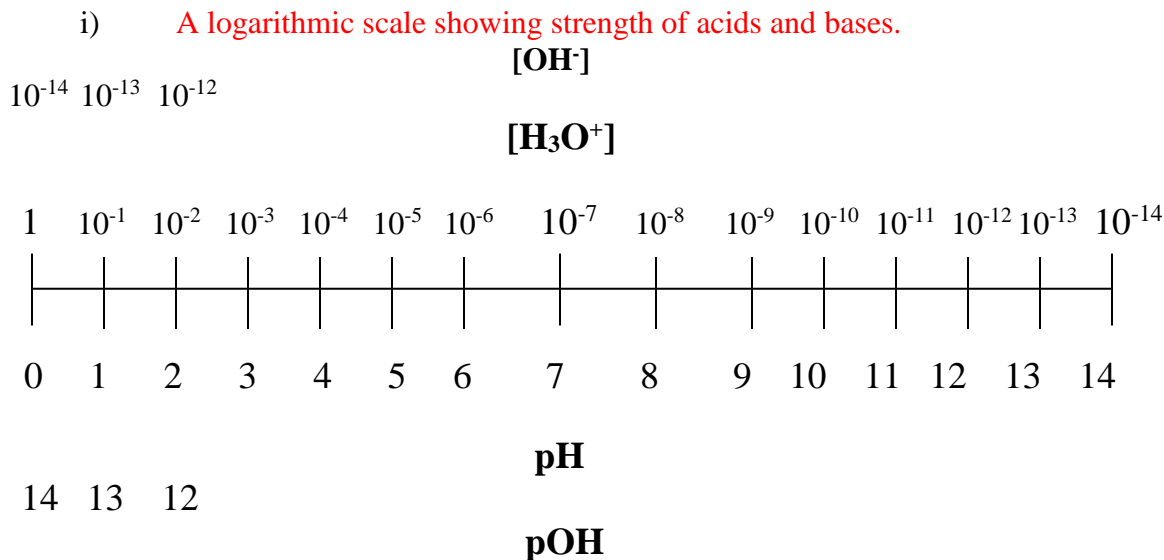


4.11 pH, pOH and pK Values

a) The pH Scale



ii) Every decrease in pH of 1 = Increase in [H₃O⁺] by 10

b) What is pH?

- i) Measure of [H₃O⁺] present in a solution
- ii) Solution is acidic when pH less than 7 (or when pOH greater than 7)

iii) **pH = -log[H₃O⁺]**

iv) What is pH when the [H₃O⁺] = 1.2 x 10⁻³ M?

$$\text{pH} = -\log(1.2 \times 10^{-3}) = \mathbf{2.92}$$

v) What is pH when the [H₃O⁺] = 4.8 x 10⁻⁸ M?

$$\text{pH} = -\log(4.8 \times 10^{-8}) = \mathbf{7.32}$$

vi) What is the [H₃O⁺] when the pH is 2.55?

$$[\text{H}_3\text{O}^+] = 10^{-2.55} \text{ or } [\text{H}_3\text{O}^+] = \text{antilog}(-2.55) = \mathbf{2.8 \times 10^{-3} \text{ M}}$$

vii) What is the [H₃O⁺] when the pH is 9.70?

$$[\text{H}_3\text{O}^+] = 10^{-9.70} = \mathbf{2.0 \times 10^{-10} \text{ M}}$$

c) What is pOH?

- i) Measure of $[\text{OH}^-]$ present in a solution
- ii) Solution is basic when pOH less than 7 (*pH greater than 7*)
- iii) **$\text{pOH} = -\log[\text{OH}^-]$**
- iv) What is pOH when the $[\text{OH}^-] = 1.5 \times 10^{-1} \text{ M}$?

$$\text{pOH} = -\log(1.5 \times 10^{-1}) = 0.82$$

- v) What is pOH when the $[\text{OH}^-] = 4.4 \times 10^{-4} \text{ M}$?

$$\text{pOH} = -\log(4.4 \times 10^{-4}) = 3.36$$

- vi) What is the $[\text{OH}^-]$ when the pOH is 12.65?

$$[\text{OH}^-] = 10^{-12.65} = 2.2 \times 10^{-13} \text{ M}$$

- vii) What is the $[\text{OH}^-]$ when the pOH is 1.70?

$$[\text{OH}^-] = 10^{-1.70} = 2.0 \times 10^{-2} \text{ M}$$

d) Relationship Between pH and pOH

- i) **$\text{pH} + \text{pOH} = 14$**
- ii) What is the pH of a solution if the pOH is 10.2?

$$\text{pH} = 14 - 10.2 = 3.8$$

- iii) What is the $[\text{OH}^-]$ if the pH is 3.25?

$$\text{pOH} = 14 - 3.25 = 10.75 \quad [\text{OH}^-] = 10^{-10.75} = 1.8 \times 10^{-11} \text{ M}$$

- iv) What is the pOH if the $[\text{H}_3\text{O}^+] = 1.7 \times 10^{-4} \text{ M}$?

$$\text{pH} = -\log(1.7 \times 10^{-4}) = 3.78 \quad \text{pOH} = 14 - 3.78 = 10.22$$

- v) What is the $[\text{H}_3\text{O}^+]$ if the $[\text{OH}^-] = 3.50 \times 10^{-5} \text{ M}$

$$\text{pOH} = -\log(3.50 \times 10^{-5}) = 4.456$$

$$[\text{H}_3\text{O}^+][\text{OH}^-] = K_w$$

$$\text{pH} = 14 - 4.456 = 9.544$$

$$\text{or} \quad [\text{H}_3\text{O}^+] = \frac{1.00 \times 10^{-14}}{3.50 \times 10^{-5}} = 2.86 \times 10^{-10} \text{ M}$$

$$[\text{H}_3\text{O}^+] = 10^{-9.544} = 2.86 \times 10^{-10} \text{ M}$$

e) pK Values

i) pK values are just for convenience!

ii) Observe the Pattern!

$$\text{pH} = -\log[\text{H}_3\text{O}^+] \qquad 6 = -\log[1.00 \times 10^{-6}]$$

$$\text{pKw} = -\log[\text{Kw}] \qquad 14 = -\log[1.00 \times 10^{-14}]$$

ii) **pKw = 14** ($\text{pH} + \text{pOH} = 14$ or $\text{pH} + \text{pOH} = \text{pKw}$)

iii) Observe the Pattern!

$$\text{pKa} = -\log[\text{Ka}] \qquad 2.12 = -\log[7.5 \times 10^{-3}]$$

$$\text{pKb} = -\log[\text{Kb}] \qquad 4.74 = -\log[1.8 \times 10^{-5}]$$

iv) **pKa + pKb = pKw**

f) Significant Figures

i) In a pH (or pOH) value, only the numbers after the decimal are significant

ii) Example:

pH = 2.465 has 3 sig. figs. The “2” give the power of 10...not significant.

iii) Example:

pH = 10.25 has 2 sig. figs.

iv) Example: $[\text{H}_3\text{O}^+] = 1.24 \times 10^{-3}$ M. What is pH?

$$\text{pH} = -\log(1.24 \times 10^{-3}) = 2.907$$

v) Example: $[\text{H}_3\text{O}^+] = 1.762 \times 10^{-6}$ M. What is pH?

$$\text{pH} = -\log(1.762 \times 10^{-6}) = 5.7540$$

g) Advanced pH and pOH Calculations

i) Example: 50.0ml of 0.200 M NaOH is reacted with 30.0ml of 0.250 M HCl. What is the pH of the resulting solution?



moles of acid or base in *excess* will determine the pH

② moles NaOH present = $0.200\text{M} \times 0.0500\text{L} = 0.0100$ moles

moles HCl present = $0.250\text{M} \times 0.0300\text{L} = 0.00750$ moles

③ NaOH is in excess by: $0.0100 - 0.00750 = 0.00250$ moles

④ $[\text{NaOH}] = [\text{OH}^-] = 0.00250 \text{ moles} / (0.0300\text{L} + 0.0500\text{L}) = 0.0312 \text{ M}$

⑤ $\text{pOH} = -\log[0.0312\text{M}] = 1.506$

⑥ $\text{pH} = 14 - \text{pOH} = 14 - 1.506 = \mathbf{12.494}$

ii) Example: Calculate the pH if 1.25 L of 0.300 M KOH is added to 0.500 L of 0.0900 M H₂SO₄.



moles of acid or base in *excess* will determine the pH

② moles $[\text{OH}^-]$ present = $0.300\text{M} \times 1.25\text{L} = 0.375$ moles

moles $[\text{H}_3\text{O}^+]$ present = $0.0900\text{M} \times 0.500\text{L} \times \mathbf{2} = 0.0900$ moles



cause each H₂SO₄
produces two H₃O⁺ 's

③ $[\text{OH}^-]$ is in excess by: $0.375 - 0.0900 = 0.285$ moles

④ $[\text{OH}^-] = 0.285 \text{ moles} / (1.25\text{L} + 0.500\text{L}) = 0.163 \text{ M}$

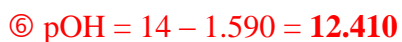
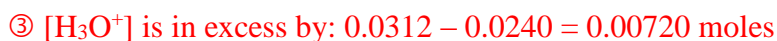
⑤ $\text{pOH} = -\log[0.163\text{M}] = 0.788$

⑥ $\text{pH} = 14 - 0.788 = \mathbf{13.212}$

iii) **Example: Calculate the pOH if 0.0300 L of 0.400 M Ca(OH)₂ is added to 0.250 L of 0.125 M HBr.**



moles of acid or base in *excess* will determine the pH



iv) **Example: How many grams of NaOH must be added to 0.800 L of 0.0400 M HBr to change the pH to 7.00? (Assume no volume change from adding NaOH)**

