of a range of treatments for urinary tract infections and kidney stones. Urinary alkalinisers are salts that form solutions of weak bases when dissolved in water. Taken orally, they are rapidly absorbed into the bloodstream, raising the pH of the blood. They also raise the pH of urine, providing relief for the sufferer. The active ingredients are usually potassium hydrogencarbonate (KHCO

3

The anions of these salts act as Brønsted–Lowry bases, accepting protons from water, and producing a basic (alkaline) solution:

HCO

3

−(aq) + H

2

O(l)  H

2

CO

3

(aq) + OH−(aq) C

6

H

5

O

7

3−(aq) + 3H

2

O(l)  C

6

H

8

O

7

(aq) + 3OH−(aq)

) and potassium citrate (K

3

C

6

H

5

O

7

)

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P

a

**Worked example 4.5.1**

**DETERMINING THE ACIDITY OF A SALT PRODUCED IN A NEUTRALISATION REACTION**

Determine and Thinking Write neutralisation Identify in Determine the sulfuric the salt.

the equation the whether anion acid.

reaction. acidity and for the of the

cation

ions the salt Working

produced by the reaction of sodium hydroxide

s

will react with water in a hydrolysis reaction. (Are the

2NaOH(aq) Na+ Na+ and cannot SO 4 accept + 2− H

2

SO

4

(aq) → Na

2

SO

4 (aq) f + 2H

2

O(l)

ions neutral, acidic or basic?)

If an ion accepts a proton from water, it is basic.

a proton o from water

If an ion donates a proton to water, it is acidic. Worked example: (Na+ SO (SO 4

4 2−(aq) 2− is a neutral + H

o 2

O(l) ion)

 HSO

4

−(aq) + OH−(aq) DETERMINING NEUTRALISATION The salt r is is a basic.

basic ion)

Determine the (Ca(OH) e 2 ) and THE oxalic acidity REACTION

**ACIDITY Try acid of yourself P the (H**

2 salt OF C 2

O A produced 4 4.5.1

**SALT PRODUCED IN A**

by the reaction of calcium hydroxide ).

Hydrolysis of salts of polyprotic acids Aqueous solutions of the salts of polyprotic acids can be neutral, acidic or basic. g

both sodium Sodium salts carbonate of carbonate carbonic are basic. acid (Na 2

(H CO

The 2 3 ) and sodium ). Solutions carbonate ion hydrogencarbonate of sodium hydrogencarbonate (NaHCO 3

) and are

undergoes hydrolysis that produces a basic solution. Figure 4.5.1 shows the reaction of the carbonate ion with water:

CO

3

CO

3

2−(aq) + H

2

O(l)  HCO

3

−(aq) + OH−(aq) –

–

carbonate ion water hydrogencarbonate

ion

– –



**FIGURE 4.5.1**

+

hydroxide ion

The reaction of the carbonate ion with water

The hydrogencarbonate ion is amphiprotic and undergoes hydrolysis reactions that produce OH− and H

3

O+: HCO

3

−(aq) + H

2

O(l)  H

2

CO

3

(aq) + OH−(aq) HCO

3

−(aq) + H

2

O(l)  CO

3

2−(aq) + H

3

O+(aq) The first reaction occurs to a greater extent than the second reaction, so the sodium hydrogencarbonate solution is basic. solution Sodium in water hydrogensulfate because the (NaHSO

following 4

) reaction is a salt of dominates:

sulfuric acid and forms an acidic

HSO

4

−(aq) + H

2

O(l)  SO

4

2−(aq) + H

3

O+(aq)

) Sodium phosphate sodium dihydrogenphosphate Sodiumdihydrogenphosphate (Na

3

PO

(NaH

4

dissolves sodium 2 PO

4 ) in are hydrogenphosphate water all salts forming of phosphoric an acidic (Na

acid.

2

solution:

) and

H

2

PO

4

−(aq) + H

2

O(l)  HPO

4

2−(aq) + H

3

O+(aq) Solutions of sodiumdihydrogenphosphate and sodium phosphate in water are basic:

HPO PO

4

3−(aq) 4

2−(aq) + + H

H

2

O(l) 2

O(l)   HPO

H

2

PO

4

2−(aq) 4

−(aq) + + OH−(aq)

OH−(aq)

CHEMISTRY IN ACTION

**Tartaric acid**

Tartaric acid is a diprotic acid that occurs naturally in many plants, but is most commonly associated with grapes. Tartaric acid and its partially neutralised salt, potassium hydrogentartrate, are responsible for the sour taste in wine. Figure 4.5.2 shows the structures of tartaric acid and potassium hydrogentartrate.

C

OH

as brass, aluminium or copper. As these metal tarnish, they react with oxygen in the air to produce metal oxides. The combined mixtures of vinegar or lemon juice and cream of tartar are all acidic, so do not react together when the paste is prepared. When the mixture is used to clean the metal, the metal oxides on the surface react with the acids in the

**(a)**

paste according to the equation:

metal oxide + acid → salt + water

C

OH

OH O

H C

C

O HO

C H

OH

**(b)**

O

HO

**FIGURE 4.5.2**

potassium naturally cooking. bread-making The both Equation Equation Potassium hydrogentartrate a P Brønsted–Lowry It from 1 2 and stabilises bitartrate HOOCC HOOCC hydrogentartrate, (b) for wine potassium its a as 2

2 egg C or leavening H

H ion Structures it 4

4 ‘cream acid O

O

HOOCC ferments whites −OOCC is 2

2 hydrogentartrate

H COO−(aq) COO−(aq) amphiprotic and g of otherwise effect. of OH H and C

2

2

H (a) H Brønsted–Lowry tartar’ and 4

4

O

O tartaric syrups, + + 2

2

C

COO−(aq) COOH(aq) H H is 2

2

crystallises and O(l) O(l) widely known acid O e O– and 

 may K+ + used as is + H base. act OH−(aq) used 3 O+(aq)

in

as

in

*K*

a

for equation 1 > K

b

for equation 2. Therefore, a solution of hydrogentartrate is acidic.

Cream of tartar (Figure 4.5.3) can be mixed with other acidic liquids such as lemon juice or white vinegar to make a paste that is an excellent cleaning agent for metals such

**Cream FIGURE 4.5.3**

of tartar is a crystalline solid.

This mixture is sometimes mistakenly made with vinegar and sodium hydrogencarbonate (baking soda). The hydrogencarbonate ion (HCO

3

−) is also amphiprotic and can act as a Brønsted–Lowry acid and base. Equation 1 HCO

3

−(aq) + H

2

O(l)  CO

3

2−(aq) + H

3

O+(aq) Equation 2 HCO

3

(aq) + OH−(aq) In this case, K

b

−(aq) + H

2

O(l)  H

2

CO

3

for equation 1 and the hydrogencarbonate solution is basic. The baking soda and vinegar react to neutralise one another creating carbon dioxide and sodium ethanoate solution. CH

3

for equation 2 is greater than K

a

COOH(aq) + NaHCO

3

(aq) → NaCH

3

COO(aq) + H

2

O(l) + CO

2

(g) The ethanoate ion in solution can only act as a proton acceptor (Brønsted–Lowry base) and so it is ineffective at cleaning tarnished metals.

P

HPO

4

r

o

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o

f

s

**4.5 Review**

SUMMARY

• Salts are ionic compounds that completely dissociate in aqueous solution.

• Salts are produced in neutralisation reactions. The anions are derived from the acid and cations from the base.

• The anions and cations of salts can act as Brønsted–Lowry acids and bases.

• The conjugate base of a weak acid reacts with

s

water to form hydroxide ions.

• The conjugate acid of a weak base reacts with water to form hydronium ions.

• Salts produced by the reaction of a strong acid and strong base are neutral salts.

KEY QUESTIONS

1 Complete the following table of neutralisation

reactions.

**Reactants Name of salt**

**formed**

• Salts produced by the reaction of a strong acid and weak base are acidic salts.

• Salts produced by the reaction of a weak acid and strong base are basic salts.

• The acidity of salts produced by the reaction of a weak acid and weak base varies and depends on the degree of hydrolysis of the anion and cation in the salt produced.

P

o

f

**Formulae of ions present in the salt solution**

2 P Write a b c equations S2−

ClO

2

−

PH

4

r for the o

hydrolysis of the following ions.

Hydrochloric acid + magnesium hydroxide

Sulfurous acid + zinc hydroxide

Potassium phosphate

Ca2+(aq) + CH

3

+

Ammonium nitrate a

g Na+(aq) + H

2

PO e COO−(aq)

4

−(aq)

d F− 3 Classify the following salts as acidic, basic or neutral.

Write relevant hydrolysis reactions to support your choice. a KNO

3 b NH

4

NO

3 c Ca(HCOO)

2 d MgBr

2

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**Chapter review**

KEY TERMS

acid acid−base reaction acidic proton acidic salt acidic solution acidity acidity constant alkali amphiprotic anion base basic salt basic solution

solute solution strong acid strong base super acid triprotic acid weak acid weak base

**Introducing acids and bases 1 The NH**

4

Brønsted−Lowry theory cation concentrated solution concentration conjugate acid conjugate acid−base pair conjugate base dilution diprotic acid dissociate hydrolysis reaction hydronium ion hydroxide ion

+ ion acts as a weak acid. Which species is likely to be most abundant in 1 mol L−1 NH

4 A NH

4

+

B Cl− C NH

3 D H

3

O+

Cl solution?

2 In which substance A NH

4

indicator ionic product of water ionisation ionisation constant of

water ionise molar monoprotic neutral pH scale polyprotic salt self-ionisation 6 Using P 7 8 a b Draw acid acid–base Identify are a concentration salt

diprotic strong molecule. acid a suitable acid r structural which acid reaction. and and one examples, Identify an o formula amphiprotic or a concentrated more which distinguish of of o the proton the substance monoprotic Brønsted–Lowry acid−base reactions. HCl(aq) + KOH(aq) → KCl(aq) + f

s

between a:

acid.

ethanoic is donated in an

following reactions of acting the following as a base?

equations is the e first-listed

b 2HNO

3 c AgNO

3

H

2

O(l) O(l)

(aq) + Mg(s) → Mg(NO

3

)

2

(aq) + H

2

(g) B Cr

2

+(aq) + OH−(aq) → NH

3

(g) + H

2 O

7

2−(aq) + 2OH−(aq) → 2CrO

4

2−(aq) + H

2

O(l) C 2Na(s) + 2H

2

(aq) + NaCl(aq) → AgCl(s) + NaNO

3

(aq) O(l) → 2Na+(aq) D CH

CH

3

3

NH

2

(aq) + CH

3 COOH(l) g

→ + 2OH−(aq) + H

2

(g)

d CuO(s) + H

2

SO

4

(aq) → CuSO

4

(aq) + H

2

O(l) CH

3

NH

3

+(aq) + COO−(aq) 3 Identify the reactant that acts as an acid in each of the

following reactions.

Strengths of acids and bases 9 The dihydrogen phosphate ion is amphiprotic. The

equilibrium constants for the H

2

a NH

4

− ion acting as an acid and as a base are provided below. H

2

PO

4

b NH

3 +(aq) (g) c HCO

3 d SO

4

+ HCl(g) + a

H

2

O(l) → → NH

NH

4

Cl(s)

3

(aq) + H

3

O+(aq)

2−(aq)

PO

4

−(aq) + H

2

O(l)  HPO

4

2−(aq) + H

3

O+(aq)

−(aq) + OH−(aq) → H

2

O(l) + CO

3

*K*

a

= 6.3 × 10−8

4 e Write a P CO CH

PO

3

4 3 balanced 2−(aq) 2−(aq) COO−(aq) + + H

3

O+(aq) → HSO

4

−(aq) + H

2

O(l) CH

3

COOH(aq) → HCO

3

−(aq) +

equations to show that in water:

*H K*

a 2

PO = 1.6 4

−(aq) × 10−7 + H

2

O(l)  H

3

PO

4

(aq) + OH−(aq)

Which one of the following can correctly be inferred from the data provided?

3− acts as a base b H

2

**A H**

2

PO

4

− readily loses two protons to water.

PO

4

− acts as an amphiprotic substance c H

2

**B H**

2

PO

4

− will form an acidic solution in water.

S acts as an acid.

**C H**

extent 2

PO

4

− in will water.

accept and donate a proton to an equal

5 Write the formula for the conjugate of:

a the acid HCl

**D H**

2

b the base OH− c the base O2− d HSO

4

PO

4

− more readily accepts a proton than donates a proton in water.

− when it acts as an acid.

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**120**

CHAPTER REVIEW CONTINUED

10 The following table lists the pH of 0.00001 mol L−1

solutions of four acids.

**Acid solution pH**

I 4.0

II 3.5

III 5.0

IV 3.0

The acid that cannot be monoprotic and weak is: A I B II C III D IV 11 Write an equation to show that perchloric acid (HClO

4

) acts as a strong acid in water. 12 Write an equation to show that hypochlorous acid

(HClO

3

) acts as a weak acid in water. 13 Write an equation to show that ammonia (NH

3

) acts as a weak base in water. 14 Write a balanced ionic equation showing the HPO

4

2−

ion acting as a weak base in water.

Acidity of solutions 15 Calculate [OH−] at 25°C in aqueous solutions with

[H

3

O+] equal to: a 0.001 mol L−1 b 10−5 mol L−1 c 5.7 × 10−9 mol L−1 d 3.4 × 10−12 mol L−1 e 6.5 × 10−2 mol L−1 f 2.23 × 10−13 mol L−1 16 What is the concentration of hydronium and hydroxide ions in solutions, at 25°C, with the following pH values? a 1 b 3 c 7 d 17 Human neutral? 18 The How coffee? 19 Calculate solutions, a b 11.7 3.0 10.0 P pH many of blood the at a more cola 25°C, concentration has drink times a with pH is acidic the of 3 7.4. Is blood acidic, basic and of black coffee is 5. is the cola than black

of H+ and OH− ions in following pH values.

or

c 8.5 d 5.8 e 9.6 f 13.5

a

g

AREA OF STUDY 2 | ACIDS AND BASES

e

20 The pH of tomato juice is 5.3 at 25°C. What is the

concentration of hydroxide ions in tomato juice?

Dilution of acids and bases 21 A solution of hydrochloric acid has a pH of 2.

a What is the molar concentration of hydrogen ions in

the solution? b What amount of hydrogen ions, in mol, would be

present in 500 mL of this solution? 22 Calculate the pH of each of the following mixtures at

25°C. a 10 mL of 0.025 mol L−1 HCl is diluted to 50 mL of

solution. b 20 mL of 0.0050 mol L−1 KOH is diluted to 500 mL of

solution. c 10 mL of 0.15 mol L−1 HCl is diluted to 1.5 L of

solution. 23 The molarity of concentrated sulfuric acid is

18.0 mol L−1. What volume of concentrated sulfuric acid is required to prepare 1.00 L solution of 2.00 mol L−1 H

2

s

24 25 pH 26 P of When diluted, volume 40.0 Will Three SO salt the mL 4 ? a of solutions of the pH 10.0 of the r water 0.10 increase pH following m L changes mol is solution o added L−1 or solutions decrease? HNO from of to o f

hydrochloric acid is

2.00 to 4.00. What the acid solution?

3

is diluted to 500.0 mL.

have a pH very close to 7. One has a pH of close to 4. Which is the solution with pH 4? A ammonium ethanoate (NH

4

CH

3

COO) B ammonium chloride (NH

4

Cl) C ammonium phosphate ((NH

4

)

3

PO

4

) D sodium bromide (NaBr) 27 a The following ionic salts were all dissolved

separately in equal volumes of water at 25°C. The approximate pH of the solutions formed was measured and recorded as being either neutral (7), acidic (5) or basic (9). Complete the following table.

**Salt dissolved Approximate pH (5, 7 or 9)**

Potassium nitrate

Ammonium chloride

Sodium phosphate

Potassium hydrogensulfate

b Another ionic salt, phosphonium ethanoate

(PH

4

CH

3

COO) was dissolved in water and was found to have a pH of below 7. Use your knowledge of acid–base chemistry to account for this observation. Use equations to support your answer.

**30 Connecting the main ideas**

Construct a concept map that demonstrates your 28 A laboratory assistant forgot to label 0.1 mol L−1

understanding solutions of sodium hydroxide (NaOH), hydrochloric terms: acid (HCl), glucose (C

6

H

12

O

6

), ammonia (NH

3

) and

conjugate. ethanoic acid (CH

3

COOH). In order to identify them, 31 Like temporary labels A–E were placed on the bottles and and the electrical conductivity and pH of each solution was measured. The results are shown in the table

NH

3

below. Identify each solution and briefly explain your reasoning.

**Solution Electrical conductivity pH**

self-ionise, water, acid, liquid base, of according ammonia the proton, links hydrogen to between is the able following to the ion, react following

hydronium equation: with itself

s ion,

A Poor 11

B Zero 7

C Good 13

D Good 1

E Poor 3

29 a Write concise definitions for:

i Brønsted–Lowry acid ii strong base iii molar concentration iv conjugate acid. b Explain, with the aid of equations, why HCO

3

−(am) where ‘am’ means ‘dissolved in ammonia’. a Rewrite the reaction and link the conjugate acid–

base pairs, clearly identifying the acids and bases. The value of the equilibrium constant (at 25°C) for this reaction is about 1.0 × 10−30. b It is found that, as the temperature of the ammonia increased, the value of K and electrical conductivity also increased. Is the ionisation of ammonia endothermic or exothermic?

P

− is classified as amphiprotic.

a

g

e

(l) + NH

3

(l)  NH

4

+(am) + NH

2

P

r

o

o

f

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