

1. Isotonic solutions have the same osmotic pressure. The osmotic pressure, Π , is given by:

$$\Pi = MRT$$

where M is the molarity of the solution. Hence,

$$M = \frac{\Pi}{RT} = \frac{(8.3 \times 10^{-5} \text{ atm})}{(0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}) \times ((25 + 273) \text{ K})} = 3.4 \times 10^{-6} \text{ M}$$

As the pressure is given in units of atm, $R = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}$ has been used, leading to molarity in units of mol L^{-1} or M.

2. The osmotic pressure for strong electrolyte solutions is given by:

$$\Pi = i \times (MRT)$$

where i is the amount (mol) of particles in solution divided by the amount (mol) of dissolved solute.

For 1 M H_2SO_4 , $M = 1$ and $i = 2$ (as $\text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{HSO}_4^{2-}$, each mole of H_2SO_4 produces two moles of particles). Note that there is very little dissociation of HSO_4^- into H^+ and SO_4^{2-} at this pH. Hence, $\Pi = 2 \times 1 \times RT = 2RT$.

For 1 M HCl , $M = 1$ and $i = 2$ (as $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$, each mole of HCl produces two moles of particles). Hence, $\Pi = 2 \times 1 \times RT = 2RT$.

For 0.5 M glucose, $M = 0.5$ and $i = 1$ (as glucose dissolves without ionizing, each mole of glucose produces one mole of dissolved particles). Hence, $\Pi = 0.5 \times 1 \times RT = 0.5RT$.

For 0.5 M CaCl_2 , $M = 0.5$ and $i = 3$ (as $\text{CaCl}_2 \rightarrow \text{Ca}^{2+} + 2\text{Cl}^-$, each mole of CaCl_2 produces three moles of particles). Hence, $\Pi = 3 \times 0.5 \times RT = 1.5RT$.

For 0.5 M NaCl , $M = 0.5$ and $i = 2$ (as $\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$, each mole of NaCl produces two moles of particles). Hence, $\Pi = 2 \times 0.5 \times RT = RT$.

Therefore, in order of increasing osmotic pressure:

$$0.5 \text{ M glucose} < 0.5 \text{ M NaCl} < 0.5 \text{ M CaCl}_2 < 1 \text{ M HCl} \approx 1 \text{ M H}_2\text{SO}_4$$

3. Using osmotic pressure, $\Pi = MRT$,

$$M = \frac{\Pi}{RT} = \frac{(95 \times 10^{-3} \text{ kPa})}{(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times ((25 + 273) \text{ K})} = 3.8 \times 10^{-5} \text{ M}$$

As the pressure is given in Pa, $R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1} = \text{kPa L mol}^{-1} \text{ K}^{-1}$ is used requiring the conversion of 95 Pa to $95 \times 10^{-3} \text{ kPa}$. This leads to the molarity in mol L^{-1} .

Alternatively, the pressure can first be converted to atmospheres, using $1 \text{ atm} = 1.01325 \times 10^5 \text{ Pa}$,

$$\Pi = 95 \text{ Pa} = \frac{95}{1.01325 \times 10^5} \text{ atm} = 9.4 \times 10^{-4} \text{ atm}$$

Using $R = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}$ M,

$$M = \frac{\Pi}{RT} = \frac{(9.4 \times 10^{-4} \text{ atm})}{(0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}) \times ((25 + 273) \text{ K})} = 3.8 \times 10^{-5} \text{ M}$$

As concentration = $\frac{\text{number of moles}}{\text{volume}}$, the number of moles present in 1.00 mL is:

$$\begin{aligned} \text{number of moles} = n &= c \times V = (3.8 \times 10^{-5} \text{ mol L}^{-1}) \times (1.00 \times 10^{-3} \text{ L}) \\ &= 3.8 \times 10^{-8} \text{ mol} \end{aligned}$$

As the amount corresponds to 1.00 mg, the molar mass of the protein is given by:

$$\text{molar mass} = \frac{\text{mass}}{\text{number of moles}} = \frac{m}{n} = \frac{(1.00 \times 10^{-3} \text{ g})}{(3.8 \times 10^{-8} \text{ mol})} = 26000 \text{ g mol}^{-1}$$

4. The molar mass of salt (NaCl) is $(22.99 \text{ (Na)} + 35.45 \text{ (Cl)}) \text{ g mol}^{-1} = 58.44 \text{ g mol}^{-1}$.

59 g of salt therefore corresponds to:

$$\text{number of moles} = n = \frac{\text{mass}}{\text{molar mass}} = \frac{59 \text{ g}}{58.44 \text{ g mol}^{-1}} = 1.0 \text{ mol}$$

As this amount is dissolved in 1000 g of water, the molality is:

$$\text{molality} = \frac{\text{number of moles of solute (mol)}}{\text{mass of solvent (kg)}} = \frac{1.0 \text{ mol}}{(1000 \times 10^{-3} \text{ kg})} = 1.0 \text{ mol kg}^{-1}$$

The freezing point depression, ΔT_f , is given by:

$$\Delta T_f = i K_f m$$

where i is the amount (mol) of particles in solution, K_f is the molar freezing point depression or cryoscopic constant of the solvent and m is the molality of the solute. As NaCl dissolves to give two ions, $i = 2$. Hence,

$$\Delta T_f = 2 \times (1.86 \text{ K kg mol}^{-1}) \times (1.0 \text{ mol kg}^{-1}) = 3.8 \text{ K}$$

As pure water freezes at 273.15 K at atmospheric pressure, the solution will freeze at $(273.15 - 3.8) \text{ K} = 269.4 \text{ K}$.

5. The reaction table in terms of concentrations is:

	$\text{C}_3\text{H}_6\text{O}_3(\text{aq})$	\rightleftharpoons	$\text{H}^+(\text{aq})$	$\text{C}_3\text{H}_5\text{O}_3^-(\text{aq})$
initial	0.100		0	0
change	-x		+x	+x
equilibrium	$0.100 - x$		x	x

As the $[\text{C}_3\text{H}_6\text{O}_3]_{\text{dissociated}} = [\text{H}^+(\text{aq})]$, the percentage of $\text{C}_3\text{H}_6\text{O}_3$ which dissociate is given by:

$$\text{percentage dissociation} = \frac{[\text{C}_3\text{H}_6\text{O}_3(\text{aq})]_{\text{dissociated}}}{[\text{C}_3\text{H}_6\text{O}_3(\text{aq})]_{\text{initial}}} \times 100 = \frac{x}{0.100} \times 100 = 3.7 \%$$

Therefore,

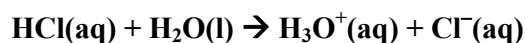
$$x = 3.7 \times 10^{-3} \text{ M}$$

Hence, at equilibrium, $[\text{C}_3\text{H}_6\text{O}_3(\text{aq})] = 0.100 - x = (0.100) - (3.7 \times 10^{-3}) = 0.096 \text{ M}$ and $[\text{H}^+(\text{aq})] = [\text{C}_3\text{H}_5\text{O}_3^-(\text{aq})] = x = 3.7 \times 10^{-3} \text{ M}$. The values of pH and K_a are therefore:

$$\text{pH} = -\log_{10}([\text{H}^+(\text{aq})]) = -\log_{10}(3.7 \times 10^{-3}) = 2,4$$

$$K_a = \frac{[\text{H}^+(\text{aq})][\text{C}_3\text{H}_5\text{O}_3^-(\text{aq})]}{[\text{C}_3\text{H}_6\text{O}_3(\text{aq})]} = \frac{(3.7 \times 10^{-3}) \times (3.7 \times 10^{-3})}{(0.096)} = 1.4 \times 10^{-4}$$

6. (a) Hydrochloric acid is a strong acid so the reaction,



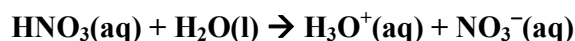
goes to completion, generating one mole of H_3O^+ for every mole of HCl.

0.14 M HCl therefore gives:

$$[\text{H}_3\text{O}^+] = 0.14 \text{ M}$$

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}[0.14] = 0.85$$

- (b) Nitric acid is a strong acid so the reaction,



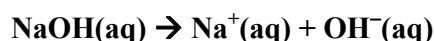
goes to completion, generating one mole of H_3O^+ for every mole of HNO_3 .

0.0025 M HNO_3 therefore gives:

$$[\text{H}_3\text{O}^+] = 0.0025 \text{ M}$$

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}[0.0025] = 2.60$$

- (c) Sodium hydroxide is a strong base so the reaction,



goes to completion, generating one mole of OH^- for every mole of NaOH .

0.048 M NaOH therefore gives:

$$[\text{OH}^-] = 0.048 \text{ M}$$

As $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$, this gives

$$[\text{H}_3\text{O}^+] = \frac{10^{-14}}{0.048} = 2.1 \times 10^{-13} \text{ M}$$

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}[2.1 \times 10^{-13}] = 12.7$$

- (d) Barium hydroxide is a strong base so the reaction,



goes to completion, generating *two* moles of OH^- for every mole of $\text{Ba}(\text{OH})_2$.

3.7×10^{-3} M $\text{Ba}(\text{OH})_2$ therefore gives:

$$[\text{OH}^-] = 2 \times (3.7 \times 10^{-3}) = 0.0074 \text{ M}$$

As $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$, this gives

$$[\text{H}_3\text{O}^+] = \frac{10^{-14}}{0.0074} = 1.4 \times 10^{-12} \text{ M}$$

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}[1.4 \times 10^{-12}] = 12$$

7. The titration is a 1:1 reaction between a strong acid and a strong base.

- (a) (i) Before any base is added, $[\text{H}^+(\text{aq})] = 0.100 \text{ M}$ and so $\text{pH} = -\log_{10}(0.100) = 1.00$.
(ii) As $\text{OH}^-(\text{aq})$ is added, the $[\text{H}^+(\text{aq})]$ decreases.

This is due to the acid-base reaction and due to the increase in volume.

Up until the equivalence point:

$$[\text{H}^+(\text{aq})] = \frac{n_{\text{H}^+}}{V} = \frac{n_{\text{H}^+}(\text{initial}) - n_{\text{OH}^-}(\text{added})}{V_{\text{total}}}$$

or

$$[\text{H}^+(\text{aq})] = \frac{(c_{\text{H}^+}(\text{initial}) \times V(\text{initial})) - (c_{\text{OH}^-}(\text{added}) \times V(\text{added}))}{(V(\text{initial}) + V(\text{added}))}$$

Thus, when 25.0 mL of 0.100 M NaOH has been added:

$$[\text{H}^+(\text{aq})] = \frac{(0.100 \text{ M} \times 0.050 \text{ L}) - (0.100 \text{ M} \times 0.025 \text{ L})}{((0.050 + 0.025) \text{ L})} = 0.033 \text{ M}$$

Hence, $\text{pH} = -\log_{10}(0.033) = 1.48$

(iii) When 45.0 mL of 0.100 M NaOH has been added:

$$[\text{H}^+(\text{aq})] = \frac{(0.100 \text{ M} \times 0.050 \text{ L}) - (0.100 \text{ M} \times 0.045 \text{ L})}{((0.050 + 0.045) \text{ L})} = 0.0053 \text{ M}$$

Hence, $\text{pH} = -\log_{10}(0.0053) = 2.28$

(iv) Addition of 50.0 mL of 0.100 M NaOH will exactly neutralize the acid.

Hence, $\text{pH} = 7.00$

(v) Beyond the equivalence point, the added base is simply being diluted as there is no acid left to react with:

$$[\text{OH}^-(\text{aq})] = \frac{n_{\text{OH}^-}}{V} = \frac{n_{\text{OH}^-}(\text{total}) - n_{\text{H}^+}(\text{initial})}{V_{\text{total}}}$$

As 50.0 mL of 0.100 M HCl was present initially, $n_{\text{H}^+}(\text{initial}) = 5.00 \times 10^{-3}$ mol.

55.0 mL of 0.100 M NaOH corresponds to 5.50×10^{-3} mol and when this has been added, the total volume is $(0.050 + 0.055) \text{ L} = 0.105 \text{ L}$. Hence

$$[\text{OH}^-(\text{aq})] = \frac{(0.00550 \text{ mol}) - (0.00500 \text{ mol})}{(0.105 \text{ L})} = 0.00476 \text{ M}$$

Hence, $\text{pOH} = -\log_{10}(0.00476) = 2.32$ and $\text{pH} = 14.00 - 2.32 = 11.7$

(v) When 75.0 mL of 0.100 M NaOH has been added, $n_{\text{OH}^-}(\text{total}) = 7.50 \times 10^{-3}$ mol and the total volume is $(0.050 + 0.075) \text{ L} = 0.125 \text{ L}$. Hence

$$[\text{OH}^- (\text{aq})] = \frac{(0.00750 \text{ mol}) - (0.00500 \text{ mol})}{(0.125 \text{ L})} = 0.0200 \text{ M}$$

Hence, $\text{pOH} = -\log_{10}(0.0200) = 1.70$ and $\text{pH} = 14.00 - 1.70 = 12.3$

- (b) Using these values, the pH curve for the titration can be constructed and is shown below.

