کلاس شیمی آمادگی آزمون IMAT-2023

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Brønsted–Lowry acids and bases

The **Brønsted–Lowry** definition is:

- an **acid** is a proton (H^+) **donor**
- a **base** / alkali is a proton (H^+) **acceptor**

a $CH_3COOH(aq) + H_2O(l) \rightarrow CH_3COO^-(aq) + H_3O^+(aq)$ (H^+ donated)

b $CH_3COOH(aq) + H_2O(l) \leftarrow CH_3COO^-(aq) + H_3O^+(aq)$ (H^+ donated)

When CH_3COOH acts as an acid and donates a proton, it forms a base, CH_3COO^-. CH_3COO^- is called the **conjugate base** of CH_3COOH. CH_3COOH and CH_3COO^- are called a **conjugate acid–base pair**. Similarly, when H_2O acts as a base and accepts a proton, it forms H_3O^+, which acts as an acid in the reverse direction. H_3O^+ is the **conjugate acid** of H_2O, and H_3O^+ and H_2O are also a conjugate acid–base pair.

We can label the conjugate acid–base pairs in the equation:

$$\underbrace{CH_3COOH}_{\text{acid 1}}(aq) + \underbrace{H_2O}_{\text{base 2}}(l) \rightleftharpoons \underbrace{CH_3COO^-}_{\text{base 1}}(aq) + \underbrace{H_3O^+}_{\text{acid 2}}(aq)$$

Conjugate acid–base pairs always differ by one proton (H^+).

Acid	Formula
hydrochloric	HCl
sulfuric	H_2SO_4
nitric	HNO_3
carbonic	H_2CO_3
ethanoic	CH_3COOH
benzoic	C_6H_5COOH

Table 8.1 The formulas for some common acids, with the H that is lost as H^+ shown in red.

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

acid — base — conjugate base — conjugate acid

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

base — acid — conjugate acid — conjugate base

5 In each of the following reactions state whether the species in **bold** is acting as an acid or a base according to the Brønsted–Lowry definition:

a $\mathbf{HSO_4^-}(aq) + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$

acid ($-H^+$) → conjugate base; base → conjugate acid

b $\mathbf{HCO_3^-}(aq) + HSO_4^-(aq) \rightleftharpoons H_2CO_3(aq) + SO_4^{2-}(aq)$

base ($+H^+$)

c $\mathbf{CH_3COOH}(aq) + H_2SO_4(aq) \rightleftharpoons CH_3COOH_2^+(aq) + HSO_4^-(aq)$

base ($+H^+$)

2 Write an equation for the reaction between ammonia (NH_3) and ethanoic acid (CH_3COOH), classifying each species as either an acid or base and indicating the conjugate acid–base pairs.

base ← ammonia

$$NH_3 + CH_3COOH \rightleftharpoons CH_3COO^- + NH_4^+$$

NH_3: base; CH_3COOH: acid; CH_3COO^-: conjugate base; NH_4^+: conjugate acid

3 Give the formula of the conjugate acid of each of the following:

a $NH_3 \xrightarrow{+H^+} NH_4^+$

b $OH^- \xrightarrow{+H^+} H_2O$

c $HSO_4^- \xrightarrow{+H^+} H_2SO_4$

d $CN^- \xrightarrow{+H^+} HCN$

e $HPO_4^{2-} \xrightarrow{+H^+} H_2PO_4^-$

f) $HSO_4^{2-} \xrightarrow{+H^+} H_2SO_4^-$

g) $SO_4^{2-} \xrightarrow{+H^+} HSO_4^-$

HCN = Sianic acid

H_2SO_4 = sulfuric acid

H_3PO_4 = phosphoric acid

4 Give the formula of the conjugate base of each of the following:

a HCO_3^- $\xrightarrow{-H^+}$ CO_3^{2-}

b H_2O $\xrightarrow{-H^+}$ OH^-

c HCOOH ⟶ $HCOO^-$

d NH_3 ⟶ NH_2^-

e) H_2CO_3 ⟶ HCO_3^-

f) HCO_3^- ⟶

g) CH_3COOH ⟶ CH_3COO^-

h) H_3PO_4 ⟶ $H_2PO_4^-$

i) $H_2PO_4^-$ ⟶ HPO_4^{2-}

The hydrated proton may be written as $H^+(aq)$ or $H_3O^+(aq)$. H_3O^+ (called the hydronium ion, hydroxonium ion or oxonium ion) has the structure shown in Figure **8.3**. It is essentially just an aqueous hydrogen ion (proton) and is often just written as $H^+(aq)$.

$H_3O^+(aq)$ ≡ $H^+(aq)$

Amphiprotic and Amphoteric

These two terms are used to describe a substance can act as both an acid and a base. Amphiprotic refers to the Bronsted-Lowry definition of acids and bases and indicates a species that can donate (acting as an acid) or accept (acting as a base) a proton. Water is a substance that is amphiprotic and we have seen above that it acted as a proton donor (to form OH^-) in its reaction with NH_3 but as a proton acceptor (to form H_3O^+) in its reaction with CH_3COOH.

H_2O / HCO_3^-

Amphoteric is a more general term and refers to a substance that can act as an acid and a base – all amphiprotic substances are also amphoteric but not all amphoteric substances are amphiprotic. The difference arises because there is another, more general, definition of acids and bases (the Lewis definition), which does not require the transfer of a proton. The most commonly encountered substances that are amphoteric but not amphiprotic are hydroxides of some metals such as aluminium or zinc.

H^+: acid
OH^-: base

OH^- = base
H^+ = acid

$$Al(OH)_3(s) + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l) \qquad \textbf{equation 1}$$

(base) (acid, reactant)

$$Al(OH)_3(s) + OH^-(aq) \rightarrow Al(OH)_4^-(aq) \qquad \textbf{equation 2}$$

(acid) (base)

⇒ $Al(OH)_3$ (Amphoteric)

eq1: $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ (H2O: base)

Bronsted_Lowry

eq2: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ (H2O: acid)

⇒ H_2O = Amphiprotic

8.2 Lewis acids and bases (HL)

The Brønsted–Lowry definition of acids and bases is only one of several definitions of acids and bases – we will now consider the Lewis definition.

Brønsted–Lowry

- An **acid** is a proton (H^+) **donor**.
- A **base** / alkali is a proton (H^+) **acceptor**.

Lewis

- An **acid** is an electron **pair acceptor**.
- A **base** is an electron **pair donor**.

$$NH_3 + H^+ \rightleftharpoons NH_4^+$$

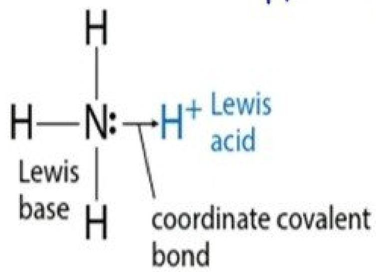

Figure 8.4 Protonation of ammonia.

$$NH_3 + BF_3 \rightleftharpoons H_3N{:}BF_3$$

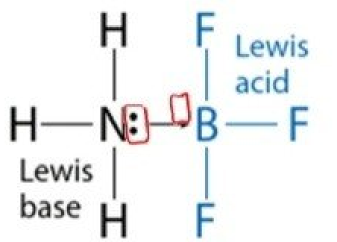

Figure 8.5 An adduct is formed when NH_3 and BF_3 react together.

toxic → □ (acid) money → •• (base)

$$Fe^{2+} + 6H_2O \rightleftharpoons [Fe(H_2O)_6]^{2+}$$

acid base

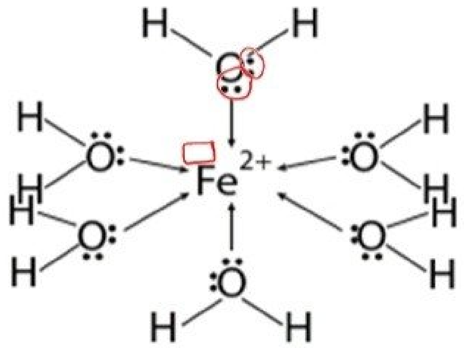

Figure 8.6 A transition metal complex ion.

For a substance to act as a Lewis base, it must have a lone pair of electrons. For a substance to act as a Lewis acid, it must have space to accept a pair of electrons in its outer shell.

7 Classify each of the following as a substance that can act as a Lewis acid or as a Lewis base:

H_2O BF_3 HCO_3^- H^+ $AlCl_3$ NH_3 CO

base acid acid acid base base

F—B—F, F

H•, H^+□

Cl—Al—Cl, Cl

:C≡O:

Group 3: Al, B

Group 6: O

Group 4: C

[H—Ö—C=Ö, :O:]⁻

e = 1 + 4 + 6(3) + 1 = 24

e = 24

e remain = 0

استثناء: HCO_3^- weak acid ⇒ base

8.4 The pH scale

Measuring pH

The pH scale can be used to indicate whether a solution is acidic, alkaline or neutral. At 25 °C, a solution with pH lower than 7 is acidic, a solution with pH 7 is neutral and a solution with pH greater than 7 is alkaline (Figure **8.7**).

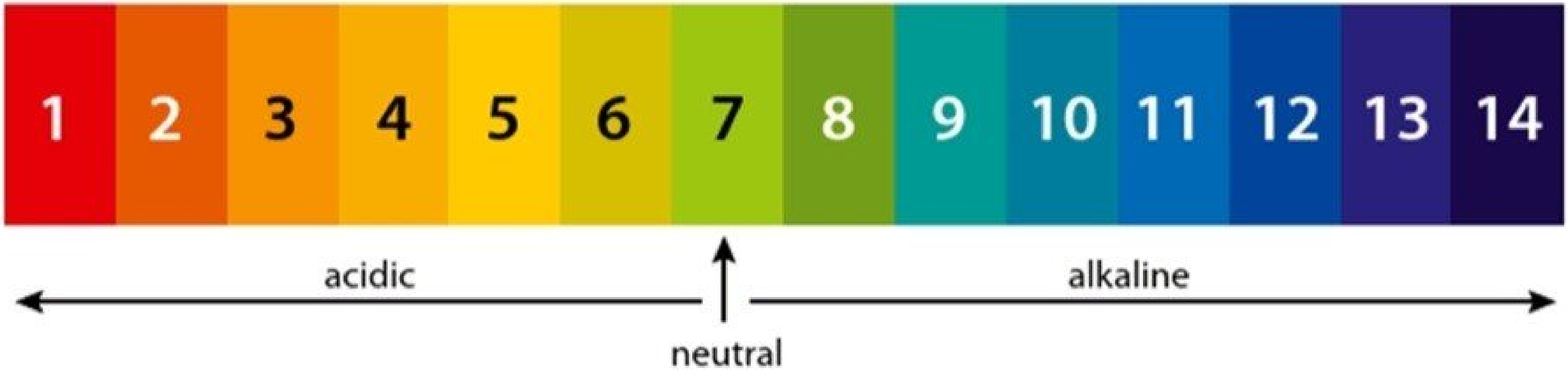

Figure 8.7 The pH scale showing the colours of universal indicator.

11 State whether each of the following statements is **true** or **false**:

F **a** The pH of 0.10 $mol\,dm^{-3}$ HCl is 2. [HCl] = 0.1 = [H^+] pH = 1

T **b** The [H^+] in a solution of pH = 3 is 100 times the [H^+] in a solution of pH = 5.

F **c** The [H^+] in a solution of pH = 13 is 1.0×10^{13} $mol\,dm^{-3}$.

T **d** The pH of 0.010 $mol\,dm^{-3}$ H_2SO_4 is 2.

$[H^+] = 10^{-pH}$

b) pH = 3 → $[H^+] = 10^{-3}$

pH = 5 → $[H^+] = 10^{-5}$

$\frac{10^{-3}}{10^{-5}} = 100$

base	acids
NaOH	HF
KOH	HCl
$Ca(OH)_2$	HBr
	H_2SO_4 , H_2CO_3
	HNO_3 , H_3PO_4

c) pH = 13 → $[H^+] = 10^{-13}$

d) $[H_2SO_4] = 0.01 = [H^+]$ ⇒ pH = $-\log[H^+] = -\log 10^{-2} = 2$

pH = ?

36 Which of the following acids is/are weak? 2022

✓ **1** $H_2CO_3(aq)$

weak

✓ **2** $HCOOH(aq)$

✗ **3** $HNO_3(aq)$

A 1 and 2 only

B 1 only

C 2 only

D 3 only

E 2 and 3 only

7 strong acids

1) H_2SO_4 = sulfuric acid
2) HNO_3 = Nitric acid
3) $HClO_4$ =
4) $HClO_3$ =
5) HCl :
6) HBr
7) HI

Bases

When a base (B) reacts with water, it accepts a proton from the water and ionises according to:

$$B(aq) + H_2O(l) \rightleftharpoons BH^+(aq) + OH^-(aq)$$

Bases are defined as strong or weak depending on how much they ionise in aqueous solution.

Strong bases

Strong bases ionise completely in aqueous solution.

For example, sodium hydroxide ionises completely to produce OH^- ions:

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

Strong bases include the group 1 hydroxides (LiOH, NaOH etc.) and $Ba(OH)_2$.

Exam tip

$Ba(OH)_2$ ionises in aqueous solution to give $2OH^-$:

$$Ba(OH)_2(aq) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$$

So 1 mol dm^{-3} $Ba(OH)_2(aq)$ contains 2 mol dm^{-3} OH^-.

Weak bases

Weak bases ionise only partially in aqueous solution.

Ammonia is a typical weak base and ionises according to the equation:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

In a 0.10 mol dm^{-3} solution of ammonia (at 25 °C) about 1.3% of the molecules are ionised.

Other weak bases are amines such as ethylamine (ethanamine), $CH_3CH_2NH_2$. Ethylamine ionises according to the equation:

$$CH_3CH_2NH_2(aq) + H_2O(l) \rightleftharpoons CH_3CH_2NH_3^+(aq) + OH^-(aq)$$

At 25 °C, a 0.10 mol dm^{-3} solution of ethylamine has about 7.1% of the molecules ionised – so ethylamine is a stronger base than ammonia.

47 X is 50.0 mL of a 0.050 $mol L^{-1}$ aqueous solution of nitric acid (HNO_3).

Y is 100 mL of a 0.050 $mol L^{-1}$ aqueous solution of sulfuric acid (H_2SO_4).

Which of the following statements about these solutions is/are correct?

1. The hydrogen ion concentration in Y is four times greater than the hydrogen ion concentration in X.
2. Only the acid in Y will be completely dissociated into ions.
3. Y has a pH value of less than 2 at 25°C.

A 3 only

B none of them

C 1 only

D 1 and 3 only

E 2 and 3 only

51 X and Y are acidic solutions.

X has a pH of 2 and Y has a pH of 4

Which of the following statements about the hydrogen ion concentrations of X and Y is correct?

A Y has a hydrogen ion concentration 2 times greater than X.

B X has a hydrogen ion concentration 100 times greater than Y.

C X has a hydrogen ion concentration 4 times greater than Y.

D Y has a hydrogen ion concentration 100 times greater than X.

E X has a hydrogen ion concentration 2 times greater than Y.

strong acids

HBr	hydrobromic acid
HI	hydroiodic acid
HCl	hydrochloric acid
$HClO_3$	chloric acid
$HClO_4$	perchloric acid
HNO_3	nitric acid
H_2SO_4	sulfuric acid

strong bases

LiOH	Lithium hydroxide
NaOH	Sodium hydroxide
KOH	potassium 〃
RbOH	Rubidium 〃
CsOH	Cesium 〃
SrOH	strontium 〃
$Ba(OH)_2$	Barium 〃
$Ca(OH)_2$	Calcium 〃

Bronsted - Lowry acids

proton donor

Bronsted - lowry bases

proton acceptor

$$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$$

NH_3: base → NH_4^+: conjugate acid

H_2O: acid → OH^-: conjugate base

Lewis acids

electron pair acceptor

Lewis bases

electron pair donor

$$NH_3 + H^+ \rightarrow NH_4^+$$

NH_3: Lewis base; H^+: Lewis acid

acidic oxides		Amphoteric oxide	basic oxides	
Non-metal oxides		some metal oxides & some semimetal ox.	metal oxides	
NO_2	nitrogen dioxide	★Al_2O_3 Aluminium oxide	Na_2O	sodium oxide
CO_2	carbon dioxide	★MgO	K_2O	potassium "
SO_2	sulphur dioxide	★BeO	Li_2O	Lithium "
SO_3	sulphur trioxide	★BaO	CaO	calsium oxide
P_2O_5		zinc oxide ★ZnO		
N_2O_5		Fe_2O_3 iron(III) oxide		
SiO_2		GeO_2 germanium oxide		
Cl_2O_7		Cr_2O_3		
B_2O_3	(B = semi metal)	SnO		
		PbO		
		N_2O		

4 kinds of oxides

1) acidic (non-metal) : CO_2 = Carbon dioxide

2) basic (metal) : Na_2O : sodium oxide

3) amphoteric : Al_2O_3 = Aluminium Oxide

4) neutral : CO = Carbon monoxide