Oceans (O)

Revision guide

Clarity, confidence, and top grades.

What this storyline covers

- Dissolving and enthalpy changes
- Lattice and hydration enthalpies
- Entropy and feasibility
- Solubility products
- Brønsted–Lowry acids and bases
- Weak acids, K_a and pH
- Buffer solutions
- The greenhouse effect

Required practicals in this storyline: the enthalpy change of solution (**PAG3**) and K_a and the pH of acids and alkalis (**PAG11**), plus solubility-product and buffer experiments.



i How to use this guide

Oceans is the most calculation-heavy storyline. It brings together **energetics** (enthalpy cycles for dissolving), **entropy and feasibility**, **solubility products**, and **acid–base equilibria with pH**. Take each calculation type slowly and learn which formula applies; many ideas here build on the Hess cycles and equilibrium work from Elements of Life, Developing Fuels and The Chemical Industry.

How it is assessed

OCR B is context-led, so Oceans content can appear in **any** of the three written papers, in any context:

- **Paper 1** — Fundamentals of chemistry (110 marks)
- **Paper 2** — Scientific literacy, including the pre-release Advance Notice article (100 marks)
- **Paper 3** — Practical skills in chemistry (60 marks)

How to revise it

- Learn the **definitions in the teal boxes** word-for-word — they are marks you can guarantee.
- Work every **worked example** yourself before reading the solution.
- For each calculation, note the **formula and the units**; pH and entropy questions punish unit slips.
- Use the **spec checklist** on the last page to confirm you can do every point.

1 Dissolving and enthalpy changes

Spec references: O(a), O(b), O(c). Practical: PAG3.

Key definitions

- **Hydrated ions** are ions surrounded by water molecules, attracted by **ion–dipole bonds**.
- **Enthalpy change of solution ($\Delta_{sol}H$)** — the enthalpy change when **one mole** of solute dissolves to form a very dilute solution.
- **Lattice enthalpy ($\Delta_{LE}H$)** — the enthalpy change when one mole of solid ionic lattice forms from its **gaseous ions** (defined here as an **exothermic** quantity).
- **Enthalpy change of hydration ($\Delta_{hyd}H$)** — the enthalpy change when one mole of **gaseous ions** is dissolved in water to form hydrated ions.

Will it dissolve? O(a)

- Dissolving an ionic solid means **breaking the ionic lattice** and **hydrating the ions**. Whether it dissolves depends on the balance between the energy needed to separate the ions and the energy released when they are hydrated.
- Non-polar (non-aqueous) solvents cannot form strong **ion–dipole bonds**, so ionic solids are usually **insoluble** in them.

The enthalpy cycle for dissolving O(b)

Worked example — enthalpy of solution

From an enthalpy cycle, $\Delta_{sol}H = -\Delta_{LE}H + \Delta_{hyd}H$ (break the lattice, then hydrate the ions). For sodium chloride:

$\Delta_{LE}H = -780\ kJ\ mol^{-1}$; $\Delta_{hyd}H\ (Na^+ + Cl^-) = -770\ kJ\ mol^{-1}$

$\Delta_{sol}H = -(-780) + (-770) = +780 - 770 = +10\ kJ\ mol^{-1}$

Charge density O(c)

- The **greater the charge density** of an ion (higher charge, smaller radius), the stronger the electrostatic attraction.
- So a greater charge density makes the **lattice enthalpy more exothermic** and the **hydration enthalpy more exothermic**.

2 Entropy

Spec references: O(d), O(e), O(g).

Key definition

Entropy (S) is a measure of the number of ways that molecules and their energy quanta can be arranged — broadly, a measure of how dispersed the matter and energy are.

Predicting the sign of $\Delta_{sys}S$ O(e)

- Entropy increases in the order **solid < liquid < gas**: a gas has far more ways to arrange its particles and energy.
- A reaction that produces **more gas molecules** than it uses has a **positive** $\Delta_{sys}S$; one that produces fewer has a negative $\Delta_{sys}S$.

Calculating $\Delta_{sys}S$ O(g)

Worked example — entropy change of the system

$\Delta_{sys}S = \Sigma S(\text{products}) - \Sigma S(\text{reactants})$, in J K^{-1} mol^{-1}. For $N_2 + 3H_2 \rightleftharpoons 2NH_3$ (S: N_2 192, H_2 131, NH_3 192):

$\Delta_{sys}S = 2(192) - [192 + 3(131)] = 384 - 585 = -201$ J K^{-1} mol^{-1}

Negative, as expected: four moles of gas become two.

3 Entropy and feasibility

Spec reference: O(f).

Key relationships

- **$\Delta_{tot}S = \Delta_{sys}S + \Delta_{surr}S$** — the total entropy change.
- **$\Delta_{surr}S = -\Delta H / T$** — the entropy change of the surroundings (ΔH in J mol^{-1}, T in K).
- A reaction is **feasible** when **$\Delta_{tot}S$ is positive**.

Worked example — is it feasible?

For the reaction above, $\Delta H = -200$ kJ mol^{-1} and $\Delta_{sys}S = -201$ J K^{-1} mol^{-1} at 298 K. Convert ΔH to joules first.

$\Delta_{surr}S = -\Delta H / T = -(-200000) / 298 = +671$ J K^{-1} mol^{-1}

$\Delta_{tot}S = -201 + 671 = +470$ J K^{-1} mol^{-1} → positive, so the reaction is feasible

Exam tip — units

Entropies are in **J** K^{-1} mol^{-1} but ΔH is usually quoted in **kJ** mol^{-1}. Convert ΔH to joules (×1000) before using $\Delta_{surr}S = -\Delta H / T$, or the answer will be out by a factor of 1000.

4 Solubility products

Spec reference: O(h).

Key definition

The **solubility product (K_{sp})** of a sparingly soluble ionic compound is the product of the concentrations of its ions in a **saturated** solution, each raised to the power of its number in the formula.

- For a salt $M_aX_b \rightleftharpoons aM + bX$: $K_{sp} = [M]^a[X]^b$.
- Examples: for AgCl, $K_{sp} = [Ag^+][Cl^-]$; for $Mg(OH)_2$, $K_{sp} = [Mg^{2+}][OH^-]^2$.

Worked example — solubility product of AgCl

The solubility of silver chloride is 1.3×10^{-5} mol dm^{-3}. Each mole gives one Ag^+ and one Cl^-, so both ions are at 1.3×10^{-5} mol dm^{-3}.

$K_{sp} = [Ag^+][Cl^-] = (1.3 \times 10^{-5})^2 = 1.7 \times 10^{-10}$ mol^2 dm^{-6}

5 Acids and bases

Spec references: O(i), O(j).

Key definitions — Brønsted–Lowry

- An **acid** is a **proton (H^+) donor**; a **base** is a **proton acceptor**.
- When an acid donates a proton it forms its **conjugate base**; when a base accepts a proton it forms its **conjugate acid**. A conjugate pair differs by one H^+.

- Example: $HCl + H_2O \rightarrow H_3O^+ + Cl^-$. HCl is the acid and Cl^- its conjugate base; H_2O is the base and H_3O^+ its conjugate acid.

Strong acids and bases O(j)

- A **strong acid** ionises **completely** in water: $HCl \rightarrow H^+ + Cl^-$.
- A **strong base** dissociates completely: $NaOH \rightarrow Na^+ + OH^-$.

6 Weak acids, K_a and pH

Spec references: O(k), O(l). Practical: PAG11.

Key relationships

- A **weak acid** ionises only **partially**, setting up an equilibrium: $HA \rightleftharpoons H^+ + A^-$.
- **Acidity constant:** $K_a = [H^+][A^-] / [HA]$; **$pK_a = -\log_{10} K_a$**.
- **$pH = -\log_{10}[H^+]$**, and $[H^+] = 10^{-pH}$.
- **$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$** $mol^2\ dm^{-6}$ (given on the Data Sheet).

Worked example — three pH calculations

Strong acid — 0.10 mol dm^{-3} HCl (fully ionised, so $[H^+] = 0.10$):

$pH = -\log_{10}(0.10) = 1.00$

Strong base — 0.10 mol dm^{-3} NaOH (use K_w):

$[H^+] = K_w / [OH^-] = (1.0 \times 10^{-14}) / 0.10 = 1.0 \times 10^{-13}$; $pH = 13.00$

Weak acid — 0.10 mol dm^{-3} with $K_a = 1.7 \times 10^{-5}$ (using $[H^+] = \sqrt{(K_a \times [HA])}$):

$[H^+] = \sqrt{(1.7 \times 10^{-5} \times 0.10)} = \sqrt{(1.7 \times 10^{-6})} = 1.3 \times 10^{-3}$; $pH = 2.89$

7 Buffer solutions

Spec reference: O(m). Practical: buffer experiments.

Key definition

A **buffer** is a solution that **resists changes in pH** when a small amount of acid or base is added. It is made from a **weak acid and its salt** (its conjugate base), for example ethanoic acid and sodium ethanoate.

How a buffer works O(m)

- Both the weak acid **HA** and its conjugate base **A^-** are present in large amounts.
- **Add acid:** $A^- + H^+ \rightarrow HA$ — the conjugate base removes the added H^+.
- **Add base:** $HA + OH^- \rightarrow A^- + H_2O$ — the weak acid removes the added OH^-.
- So $[H^+]$ stays almost constant. Blood is buffered near pH 7.4 by a carbonic acid / hydrogencarbonate system.

Worked example — pH of a buffer

A buffer has $[HA] = 0.20$ and $[A^-] = 0.10$ mol dm^{-3}, with $K_a = 1.7 \times 10^{-5}$. Rearrange the K_a expression:

$[H^+] = K_a \times [HA] / [A^-] = (1.7 \times 10^{-5}) \times (0.20 / 0.10) = 3.4 \times 10^{-5}$

$pH = -\log_{10}(3.4 \times 10^{-5}) = 4.47$

8 The greenhouse effect

Spec reference: O(n).

- Solar energy reaches the Earth mainly as **visible and ultraviolet** radiation.
- The Earth **absorbs** some of this, warms up, and radiates energy back as **infrared (IR)**.
- **Greenhouse gases** (such as carbon dioxide and methane) in the troposphere **absorb some of this IR**, in the 'IR window' that the atmosphere would otherwise let escape.
- Absorbing IR makes the **bonds vibrate more**; this energy is passed to other molecules by **collisions**, raising their kinetic energy and the temperature.
- Greenhouse gas molecules also **re-emit** IR in all directions, some back towards the Earth, warming it further.
- **Higher concentrations** of greenhouse gases lead to an **enhanced greenhouse effect**.

In a nutshell

Sun in as visible and UV; Earth out as IR; greenhouse gases catch the IR, vibrate, and share that energy by collisions and re-emission — so the lower atmosphere warms.

P Required practicals in this storyline

These PAG activities are most likely to be examined in Paper 3, but the techniques can appear in any paper.

Activity	What to be able to do
Enthalpy of solution (PAG3)	measure the temperature change when a known amount of solid dissolves, and use q = mcΔT to find $\Delta_{sol}H$ per mole
K_a and pH (PAG11)	measure pH with a calibrated pH meter; find the K_a of a weak acid from its pH and concentration
Solubility products	carry out experiments to determine the solubility product of a sparingly soluble salt
Buffer solutions	make a buffer and show how its pH resists change when small amounts of acid or base are added

Exam tip — choosing the right pH method

Strong acid: $[H^+]$ equals the acid concentration. **Strong base:** find $[H^+]$ from K_w and $[OH^-]$. **Weak acid:** $[H^+] = \sqrt{(K_a \times [HA])}$. **Buffer:** $[H^+] = K_a \times [HA] / [A^-]$. Always finish with $pH = -\log_{10}[H^+]$.

Spec coverage checklist — can you do every point?

Every assessable learning outcome for Oceans, with where it is covered in this guide. Tick each one once you can do it confidently.

O(a)	Factors determining the solubility of a solute in different solvents — §1
O(b)	Enthalpy of solution, lattice enthalpy and hydration enthalpy; enthalpy cycles and calculations — §1
O(c)	Dependence of lattice and hydration enthalpies on charge density — §1
O(d)	Entropy as the number of ways of arranging molecules and energy — §2
O(e)	Qualitative predictions of the sign of $\Delta_{sys}S$ — §2
O(f)	$\Delta_{tot}S = \Delta_{sys}S + \Delta_{surr}S$, $\Delta_{surr}S = -\Delta H/T$, and feasibility — §3
O(g)	Calculating $\Delta_{sys}S$ from entropies of reactants and products — §2
O(h)	Solubility product and its calculations — §4
O(i)	Brønsted–Lowry acids and bases; conjugate pairs — §5
O(j)	Strong acids and bases and their ionisation in water — §5
O(k)	Weak acids, the acidity constant K_a and pK_a — §6
O(l)	pH calculations for strong acids, strong bases and weak acids — §6
O(m)	Buffer solutions: meaning, how they work and calculations — §7
O(n)	The greenhouse effect — §8