Solutions for „Simple pH and concentration calculation problems”

Revision:

1. What is the mass of a 51.6 mL sample of gasoline, which has a density of 0.70 g/cm³?

\[ \rho (\text{density}) = \frac{m (\text{mass})}{V (\text{volume})} \], from this \[ m = \rho \times V = 0.70 \text{ g/cm}^3 \times 51.6 \text{ cm}^3 = 36.12 \text{ g} \]

2. A flask contains 25.0 mL of ether weighing 17.84 g. What is the density of ether?

\[ \rho = \frac{m}{V} = \frac{17.84 \text{ g}}{25 \text{ cm}^3} = 0.7136 \text{ g/cm}^3 \]

3. How many moles of sodium chloride should be put in a 50.0 mL volumetric flask to give a 0.15 M NaCl solution when the flask is filled with water? How many grams of NaCl is this? (NaCl: M.W. = 58.5 g/mol)

\[ c (\text{molar concentration or molarity}) = \frac{n (\text{molar amount})}{V (\text{volume})} \], from this \[ n = c \times V = 0.15 \text{ mmol/cm}^3 \times 50 \text{ cm}^3 = 7.5 \text{ mmol} = 0.0075 \text{ mol} \].

1 mol of NaCl is 58.5 g, that is 0.0075 moles are: \[ m = n \times \text{M.W.} \text{ (molar weight)} = 0.0075 \text{ mol} \times 58.5 \text{ g/mol} = 0.43875 \text{ g} \]

4. A sample of NaOH weighing 0.38 g is placed in a 50.0 mL volumetric flask. The flask is then filled with water to the mark. What is the molarity of the resulting solution? Is it necessary to use a volumetric flask to make this solution? (NaOH: M.W. = 40.0 g/mol)

To calculate the molarity (\( c \)) of the solution, we need the molar amount of the solute and the volume of the solution. The volume (50 mL) is given. To calculate the molar amount, we use the following equation: \[ m = n \times \text{M.W.} \text{, that is } n = \frac{m}{\text{M.W.}} = \frac{0.38 \text{ g}}{40 \text{ g/mol}} = 0.0095 \text{ mol} = 9.5 \text{ mmol} \].

Than: \[ c = \frac{n}{V} = \frac{9.5 \text{ mmol}}{50 \text{ cm}^3} = 0.19 \text{ M} \]

It is not necessary to use a volumetric flask, since, anyway, the concentration cannot be very accurate because of the carbonate content of the solid NaOH.

5. How many moles of NaOH are contained in 31 mL of 0.15 M NaOH?

\[ c = \frac{n}{V} \text{, that is } n = c \times V = 0.15 \text{ mmol/cm}^3 \times 31 \text{ cm}^3 = 4.65 \text{ mmol} = 0.00465 \text{ mol} \]

6. How would you prepare 425 g of an aqueous solution containing 2.40% (by mass) of sodium acetate?
The mass of sodium acetate needed is 2.40 % of the total mass that is 425 g \times \frac{2.40}{100} = 10.2 \text{ g of sodium acetate.} The amount of water needed is the remaining mass that is 425 \text{ g} - 10.2 \text{ g} = 414.8 \text{ g of water.}

7. You are given 5.00 mL of 14.8 M \text{NH}_3 solution. What will the final volume be after this solution is diluted with water to give 1.00 M \text{NH}_3 solution?

To calculate the final volume of the solution, we need to know the molar amount of the solute and the concentration of the final solution. The final concentration (1.00 M) is given. The molar amount of \text{NH}_3 is the same in the initial 5 mL and in the final solution, because we do not add any more \text{NH}_3 during the dilution. This molar amount (calculated from the initial solution) is:

\[ n = c \times V = 14.8 \text{ mmol/cm}^3 \times 5 \text{ cm}^3 = 74 \text{ mmol} = 0.074 \text{ mol NH}_3 \]

Now, we know the concentration and also the molar amount of \text{NH}_3 in the final solution, so we can calculate the volume of it:

\[ c = \frac{n}{V} \text{ that is V} = \frac{n}{c} = \frac{74 \text{ mmol}}{1 \text{ mmol/cm}^3} = 74 \text{ cm}^3 = 74 \text{ mL} \]

8. Commercially available concentrated hydrochloric acid is an aqueous solution containing 38% \text{HCl} by mass. It’s density is 1.19 g/mL. What is the molarity of this solution? How many milliliters of this concentrated solution is required to make 1.00 L of 0.10 M \text{HCl} solution?

To calculate the molarity (c) of the solution, we need the molar amount of the solute (\text{HCl}) and the volume of the solution. Let’s take 1 mL of this solution (you can choose also a different volume). In this case, from the density of the solution, we can calculate the mass of the solution, which is:

\[ m = \rho \times V = 1.19 \text{ g/mL} \times 1 \text{ mL} = 1.19 \text{ g} \]

38% of this mass is \text{HCl}, that is the mass of \text{HCl} (solute) is \[1.19 \times \frac{38}{100} = 0.4522 \text{ g} \]

So, the molar amount of \text{HCl} is:

\[ n = \frac{m}{\text{M.W.}} = \frac{0.4522 \text{ g}}{36.5 \text{ g/mol}} = 0.0124 \text{ mol} = 12.4 \text{ mmol.} \]

Finally, the concentration of this solution is:

\[ c = \frac{n}{V} = \frac{12.4 \text{ mmol}}{1 \text{ cm}^3} = 12.4 \text{ M} \]

If we want to make 1 L of 0.10 M \text{HCl} solution, than we need the volume which contains the required amount of \text{HCl}. In the final solution we need:

\[ n = c \times V = 0.1 \text{ M} \times 1 \text{ L} = 0.1 \text{ mol of HCl.} \]

The same amount is needed in the concentrated solution, that is (for the concentrated solution):

\[ V = \frac{n}{c} = \frac{0.1 \text{ mol}}{12.4 \text{ mol/L}} = 0.0081 \text{ L} = 8.1 \text{ mL} \text{ of concentrated HCl solution.} \]
Simple pH calculations:

1. Calculate the pH of the solution:
   a) $[H_3O^+] = 0.01 \text{ mol/dm}^3$
   b) $[H_3O^+] = 4.1 \times 10^{-3} \text{ mol/dm}^3$
   c) $[H_3O^+] = 2.65 \times 10^{-5} \text{ mol/dm}^3$

   a) pH = 2.00
   b) pH = 2.39
   c) pH = 4.58

2. Calculate the $[H_3O^+]$ and the $[OH^-]$ concentrations in the following solutions:
   a) pH = 2.32
   b) pH = 6.45
   c) pH = 11.32

   a) $[H_3O^+] = 4.786 \times 10^{-3} \text{ mol/dm}^3$  $[OH^-] = 2.09 \times 10^{-3} \text{ mol/dm}^3$
   b) $[H_3O^+] = 3.548 \times 10^{-7} \text{ mol/dm}^3$  $[OH^-] = 2.82 \times 10^{-8} \text{ mol/dm}^3$
   c) $[H_3O^+] = 4.785 \times 10^{-12} \text{ mol/dm}^3$  $[OH^-] = 2.09 \times 10^{-3} \text{ mol/dm}^3$

3. A sample of orange juice has an equilibrium hydrogen ion concentration of $2.9 \times 10^{-4} \text{ M}$. What is the pH? Is the solution acidic, basic or neutral?

   pH = -log[$H^+]$ = -log(2.9\times10^{-4}) = 3.54
   pH < 7.00 so the solution is acidic.

4. A saturated solution of potassium hydroxide has an equilibrium hydroxide ion concentration of 0.05 M. What is the pH? Is the solution acidic, basic or neutral?

   pOH = -log[$OH^-]$ = -log(0.05) = 1.30
   pH + pOH = 14.00 that is pH = 14.00 – pOH = 14.00 – 1.30 = 12.70
   pH > 7.00 so the solution is basic.

5. A HClO$_4$ solution has an analytical concentration of 0.012 M. What is the equilibrium hydrogen ion concentration and the pH?

   In the case of strong acids, the dissociation is 100%, so the concentration of the acid and the $H^+$ ion is the same (expect when $c_{\text{strong acid}} < 10^{-6} \text{ M}$, but this is not our case). So, $[H^+] = c_{\text{strong acid}} = 0.012 \text{ M}$, and pH = -log[$H^+] = -log(0.012) = 1.92$

6. A NaOH solution has an analytical concentration of $10^{-7} \text{ M}$. What is the equilibrium hydroxide ion concentration, the pOH and the pH?

   In the case of strong bases, the dissociation is 100%, but now $c_{\text{strong base}} < 10^{-6} \text{ M}$, so the self-dissociation of the water has also to be taken into account:

   $[OH^-] = c_{\text{strong base}} + \frac{K_w}{[OH^-]}$
   $[OH^-] = 10^{-7} + \frac{10^{-14}}{[OH^-]}$
   $[OH^-]^2 = 10^{-7}[OH^-] + 10^{-14}$
   $[OH^-]^2 - 10^{-7}[OH^-] - 10^{-14} = 0$
\[ [OH^-] = \frac{(-10^{-7}) \pm \sqrt{(-10^{-7})^2 - 4 \times 1 \times (-10^{-14})}}{2 \times 1} = 1.618 \times 10^{-7} \text{ M} \]
pOH = \log([OH^-]) = \log(1.618 \times 10^{-7}) = 6.79
\[ \text{pH + pOH} = 14.00 \text{ that is pH} = 14.00 - \text{pOH} = 14.00 - 6.79 = 7.21 \]

7. **Calculate the pH of pure water at 25.0 and 50.0 °C (K_w = 1.0 \times 10^{-14} and 5.0 \times 10^{-14}, respectively)**

In pure water there is no extra H^+ or OH^- so [H^+] = [OH^-]. At the same time so [H^+] \times [OH^-] = K_w. From these two equations:

\[ [H^+] = \sqrt{K_w} \]

At 25°C, [H^+] = \sqrt{1.0 \times 10^{-14}} = 1.0 \times 10^{-7}, so pH = -\log(1.0 \times 10^{-7}) = 7.00

At 50°C, [H^+] = \sqrt{5.0 \times 10^{-14}} = 2.236 \times 10^{-7}, so pH = -\log(2.236 \times 10^{-7}) = 6.65

8. **10.00 ml of HCl of unknown concentration is titrated with potassium hydroxide whose concentration is 0.085 M. Calculate the pH at: a) 0%; b) 40%; c) 100%; d) 160% degree of titration. Up to the equivalence point 12.55 mL of KOH solution is consumed.**

a) **0%**

n(NaOH) = c \times V = 0.085 \times 0.01255 = 0.0010668 mol
n(NaOH) = n(HCl) = 0.0010668 mol

c(HCl) = \frac{0.0010668}{0.010} = 0.1067 \text{ M}
\[ \text{pH} = -\log[H^+] \]
\[ \text{pH} = 0.972 \]

b) **40%**

40% of HCl is titrated, 60% remains.

n(HCl) = 0.6 \times 0.001066 = 0.00064008 mol
V(sol) = 0.010 + 0.00502 = 0.01502 dm^3
n(NaOH) = 0.4 \times 0.0010668 = 0.00042672 mol
V(NaOH solution) = 0.00502 dm^3

c(HCl) = \frac{0.00064008}{0.01502} = 0.04262 \text{ M}
\[ \text{pH} = -\log[H^+] \]
\[ \text{pH} = 1.370 \]

c) **100%**

\[ \text{pH} = 7.000 \]

d) **160%**

V(solution) = V(HCl sample) + V(NaOH eqiu.) + V(NaOH extra)
\[ V(\text{solution}) = 0.010 + 0.01255 + 0.00753 = 0.03008 \text{ dm}^3 \]
\[ n(\text{NaOH}) = 0.00064008 \text{ mol} \]

\[ c(\text{NaOH}) = \frac{0.00064008}{0.03008} = 0.02128 \text{ M} \]

\[ \text{pOH} = -\log[\text{OH}^-] \]
\[ \text{pOH} = -\log (0.02128) \]
\[ \text{pOH} = 1.672 \quad \text{pH} = 14-1.672 = 12.328 \]

9. 20.00 ml of HCl of 0.100 mol/dm\(^3\) concentration is titrated with sodium hydroxide whose concentration is 0.200 mol/dm\(^3\). Calculate the pH at: a) 0 %; b) 20 %; c) 50 %; d) 120 % degree of titration.

a) 0 %

\[ c(\text{HCl}) = 0.100 \text{ mol/dm}^3 \]

\[ \text{pH} = -\log [\text{H}^+] \]
\[ \text{pH} = -\log 0.1 \]
\[ \text{pH} = 1.000 \]

b) 20 %

\[ n(\text{HCl}) = 0.80 \times 0.002 = 0.0016 \text{ mol} \]

The amount of NaOH needed for the neutralization of 20 % of the acid:
\[ n(\text{NaOH}) = 0.20 \times 0.002 = 0.0004 \text{ mol} \]
\[ V(\text{NaOH}) = \frac{0.0004}{0.200} = 0.002 \text{ dm}^3 \]

\[ V(\text{solution}) = V(\text{HCl sample}) + V(\text{NaOH added to the system}) \]
\[ V(\text{solution}) = 0.020 + 0.002 = 0.022 \text{ dm}^3 \]

\[ c(\text{HCl}) = \frac{n(\text{HCl})}{V(\text{sol})} = \frac{0.0016}{0.022} = 0.07273 \text{ mol/dm}^3 \]

\[ \text{pH} = -\log[\text{H}^+] \]
\[ \text{pH} = -\log 0.07273 \]
\[ \text{pH} = 1.140 \]

c) 50 %

\[ n(\text{HCl}) = 0.50 \times 0.002 = 0.001 \text{ mol} \]

The amount of NaOH needed for the neutralization of 50 % of the acid:
\[ n(\text{NaOH}) = 0.50 \times 0.002 = 0.001 \text{ mol} \]
\[ V(\text{NaOH}) = \frac{0.001}{0.2} = 0.005 \text{ dm}^3 \]
\[ V(\text{solution}) = 0.020 + 0.005 = 0.025 \text{ dm}^3 \]
\[ c(HCl) = \frac{n(HCl)}{V(sol)} = \frac{0.001}{0.025} = 0.04 \text{ mol/dm}^3 \]

\[ \text{pH} = - \log[H^+] \]
\[ \text{pH} = - \log 0.04 \]
\[ \text{pH} = 1.398 \]

d) 120 \%

After the equivalence point the pH of the solution is determined by the excess of NaOH.

\[ n(\text{NaOH}) = 0.020 \times 0.100 \times 1.20 = 0.0024 \text{ mol} \]
\[ n(\text{NaOH})_{\text{excess}} = 0.0024 - 0.002 = 0.0004 \text{ mol} \]

\[ V(\text{NaOH}) = \frac{0.0024}{0.200} = 0.012 \text{ dm}^3 \]

\[ V(\text{solution}) = 0.020 + 0.012 = 0.032 \text{ dm}^3 \]

\[ c(\text{NaOH}) = \frac{n(\text{NaOH})}{V(\text{NaOH})} = \frac{0.0004}{0.032} = 0.0125 \text{ mol/dm}^3 \]

\[ \text{pOH} = - \log [\text{OH}^-] \]
\[ \text{pOH} = - \log 0.0125 \]
\[ \text{pOH} = 1.903 \]

\[ \text{pH} = 14 - \text{pH} = 14 - 1.903 = 12.097 \]