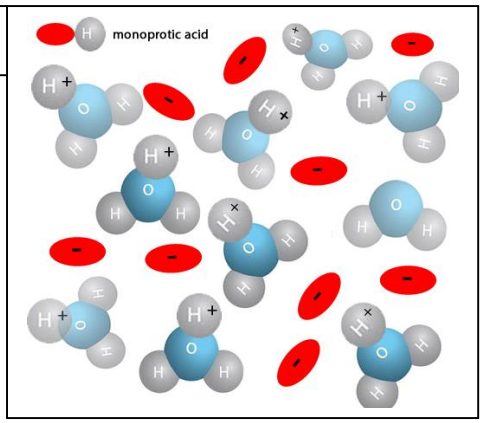
1) Consider the four diagrams of a monoprotic acid ionising in water.

A



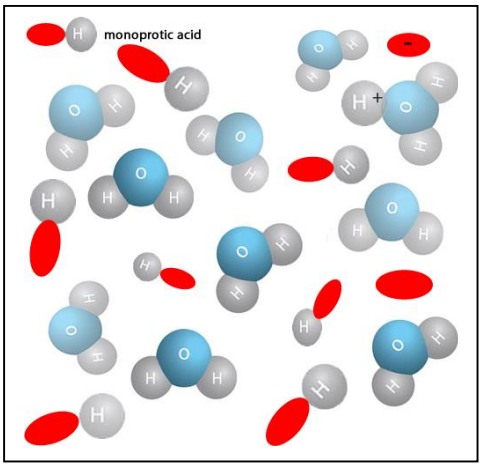
Identify the diagram on the right and give a reason as to which one represents:

- a dilute solution of a weak acid

concentrated, strong acid.
High percentage of acid molecules are ionised = strong acid

relative to the number of water molecules there is a significant number of acid molecules packed in → concentrated

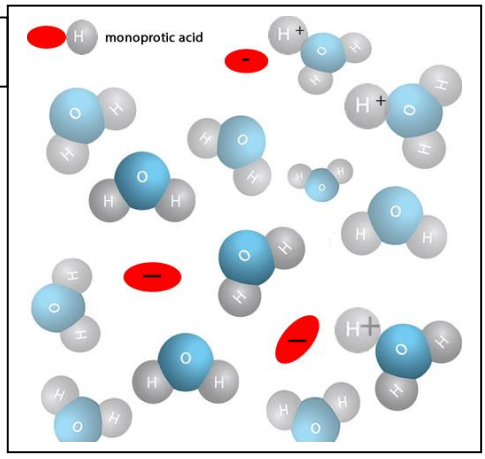
B



- a concentrated solution of a strong acid
- *concentrated, weak acid.*
  Low percentage of acid molecules are ionised = weak acid

  relative to the number of water molecules there is a significant number of acid molecules packed in→ concentrated

C

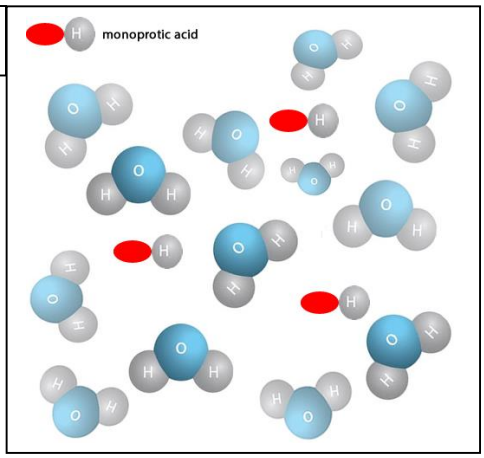


- a dilute solution of a strong acid

dilute, strong acid.
High percentage of acid molecules are ionised = strong acid

relative to the number of water molecules there is a low number of acid molecules packed in → dilute

D



- a concentrated solution of a weak acid
- *dilute, weak acid.*
  Low percentage of acid molecules are ionised = weak acid

  relative to the number of water molecules there is a low number of acid molecules packed in→ dilute

2) Calculate the pH of the following solutions with the given concentration of hydronium or hydroxide ions. Use the formulae on the right.

$$pH = -\log_{10}[H_3O^+]$$

$$10^{-14} = [H_3O^+][OH^-]$$

a. $[H_3O^+] = 0.100M$
$pH = -\log_{10}[H_3O^+] = -\log_{10}[10^{-1}] = 1$

b. $[H_3O^+] = 0.00100M$
$pH = -\log_{10}[H_3O^+] = -\log_{10}[10^{-3}] = 3$

c. $[H_3O^+] = 0.500M$
$pH = -\log_{10}[H_3O^+] = -\log_{10}[10^{-0.301}] = 0.301$

d. $[H_3O^+] = 0.0703M$
$pH = -\log_{10}[H_3O^+] = -\log_{10}[10^{-1.153}] = 1.153$

e. $[H_3O^+] = 3.45 \times 10^{-6}$ M
$[H_3O^+] = 10^{0.538} \times 10^{-6} = 10^{-5.462}$
$pH = -\log_{10}[H_3O^+] = -\log_{10}[10^{-5.462}] = 5.462$

3) Calculate the $[OH^-]$ of the solutions given in a) above if the solution is at 25°C.

a. $[H_3O^+] = 0.100M$ *$= 10^{-1}M$*
$[OH^-]$ *$= 10^{-14} / 10^{-1} = 10^{-13}M$*

b. $[H_3O^+] = 0.00100M$ *$= 10^{-3}M$*
$[OH^-]$ *$= 10^{-14} / 10^{-3} = 10^{-11}M$*

c. $[H_3O^+]$ *$= 0.500M = 10^{-0.301}M$*
$[OH^-]$ *$= 10^{-14} / 10^{-0.301} = 10^{-13.699}M$*

d. $[H_3O^+]$ *$= 0.0703M = 10^{-1.153}M$*
$[OH^-]$ *$= 10^{-14} / 10^{-1.153} = 10^{-12.847}M$*

f. $[H_3O^+]$ *$= 3.45 \times 10^{-6}\ M = 10^{0.538} \times 10^{-6} = 10^{-5.462}$*
$[OH^-]$ *$= 10^{-14} / 10^{-5.462} = 10^{-8.538}M$*

4) Consider a solution made by dissolving 3.65g of pure HCl in a 250mL volumetric flask using distilled water.

a. Calculate the $[OH^-]$ of the solution. Show all calculations.
Step 1 Find the mol of HCl
=> 3.65 /36.5 = 0.100 mol
Step 2 find the concentration of HCl
=> 0.100 / 0.250 = 0.400M = $10^{-0.398}M$
Step 3 Since HCl is a strong acid, mol of HCl = mol of H_3O^+.
Calculate the $[OH^-]$
=> $[OH^-] = 10^{-14} / 10^{-0.398} = 10^{-13.602}M$

b. Find the pH of the solution. Show all calculations.

$pH = -\log_{10}[H_3O^+] = -\log_{10}[10^{-0.398}] = 0.398$

5) Find the pH of the following solutions given their $[OH^-]$

a. $[OH^-] = 0.00100M = 10^{-3}M$

$[H_3O^+] = 10^{-14}/10^{-3} = 10^{-11}M$

$pH = 11$

b. $[OH^-] = 0.900M = 10^{-0.0458}$

$[H_3O^+] = 10^{-14}/10^{-0.0458} = 10^{-13.95}M$

$pH = 13.95$

$[OH^-] = 5.00 \times 10^{-4} = 10^{0.699} \times 10^{-4} = 10^{-3.301}$

$[H_3O^+] = 10^{-14}/10^{-3.301} = 10^{-10.699}M$

$pH = 10.7$

6) 9.60 grams of a weak, monoprotic acid, known as acetic acid ($C_2H_4O_2$), was dissolved in 200 mL of distilled water.

a. Explain why it is not possible to determine accurately the pH of the resulting solution.

Since the degree of ionisation of the weak acid is very low, extra information, such as the percentage ionisation, is needed.

b. The solution was neutralised by the addition of NaOH.

i. Write the balanced overall equation for the reaction between acetic acid and NaOH.

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

ii. Write the balanced ionic equation for the reaction in i. above.

$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

7) 0.400 grams of NaOH was totally dissolved in 500 mL of distilled water.

a. Calculate the $[OH^-]$ of the resulting solution.

Step 1 calculate the mol of NaOH

=> 0.400/40.0 = 0.00400 mol

Step 2 find the concentration in mol/L

=> 0.400 mol / 0.500 = 0.800M =

b. Calculate the pH of the solution.

Step 1 find the $[H_3O^+]$

=> $10^{-14} = [H_3O^+][OH^-] = 10^{-0.0969}$

=> $[H_3O^+] = 10^{-14}/10^{-0.0969} = 10^{-13.903}$

Step 2 find the pH

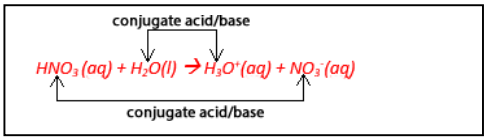
=> $pH = -\log_{10}[10^{-13.903}] = 13.903$

8) An acidic solution was formed by dissolving HNO_3 in water.

a. Give the overall equation of the reaction that takes place between the acid and the water

$HNO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$

b. Indicate the conjugate pairs in question a. above.

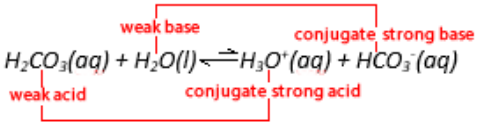


c. Which of the following are acid/base conjugate pairs.

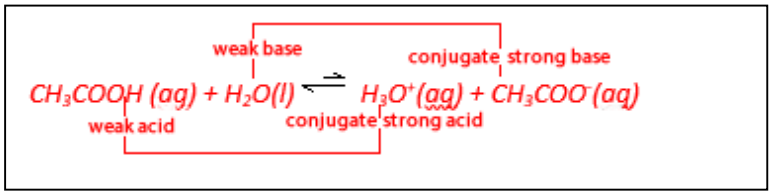
i. HSO_4^- / SO_4^{-2}
ii. H_2SO_4/SO_4^{-2}
iii. H_2CO_3/CO_3^{-2}
iv. NH_4^+/NH_3
v. HCO_3^-/CO_3^{-2}
vi. CO_3^{-2}/CO_2
vii. H_3PO_4/H_2PO_4

Conjugate pairs differ by just one H^+

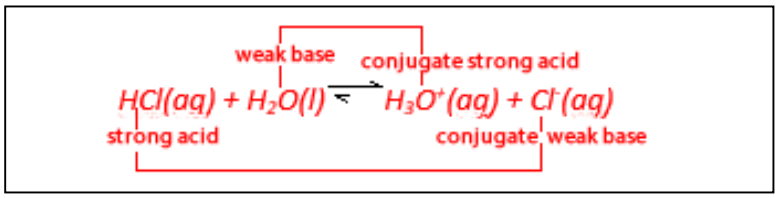
d. Write balanced equations to the reactions taking place and label the reactatns and products with the labels provided below. The first one is labelled for you.
Weak acid, weak base, weak conjugate acid, weak conjugate base.
Carbonic acid ionises when placed in water
$H_2CO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HCO_3^-(aq)$



i. Acetic acid (CH_3COOH) ionises in water



ii. HCl ionises in water



iii. Ammonia (NH_3) acts as a base and ionises in water.