

CHEM 9.3.4 DEFINING ACIDS AND BASES

Because of the prevalence and importance of acids, they have been used and studied for hundreds of years. Over time, the definitions of acid and base have been redefined

4.1 Outline the **historical development** of ideas about **acids** including those of **Lavoisier, Davy and Arrhenius**

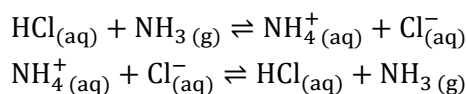
- Initially: Acids and bases recognised, **acids** as **sour**, **bases** as **bitter**

Chemist	Year	Theory	Limitation of theory
Antoine Lavoisier	1780	All acids contain oxygen and that causes acidity	Muriatic acid (HCl) did not contain oxygen, most bases contained oxygen
Humphry Davy	1815	Metals could displace hydrogen in acids Acids and bases formed salts and water	Only classifies substances, without interpreting properties
Svante Arrhenius	1884	Acid ionises in water to produce H ⁺ ions Base ionises in water to produce OH ⁻ ions	Excludes oxides, only applies to aqueous solutions
Johannes Brönsted Thomas Lowry	1923	Independently (Denmark and Britain) Acids are proton donors, Bases are proton acceptors Looked at properties relative to the solvent	

4.2 Outline the **Brönsted-Lowry theory** of acids and bases

- A **proton is transferred** from an acid to a base in an acid-base reaction – water does not need to be present
 - Acids are proton **donors**, bases are proton acceptors
- $\text{HA}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightarrow \text{H}_3\text{O}^+_{(\text{aq})} + \text{A}^-_{(\text{aq})}$, where acid HA has a **greater tendency to give up protons** than solvent
- $\text{B}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightarrow \text{HB}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$, where base B has a **greater tendency to accept protons** than solvent
- Therefore, acids and bases are related to not just properties of acid/base, but **properties of the solvent**

4.3 Describe the **relationship** between an **acid** and its **conjugate base** and a **base** and its **conjugate acid**



- Acid** has a **conjugate base**, has one less proton than acid – both make a **conjugate pair**
 - In equation 1, $\text{HCl}_{(\text{aq})}$ donates proton ('acid') and becomes $\text{Cl}^-_{(\text{aq})}$
 - In equation 2 (reversed equilibrium), $\text{NH}_4^+_{(\text{aq})}$ accepts a proton ('base') and becomes $\text{NH}_3(\text{g})$
- Similarly, base** has a **conjugate acid** – $\text{NH}_3(\text{g})$ donates to become $\text{NH}_4^+_{(\text{aq})}$, and vice versa
- Relationship** between strength of acid/base:
 - Strong acid** has a **weak conjugate base**, **weak acid** has a **strong conjugate base**
 - Similarly, **strong base** has a **weak conjugate acid**, **weak base** has a **strong conjugate acid**

4.4 Identify a **range of salts** which form **acidic, basic or neutral solutions** and explain their acidic, neutral or basic **nature**

- Salt forms when **acid neutralises base**, but not necessarily **neutral** – ions can act as acids/bases in solution
- Depends on reactants of a neutralisation reaction
 - **Neutral** salt formed by **strong acid and strong base**, e.g. NaCl, KNO₃, Na₂SO₄
 - **Neutral** salt formed by **weak acid and weak base**, e.g. CH₃COONH₄ (ammonium acetate)
 - **Acidic** salt formed by **strong acid and weak base**, e.g. NH₄Cl, NH₄NO₃
 - **Basic** salt formed by **weak acid and strong base**, e.g. CH₃COONa, KNO₂, NaCO₃
- Salts are acidic or basic, as ions react with water to form H₃O⁺ or OH⁻ ions
 - NaCl is neutral as both **Na⁺ and Cl⁻ do not react** with water (they ionise if attracted to H/OH)
 - NH₄Cl is acidic as NH₄⁺ reacts with water (NH₄⁺ + H₂O → NH₃ + H₃O⁺) while Cl does not react, so therefore **NH₄Cl + H₂O → Cl⁻ + NH₃ + H₃O**
 - CH₃COONa is basic as CH₃COO⁻ reacts with water (CH₃COO⁻ + H₂O → CH₃COOH + OH⁻) while Na does not react, so therefore **CH₃COONa + H₂O → Na⁺ + CH₃COOH + OH⁻**

4.5 Identify **conjugate acid/base pairs**

- Conjugate pairs identified by looking at its reaction (see 4.3):
 - The **acid pairs** for HCl_(aq) + H₂O_(l) → H₃O_(aq)⁺ + Cl_(aq)⁻ are HCl/Cl⁻ and H₂O/H₃O⁺
 - The **conjugate acid** of Cl⁻ is the base plus a proton – HCl
 - The **conjugate base** of H₃O is the acid minus a proton – H₂O

4.6 Identify **amphiprotic substances** and construct equations to describe their **behaviour in acidic and basic solutions**

- **Amphiprotic** – can act as both **acid and base** (different from amphoteric – they are oxides)
- **HCO₃⁻ (hydrogen carbonate), HSO₃⁻ (hydrogen sulfite) and water** are amphoteric
- Shown in reaction with **water** to a small extent:
 - **HCO₃⁻_(aq) + H₂O_(l) ⇌ CO₃²⁻_(aq) + H₃O_(aq)⁺** where HCO₃⁻ donates proton and acts as **acid**
 - **HCO₃⁻_(aq) + H₂O_(l) ⇌ H₂CO_{3(aq)} + OH_(aq)⁻** where HCO₃⁻ accepts proton and acts as **base**
- However, when placed into acidic/basic environments, reaction completes:
 - **Acidic environment: HCO₃⁻_(aq) + H₃O_(aq)⁺ → H₂CO_{3(aq)} + H₂O_(l)** where HCO₃⁻ acts as **base**
 - **Basic environment: HCO₃⁻_(aq) + OH_(aq)⁻ → CO₃²⁻_(aq) + H₂O_(l)**
- Water is also amphoteric – H₂O acts as **acid** and becomes OH⁻ and H₂O acts as **base** and becomes H₃O⁺

4.7 Identify **neutralisation** as a **proton transfer reaction** which is **exothermic**

- Neutralisation – acid + base → salt + water
- E.g. **HCl_(aq) + NaOH_(aq) → H₂O_(l) + NaCl_(aq)** (not always H and OH)
 - **Acid transfers a proton to the base**
 - As a net equation, **H_(aq)⁺ + OH_(aq)⁻ → H₂O_(l)**, where the H⁺ ion transfers proton to the OH⁻
- Is exothermic: **H_(aq)⁺ + OH_(aq)⁻ → H₂O_(l) ΔH = -57 kJ/mol**

4.8 Describe the correct **technique** for **conducting titrations** and **preparation of standard solutions**

CONDUCTING TITRATIONS

- Titration – using **volumetric analysis**, find the concentration of a solution
 - **Standard** solution (known concentration) is added to a solution (unknown) until **reaction is complete**
 - **Equivalence point** – when reactants and products are **in ratio** with its balanced equation – **indicator** used
 - **Indicator** based on strength of acid/base (position of equivalence):
 - **Strong base and strong acid** – equivalence at 7 – bromothymol blue
 - **Strong base and weak acid** – equivalence around 9 – phenolphthalein
 - **Weak base and strong acid** – equivalence around 5 – methyl orange
 - **Weak base and weak acid** – no sudden change in pH, titration with indicator not used
1. **Clean and rinse** equipment with correct liquid
 - **Conical flask** – rinsed with distilled water – left wet
 - **Burette** – rinsed with distilled water, open tap to flow out, then rinsed with solution
 - **Pipette** – rinsed with distilled water, then rinsed with solution
 2. Add solution into **burette** (funnel, meniscus on top of mark) and use **pipette** to measure volume of the other solution
 3. Place solution in pipette into conical flask and add 3 to 4 drops of indicator
 4. Add solution from burette into conical flask slowly, while swirling, until indicator changes colour (white paper under flask)

PREPARING STANDARD SOLUTIONS

- Common acids and bases have **changing concentrations** over time
 - HCl and HNO₃ volatile, H₂SO₄ absorbs water, NaOH and KOH absorbing moisture, reacts with CO₂
 - **Primary standards** – suitable for creating standard solutions – **has a known concentration**
 - Must be **pure, chemically stable** and dissolves in **water**
 - **E.g.** sodium carbonate, sodium hydrogen carbonate, oxalic acid
1. **Primary standard purified** – placed in an oven, cooled in a desiccator (remove moisture)
 2. **Volumetric flask, watch glass and funnel cleaned** (rinse with distilled water)
 3. Place primary standard onto watch glass and **measure weight** on electronic scale
 4. Wash (with distilled water) into **volumetric flask** – all solute flows into volumetric flask
 5. Fill flask until meniscus is on the graduation mark, invert flask few times

4.8 Qualitatively describe the **effect of buffers** with reference to a specific **example in a natural system**

- Buffer can **maintain constant pH**, even when strong acid/base is added
 - E.g. CH₃COOH and CH₃COONa, H₂CO₃ and NaHCO₃ (sodium bicarbonate), NH₃ and NH₄Cl
- Equilibrium established between acid and conjugate base (in the salt):
 - $CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$
 - Approximately **equal amounts of weak acid and the salt of that acid (its conjugate base)**
 - When **H⁺ added** (acid), equilibrium moves to **left** and pH change minimised
 - When OH⁻ added, reacts with the H⁺ ions, equilibrium moves to the **right** and pH change minimised
- In nature, freshwater **lakes and rivers** are buffer systems
 - CO₂ dissolves in water to form **H₂CO₃**, conjugate base HCO₃⁻ from **carbonate rocks and minerals**
 - $H_2CO_{3(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HCO_3^-_{(aq)}$
 - Therefore protects against **acid rain**