Chem Factsheet

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Lewis Acids and Bases

To succeed in this topic you need to understand:
- the pH scale and Bronsted Lowry definitions of acids and bases
- dot and cross diagrams to show the outer electrons of atoms
- co-ordinate (dative) covalent bonding

After working through this Factsheet you will be able to:
- define and identify Lewis acids and bases
- explain how Lewis acids and bases react with each other
- describe some of the applications of Lewis acids and bases

What are Lewis acids and bases?
At GCSE level, acids and bases are defined in terms of their pH values. This is the basis of the Bronsted-Lowry theory of acids and bases, in which an acid is defined as a proton donor (that is, a donor of H^+ ions) and a base is defined as a proton acceptor.

The problem with these definitions is that they are restricted to the behaviour of acids and bases in aqueous solution. The Lewis theory of acids and bases extends the concept to include substances which do not dissolve readily in water.

> **A Lewis acid is an electron pair acceptor.**
> **A Lewis base is an electron pair donor.**

Examples of Lewis bases
Lewis bases have lone pairs of electrons which they can donate to form a co-ordinate covalent bond. Common examples of Lewis bases include: H_2O, NH_3, Cl^- and OH^-. Note that these bases are also Bronsted-Lowry bases (and therefore alkalis) because they can accept H^+ ions.

> **What is an alkali?**
> An alkali is a base that is soluble in water. All Bronsted-Lowry bases are alkalis, since they are in aqueous solution. Not all Lewis bases are soluble in water, so they are not necessarily alkalis.

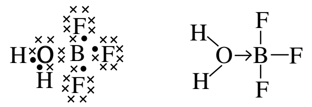
Examples of Lewis acids
Lewis acids can accept a bonding pair of electrons. The most common examples of Lewis acids are transition metal ions (Fe^{2+}, for example) and H^+ ions. Other Lewis acids include $AlCl_3$ and BF_3 (Fig 1). Aluminium and boron are in Group 3, so they have three outer electrons. When bonded in compounds like $AlCl_3$ and BF_3, they have only six outer electrons, so there is space available for a bonding pair of electrons to join the central atom.

Fig 1. The Lewis acids BF_3 and $AlCl_3$

$AlCl_3$ BF_3

How do Lewis acids and bases react with each other?
When Lewis bases react with Lewis acids, they form a **complex**, where a **co-ordinate bond** (also called a dative covalent bond) forms between the acid and the base. A co-ordinate bond is a covalent bond where both shared electrons originate from the same atom. In Lewis acid-base complexes, the pair of electrons originates from the Lewis base. Fig 2 shows the complex formed between water, a Lewis base, and boron trifluoride, a Lewis acid. The diagram on the right uses the convention of drawing the co-ordinate bond as an arrow, with the bonding pair of electrons originating from the oxygen.

Fig 2. The complex formed between a Lewis acid (BF_3) and a Lewis base (H_2O)

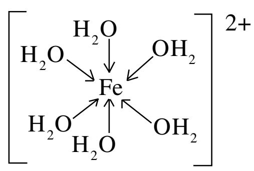


Applications of Lewis acids and bases

1. Transition metal chemistry
One of the characteristic properties of transition metals is that they form **complex ions**. A complex ion contains a central transition metal ion surrounded by **ligands**, which are molecules or ions forming a co-ordinate covalent bond with the transition metal (Fig 3).

The pair of electrons in the co-ordinate bond originates from the ligand, which means that ligands are Lewis bases. Many of the Lewis bases in the discussion above form common transition metal complexes.

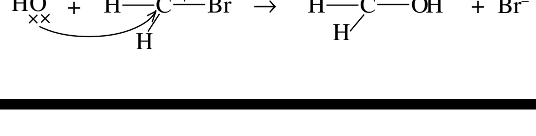
Fig 3. The complex ion $[Fe(H_2O)_6]^{2+}$, formed from an Fe^{2+} ion surrounded by six water ligands



> **What is the difference between a Lewis base and a reducing agent?**
> A Lewis base is an **electron pair** donor. It uses its lone pair to form a covalent bond, thus sharing its electrons with another atom. A reducing agent is an electron donor. It transfers **one or more electrons** to another atom, thus becoming oxidised itself. Conversely, a Lewis acid is an electron pair acceptor, while an oxidising agent is an acceptor of one or more electrons.

2. Nucleophilic substitution reactions
Nucleophiles are Lewis bases because they use a lone pair of electrons to make a bond with an organic molecule (Fig 4).

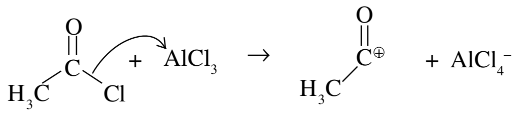
Fig 4. Nucleophilic substitution reaction of bromomethane by the OH^- ion



3. Acylation of benzene

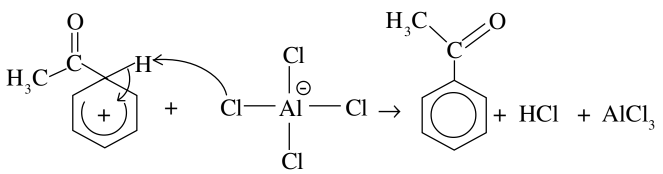
Aluminium chloride is used as a catalyst in the Friedel-Crafts acylation of benzene. The ability of aluminium chloride to act as a Lewis acid allows it to react with an acyl chloride and accept a pair of electrons from the chlorine (Fig 5). This forms the complex $AlCl_4^-$ and the electrophile CH_3CO^+, which goes on to react with the benzene ring.

Fig 5. The first step in the Friedel-Crafts acylation of benzene



In the final step of this electrophilic substitution reaction, the $AlCl_3$ is regenerated (Fig 6).

Fig 6. Regeneration of the $AlCl_3$ catalyst

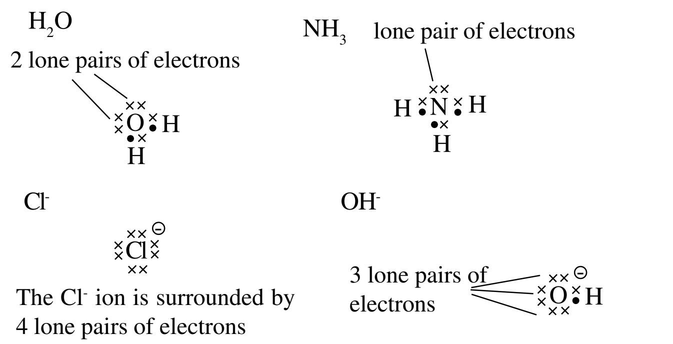


Practice Questions

1. **Define the following terms: (a) Lewis acid, (b) Lewis base, (c) co-ordinate bond.**
2. Draw dot and cross diagrams of these Lewis bases:
   (a) H_2O
   (b) NH_3
   (c) Cl^-
   (d) OH^-
   Identify the lone pairs of electrons.
3. Write equations to show the reaction of each of the following Lewis bases with an H^+ ion:
   (a) H_2O
   (b) NH_3
   (c) Cl^-
   (d) OH^-
4. Classify each of these substances as a Lewis acid or a Lewis base: CN^-, CH_3NH_2, BF_3, Cu^{2+}, Br^-, Fe^{3+}.
5. Draw the complexes formed between the following Lewis acids and bases, showing the co-ordinate bond clearly.
   (a) H_2O and H^+
   (b) NH_3 and H^+
   (c) $AlCl_3$ and Cl^-
   (d) NH_3 and BF_3
   (e) $AlCl_3$ and $AlCl_3$
6. Identify the Lewis base in each of the following transition metal complex ions.
   (a) $[V(H_2O)_6]^{2+}$
   (b) $CuCl_4^{2-}$
   (c) $[Cr(CN)_4]^-$
   (d) $[Fe(NH_3)_6]^{3+}$
   (e) $[Ni(NH_2CH_2CH_2NH_2)_3]^{2+}$

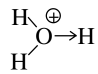
Answers

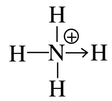
1. (a) A Lewis acid is an electron pair acceptor.
   (b) A Lewis base is an electron pair donor.
   (c) A co-ordinate bond is a covalent bond in which both electrons come from the same atom.

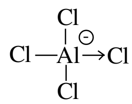
2. 

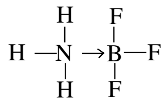
3. (a) $H_2O + H^+ \rightarrow H_3O^+$
   (b) $NH_3 + H^+ \rightarrow NH_4^+$
   (c) $Cl^- + H^+ \rightarrow HCl$
   (d) $OH^- + H^+ \rightarrow H_2O$

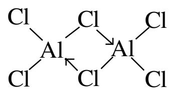
4. | **Lewis acids** | **Lewis bases** |
   |---|---|
   | BF_3, Cu^{2+}, Fe^{3+} | CN^-, CH_3NH_2, Br^- |

5 (a) $H_2O + H^+ \rightarrow H_3O^+$ 

(b) $NH_3 + H^+ \rightarrow NH_4^+$ 

(c) $AlCl_3 + Cl^- \rightarrow AlCl_4^-$ 

(d) $NH_3 + BF_3 \rightarrow H_3NBF_3$ 

(e) $AlCl_3 + AlCl_3 \rightarrow Al_2Cl_6$ 

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	Complex ion	Lewis base
(a)	$V(H_2O)_6]^{2+}$	H_2O
(b)	$CuCl_4^{2-}$	Cl^-
(c)	$[Cr(CN)_4]^-$	CN^-
(d)	$[Fe(NH_3)_6]^{3+}$	NH_3
(e)	$[Ni(NH_2CH_2CH_2NH_2)_3]^{2+}$	$NH_2CH_2CH_2NH_2$

Note: this is a **bidentate** ligand, since both N atoms act as Lewis bases and form bonds with the central metal ion.

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