Module 3: Water

Further exercises

For pages 185–6

A1. Explain why lakes and rivers freeze from the top down and not from the bottom up. Answering the following questions may be the best way to do this. If the air above a shallow lake gradually cools from say 10^{o}C downwards, what happens to the temperature of the water at the surface of the lake? What mixing, if any does this cause? Why? Does there come a point at which this mixing stops? Why? As the air temperature cools to below 0^{o}C, where does ice form first? Does it stay there? Why? Data in Exercise 2 on page 190 may be useful. Note also that the density of ice at 0^{o}C is 0.917 g/mL.
What is the significance of this for fish?

A2. Some students made up solutions of sodium chloride in water of known composition (grams of NaCl per 100 g solution). They then pipetted 5.00 mL of each solution into a weighed beaker and determined the mass of the solution. The experiment was performed at 25^{o}C. Their results were:

Composition (g NaCl per 100 g solution)	5.00	10.00	15.00	20.00
Mass of 5.00 mL of solution (g)	5.16	5.32	5.54	5.73

(a) Draw a graph of density of solution versus per cent composition. They also knew that the density of pure water at 25^{o}C was 0.997 g/mL. Include this point on your graph also. How does density vary with composition?

(b) Sea water is 3.5% sodium chloride. What is its density at 25^{o}C?

(c) A sample of bore water had a density of 1.01 g/mL at 25^{o}C. Assuming that the only salt it contained was sodium chloride, estimate its per cent composition.

(d) Using your answer in (b), what happens to a ship when it moves out of fresh water into sea water? Does it sink further into the water, rise further out of the water or remain unaffected? Explain.

A3. (a) Use the data below to draw a graph of freezing point versus per cent composition (g NaCl per 100 g solution) for solutions of sodium chloride.

Composition	0	2.0	6.0	10.0	16.0	20.0	23.0
Freezing point (^{o}C)	0.0	–0.4	–1.6	–3.6	–9.0	–15.0	–21.2

(b) What is the freezing point of **(i)** sea water (3.5% NaCl) **(ii)** a 17.5% NaCl solution? What is the composition of a solution that freezes at **(iii)** 12.5^{o}C **(iv)** 22^{o}C?

(c) From your graph, what would be the freezing point of a 25% solution? How confident are you about your prediction? Why?

(d) In countries that have regular snowfalls leading to ice formation on roads, the roads are often sprinkled with salt to cause the ice to melt. Explain how this works.

A4. If a sample of pure water is boiled in an open beaker, the temperature remains constant throughout the process. However if a sample of sea water is similarly boiled in an open beaker the temperature gradually increases during the boiling process. Explain these different observations.

A5. Is clean drinking water a renewable resource? Explain. Relate your answer to the current efforts being made by Sydney Water to persuade people to use less water.

For page 190

B1. Draw electron-dot diagrams and structural formulae for the molecules formed from the following atoms:
(a) two fluorine atoms
(b) an oxygen and two chlorine atoms
(c) silicon and hydrogen atoms
(d) arsenic and fluorine atoms

B2. Which (if any) of the molecules in Exercise B1 is(are) polar? The shapes of these molecules are: (a) linear (b) bent (c) tetrahedral (d) pyramidal. Explain your reasoning.

B3. Place molecular chlorine, sodium chloride and trichloromethane, $CHCl_3$ in order of increasing boiling points. Give reasons for your predictions.

B4. Both the B–Cl and P–Cl bonds are polar. However boron trichloride is non-polar while phosphorus trichloride is polar. Offer an explanation for this.

For page 194

C1. The molecular weight of hydrogen fluoride is the same as the atomic weight of neon Give a reason for the boiling point of hydrogen fluoride being so much greater than that of neon (HF 19°C, Ne –246°C)

C2. Methanol, CH_3OH has an OH group and three H atoms attached to a central C atom; methanal, CH_2O, has an O atom and two H atoms attached to a central C atom. Which compound would you expect to have the higher boiling point? Explain why.

C3. What, if anything, is wrong with the following statement: despite having similar molecular weights (32 and 34 respectively) hydrazine, $H_2N–NH_2$, has a higher boiling point than phosphine, PH_3, because the chemical bonds in hydrazine are stronger than those in phosphine.

C4. Hydrogen peroxide, HOOH, is a liquid at room temperature whereas oxygen, O_2 or OO, is a gas; hydrogen peroxide is readily soluble in water while oxygen is only sparingly soluble. Offer explanations for these facts.

C5. Give an explanation for the difference in boiling points for the following pairs of substances.
(a) sulfur dichloride (59°C) and magnesium chloride (1437°C)
(b) Methanol, $H_3C–OH$ (64.7°C) and bromomethane, $H_3C–Br$ (3.6°C)

C6. (a) Using values from Table 7.2 on page 185 and others given here, list the following compounds in order of decreasing surface tension.
water, ethanol, $H_3C–CH_2–OH$, methanol, $H_3C–OH$ (23 mN m^{-1}), ethylene glycol, $HO–CH_2–CH_2–OH$, hexane, C_6H_{14} (18), pentane, C_5H_{12}, (16).
(b) What, if any, correlation is there between surface tension and magnitude of intermolecular forces in these compounds. Offer an explanation for any correlation that exists.

For page 200

D1. Draw diagrams on the molecular or ionic level of aqueous solutions of
(a) magnesium sulfate **(b)** urea, $H_2N–CO–NH_2$
Identify the ions, if any, in each solution.

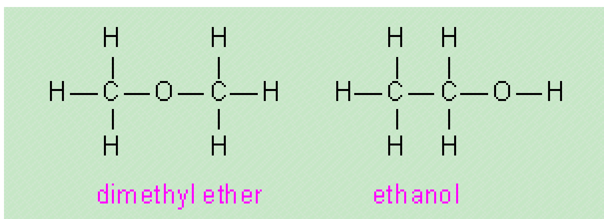
D2. Which of the following substances do you expect to be soluble in water? Give a reason for your prediction for each compound in the list.
calcium bromide, tetrafluoromethane, CF_4, nickel, ammonia, carborundum, SiC, nitrous acid, HONO, glucose, neon.

D3. Nickel chloride crystallises with six moles of water per mole of nickel chloride. Give the formula and name for the hydrated compound.
If 4.82 g of the hydrate were heated until all the water had been vaporised, what mass would remain?

D4. Magnesium chloride exists as a hydrate. When the compound is heated it loses its water of crystallisation. 2.64 g of this hydrate was heated to a constant mass of 1.24 g. Calculate the formula of the hydrate and name it.

D5. **(a)** If a sewing needle is very gently placed horizontally on the surface of water in a large beaker, the needle floats. Explain why, bearing in mind that the density of the needle is much greater than that of the water.
(b) However if one repeats the experiment using water that contains a small amount of detergent (surfactant), the needle does not float, no matter how carefully one does the experiment. Explain why.

D6. Dimethyl ether and ethanol both have the same molecular formula, C_2H_6O, but they have different structures:



Each dash represents a shared pair of bonding electrons.
Ethanol is very soluble in water whereas dimethyl ether is insoluble. Explain why.

For page 207

E1. Give the name and formula for the precipitate (if any) that forms when the following pairs of solutions are mixed:
(a) silver nitrate and potassium iodide
(b) aluminium chloride and sodium hydroxide
(c) ammonium nitrate and copper sulfate
(d) lead nitrate and sodium sulfide
(e) barium chloride and sulfuric acid
(f) sodium carbonate and potassium chloride
(g) potassium carbonate and iron(II) sulfate
(h) silver nitrate and calcium chloride
(i) zinc nitrate and barium hydroxide
(j) ammonium sulfide and iron(III) sulfate

E2. Write net ionic equations for all reactions that occurred in Exercise E1.

E3. **(a)** For the first, third and fifth equations you wrote in Exercise E2 rewrite them as complete ionic equations.
(b) For the second, fourth and eighth equations rewrite them as neutral species equations.

E4. How would you prepare samples of the following substances using precipitation reactions?
(a) lead sulfate **(b)** magnesium carbonate **(c)** copper sulfide

For page 212

F1. 100.0 g ammonium chloride was stirred with 200.0 g water for a considerable time then allowed to stand overnight. The solid remaining at the bottom of the container was filtered off, dried and weighed. It had a mass of 21.3 g. Calculate the solubility of ammonium chloride in g per 100 g water.

F2. The solubility of potassium nitrate is 37 g / 100 g water at 25^{o}C. If you mixed **(a)** 22 g potassium nitrate with 50.0 g water **(b)** 71 g with 200 g water, would all the solid dissolve? Explain.

F3. (a) Use the data below to draw a graph of solubility of potassium chlorate, $KClO_3$, as a function of temperature. Solubility is given as grams of solute per 100 g water.

Temperature (^{o}C)	10	30	50	70	80	90	100
Solubility (g / 100 g)	5	10	18	28	34	42	52

(b) What is the solubility of potassium chlorate at **(i)** 14^{o}C **(ii)** 87^{o}C?

(c) At what temperature is the solubility of potassium chlorate **(i)** 15 g / 100 g **(ii)** 48 g / 100 g?

(d) If 70 g $KClO_3$ was mixed with 200 g water at 90^{o}C, would a clear solution result? Why? If the mixture was cooled to 10^{o}C, what would you observe? Be as quantitative as possible. Explain your observation.

F4. Solutions of sodium hydroxide and calcium nitrate were mixed; the total volume was 100 mL. A precipitate of calcium hydroxide formed. After many hours analysis showed that 0.016 mol Ca^{2+} ions remained in solution. 2.00 g solid calcium hydroxide was shaken with 100 mL water. How many moles of calcium ion would you expect to be in the solution after several hours? Explain why. Why was a long time allowed before the analysis was performed in the first experiment and why could you only make a prediction for a long time after mixing the solid and water in the second experiment?

F5. The table below gives relative electrical conductivities (R.E.C.) for aqueous solutions (of suitable concentrations) of several substances.

Substance	sodium chloride	sucrose	*A*	*B*	*C*	*D*	*E*	*F*
R.E.C.	100	0.01	200	0.02	80	7	0.01	4

(a) First classify substances *A*, *B*, *C*, *D*, *E* and *F* as *electrolytes* or *non-electrolytes*. Then, for the electrolytes, decide whether they are strong or weak.

(b) State whether substances *A*, *B*, *C*, *D*, *E* and *F* would be present in aqueous solution as: **(i)** molecules **(ii)** ions **(iii)** mostly molecules but a few ions.

F6. The weak electrolytes in the table in Section 3 of the Supplementary material section above were acetic acid, CH_3CO_2H and ammonia NH_3. They were weak electrolytes because a small percentage of the molecules reacted with water to form H_3O^+ and $CH_3CO_2^-$ ions (from acetic acid) and NH_4^+ and OH^- (from ammonia). On the other hand nitric acid was a strong electrolyte, meaning that all of its molecules reacted with water to form H_3O^+ and NO_3^- ions. For these three ionisations write chemical equations that show the strength or weakness of the ionisations. Which (if any) of these ionisations is (are) an equilibrium reaction?

F7. Sodium hydroxide reacts with nitrous acid, HNO_2, to form sodium nitrite. Write an equation for this reaction. The relative electrical conductance of a nitrous acid solution was 12 (on a scale where hydrochloric acid was 250 and sodium hydroxide 150). After addition of just sufficient sodium hydroxide to convert all the nitrous acid to sodium nitrite, the relative electrical conductance was 90. Is **(a)** nitrous acid **(b)** sodium nitrite a strong or weak electrolyte? Explain how you decided this. What does this tell you about the species present in a solution of **(a)** nitrous acid **(b)** sodium nitrite?

For page 217

G1. Calculate the per cent composition of a saturated solution of sodium chloride which contains 36 g solute per 100 g water.

G2. **(a)** The solubility of Epsom salts (magnesium sulfate) in water is 0.36 g / g water. Express this as grams per kilogram of water and as mass percent, %(w/w).
(b) The solubility of copper sulfate is 220 g / kg water. Express this in g / g water and as mass per cent.
(c) Using an approximate density of the solutions of 1.0 g / mL, express the concentrations in (a) and (b) in grams per litre.

G3. **(a)** What mass of an 8.0% aqueous sodium chloride solution do you need to take in order to have 40 g NaCl?
(b) If an experiment requires 5.0 g sulfuric acid, what mass of a 1.0% solution of sulfuric acid in water should you take?
(c) A saturated aqueous solution of ammonium chloride at 25^{0}C contains 39 g solute per 100 g water. The density of the solution is 1.1 g /mL. Calculate the solubility in g / L.

G4. **(a)** To make 200 g of a 5% solution of sodium chloride in water, what mass of NaCl is needed?
(b) How much sucrose should be dissolved in 25 mL water to make a 15%(w/v) solution? Assume negligible volume change from water to solution upon mixing.

G5. **(a)** 12.3 g lead nitrate was dissolved in water and the volume make to 250.0 mL Calculate the concentration of the solution in g / 100 mL and in g / L.
(b) 10 mL (by pipette) of this solution was diluted to 500 mL (volumetric flask). Calculate the concentration of this diluted solution in g / L and in ppm. Assume the density of the diluted solution is 1.00 g/mL.

G6. **(a)** A solution was 24 ppm in magnesium ion. Calculate the concentration in g / L (assuming a density of 1.00 g/mL).
(b) What mass of copper sulfate pentahydrate would you dissolve in 2.00 L of water to make a solution that was 16 ppm in Cu^{2+}?

G7. If 2.00 mL of a 4.8%(w/v) solution of lead nitrate was diluted to 100 mL, what would the new concentration be? What volume of this diluted solution would you dilute to 1.00 L to make a solution that was 3.0 ppm in Pb^{2+} ion?

For pages 220–1

H1. Calculate the concentrations in mol/L of solutions that contain:
(a) 0.048 mol $Ca(OH)_2$ in 2.00 L
(b) 0.214 mol $MgSO_4$ in 250 mL
(c) 37.0 g KCl in 100 mL
(d) 3.97 g $Pb(NO_3)_2$ in 500 mL

H2. **(a)** How many moles of sodium hydroxide are needed to make 250 mL of a 0.500 mol/L solution?
(b) How many moles of ammonia are needed to make 100 mL of a 0.355 mol/L solution?
(c) What masses of the solutes are needed for the solutions in (a) and (b)?

H3. **(a)** What mass of sodium nitrate is required to make 500 mL of a 0.133 mol/L solution?
(b) What mass of barium hydroxide is required to make 2.00 L of a 4.5×10^{-3} mol/L solution?

H4. How many moles of solute are there in

(a) 37.5 mL of 0.207 mol/L Na_2CO_3

(b) 18.7 mL of 0.022 mol/L $Ca(OH)_2$

(c) 25 mL of 2.33×10^{-3} mol/L $AlCl_3$

(d) 0.250 L of 1.50 mol/L KI

H5. In the solutions in Exercise H4 how many moles are there of
(a) Na^+ ion in (a); **(b)** OH^- ion in (b); **(c)** Cl^- ion in (c)

H6. How many grams of solute are there in solutions (a) and (c) of Exercise H4?

H7. **(a)** How many moles of lead iodide are formed when 25 mL 0.266 mol/L potassium iodide is added to a solution containing excess lead nitrate?

(b) If a small excess of magnesium chloride solution is added to 50.0 mL 0.415 mol/L solution of sodium hydroxide how many moles of magnesium hydroxide precipitate out?

(c) What mass of **(i)** lead iodide **(ii)** magnesium hydroxide is formed in (a) and (b) respectively?

H8. Sodium sulfate solution was added to 100.0 mL of a solution containing lead ions until all the lead had been precipitated as lead sulfate. 0.855 g lead sulfate was formed. What was the concentration in moles per litre of lead ions in the original solution?

H9. 25.0 mL of sea water was accurately diluted to 250 mL. 50.0 mL of this diluted solution required 26.7 mL 0.114 mol/L silver nitrate solution to precipitate out all the chloride as silver chloride. Calculate the molarity of chloride in the sea water. Assuming that all the chloride had been present as sodium chloride, calculate the percentage (w/v) of sodium chloride in the sea water.

H10. To determine the concentration of cadmium ion in the waste water from an electroplating factory, an analyst took 250.0 mL of the waste water and slowly added sodium hydroxide solution until no more cadmium hydroxide precipitated. The precipitate was filtered off, washed and dried and weighed; it had a mass of 0.642 g. Calculate the concentration of cadmium ion in the original waste water in mol/L, g/100 mL and in ppm (assuming a density of 1.00 g/mL).

For page 224

J1. What quantity of heat is required to heat

(a) 45 g ethanol from 10^{o}C to 55^{o}C

(b) 32 g ethyl acetate through 25 K?

J2. If you spill 10 mL water at 60^{o}C on your hand it burns (stings, hurts) significantly more than if it had been 10 mL ethyl acetate at 60^{o}C. Use Table 8.4 on page 237 to explain why.

J3. If 100 g water at 50^{o}C is added with stirring to 250 g water at 20^{o}C in a well insulated container of negligible heat capacity, what will the final temperature be? (Hint: heat flows from the hot water to the cold water: the heat lost by the hot water equals the heat gained by the cold water. Let the final temperature be *x*; set up an equation for heat lost equals heat gained and solve for *x*.)

J4. If 100 g ethanol at 50^{o}C is added to 250 g water at 20^{o}C as in Exercise J3, what will be the final temperature? The hint in J3 applies here also.

For page 226

For Exercises K1 and K2 take the density and specific heat capacity of the final solutions as 1.00 g/mL and 4.2 J K^{-1} g^{-1} respectively and assume negligible heat losses.

K1. When 8.5 g sodium hydroxide was dissolved in 200 mL water at 18.2°C, the temperature rose to 28.1°C. Calculate the molar heat of solution of sodium hydroxide. Write a chemical equation for the process this refers to.

K2. The molar heat of solution of sulfuric acid is –90 kJ/mol. 10 mL (=18 g) concentrated sulfuric acid (take as 100%) was added to 100 mL water at 20°C. What was the final temperature?

K3. A convenient form of heat pack sometimes used by athletes consists of a plastic bag containing a supersaturated solution of sodium thiosulfate. To use it the athlete shakes the bag vigorously to cause the excess thiosulfate to crystallise out of solution. Using information in Exercise 34(b) on page 241, explain how this causes the bag to become hot. What would you do in order to be able to re-use this heat pack?
There is another type of heat pack that consists of a pouch of calcium chloride crystals inside a plastic bag of water. To use this the athlete ruptures the pouch containing the calcium chloride which causes the bag to get hot. Using information in Exercise 34(a) on page 241 explain why this bag of water becomes hot. Is this type of heat pack re-usable? Explain.

K4. **(a)** If a power station discharges 5×10^4 L of used cooling water at 60°C into a lake containing 2×10^6 L of water at 15°C, what will be the final temperature of the lake, assuming complete mixing and no heat losses? (See the hint for Exercise J3.)

(b) The assumptions in (a) are unrealistic. Describe what would be the real effect of discharging this hot water into the lake.