Water is distributed on Earth as a solid, liquid and gas

- Solute: the substance that is dissolved
  Solvent: the substance that dissolves the solute
  Solution: a mixture of a solute dissolved in a solvent in which particles are evenly distributed as a single particles
- The importance of water as a solvent
  - Allows biological processes to occur in aqueous solutions
  - Serves as a transport system for nutrients and waste products in living things (blood, lymph systems) → gas exchange of O_2 and CO_2
  - Dissolves oxygen and carbon dioxide → important for aquatic animals and all plants; also modifies the greenhouse effect
  - Used as a base in cleaning products, paints, etc.
- Water on earth

	Biosphere	**Lithosphere**	**Hydrosphere**	**Atmosphere**
% of water	70%	Variable	96-100%	0-5%
State of water	Liquid Water of crystallisation Solid ice	Liquid Solid ice	Liquid	Gas

- Water for life
  - Essential reactants in the production of glucose by process of photosynthesis, as a solvent of carbon dioxide
  - A product of cellular respiration as a decomposer of carbon dioxide → provides the energy required to sustain life
  - A solvent that dissolves oxygen, various salts and nutrients
  - A solvent of waste products, such as carbon dioxide and urea in plants and animals
  - A life-maintaining transport system as a major constituent of blood. Also transports essential nutrients and waste products
  - A major component of the lymph system and respiratory system as the moisture lining our lungs, and the diffusing of the gases of oxygen and carbon dioxide
  - Respiration through the skin → transfers heat around the body
- Water moderates temperature
  - The large heat capacity of the ocean has a significant moderating effect on the temperatures experienced in regions close to the coast
  - Water can absorb large quantities of heat energy with relatively small temperature changes. The moderation of temperature is a direct result of the ability of water to absorb and re-radiate large quantities of heat energy. The quantity of heat energy required to convert a particular amount of water from the liquid to gaseous state
  - Has a large heat of vapourisation
  - The process of evaporation absorbs heat energy and therefore helps to keep the human body at its natural temperature
- Weathering of the lithosphere
  - Extreme temperatures → expansion of water as it freezes in joints and cracks of rocks causes weathering
  - Rainfall – shapes the environment and determines the vegetation that grows
  - Slow movement of glaciers → gradual grinding away of underlying rocks → produces rich soils
  - Weathering occurs through the constant motion of water carrying suspended matter and constantly wearing away the rock surface
  - Chemical weathering – water reacting with some minerals in rocks as a weak acid erosion → from beaches into oceans
- Water for humans
  - Everyday uses: drinking, washing, growing crops, industrial processes
  - Transport system: oceans and rivers move great quantities of food, raw materials and consumer products around the planet
  - Source of entertainment and enjoyment: fishing, swimming, sailing and other water sports

- <u>Differing densities of ice and water:</u> In solid ice state, the bent shape of the polar water molecules allows the two positive hydrogen atoms, joined to the negative oxygen atom by dipole-dipole forces, to form hydrogen bonds with two negative oxygen atoms of neighbouring water molecules. These water molecules form a tetrahedron shape, creating a highly-ordered, open-cage structured water crystal made up of hexagonal rings with empty space inside. When ice melts, the rise in temperature allows the molecules to gain kinetic energy to move so that their positions become more random and they are able to pack more closely together. Thus, the volume of the molecules is decreased. As density = mass/volume, and with mass remaining constant, the density of the molecules in liquid state is greater than that in solid ice. The hydrogen bonds between them, however, are still strong enough to prevent the molecules from completely separating and escaping into the gaseous state.

Practicals

<u>Investigating the density of water</u>

- As a liquid:

Apparatus: 100mL beaker, electronic balance, distilled water, 10mL pipette, pipette filler

Method:

1. Accurately measure the mass of the empty beaker. Record
2. Add a known volume of distilled water using the pipette to the beaker and reweigh. Record.
3. Repeat until there are five measurements, using 10mL increments from the starting volume. Record all measurements

Results: plotted on a mass of water (g) vs. volume of water (mL) graph, then gradient used to determine density = 0.96 g/mL

- As a solid:

Apparatus: 500mL cylinder, ice cube, distilled water, straw, electronic balance

Method:

1. Fill the cylinder up to a suitable volume (eg. 210mL)
2. Measure mass of cube of ice. Record
3. Drop ice into cylinder, holding it just beneath the surface of the water with a straw. Record how much the ice cube has displaced the water. This is a measurement of its volume.
   Note: 1mL = 1g
4. Repeat procedure until there are five measurements

Results: Find density for each result of mass and volume of ice and average to find average density = 0.98g/mL

Discussion

Errors: accuracy of instruments, ice didn’t stay underneath water, ice caused water to spill of out cylinder when it was dropped into it, ice melted before it was placed in water

Improvements: perform experiment as fast as possible

<u>The effect of salt on the boiling point of water</u>

Apparatus:

2. The wide distribution and importance of water on Earth is a consequence of its molecular structure and hydrogen bonding

- Water:

  Ammonia:

  Hydrogen sulfide:

- <u>Molecular structure:</u>

Substance	Molecular structure	Shape	Melting point (°C)	Boiling point (°C)
H_2O		Angular, V-shaped or bent	0	100
H_2S		Angular, V-shaped or bent	–83	–62
NH_3		Pyramidal	–78	–33

- Hydrogen bonding: A strong form of a dipole-dipole attractive force, which can only occur between a relatively low electronegative hydrogen atom and a highly electronegative atom, which is oxygen in a water molecule. The highly electronegative oxygen atom of the molecule strips the hydrogen atom of its only electron and so the hydrogen atom virtually becomes a partially positive charge. The partially positive hydrogen atom then attracts an unshared electron pair of an oxygen atom in a neighbouring molecule.
- Water molecules as a polar molecule: the electron pair shared in the covalent bond between the oxygen and hydrogen atoms is not equally shared by the atoms. It spends more time in the vicinity of the oxygen atom, which has a higher electronegativity, or ability to attract electrons. Thus, there is an uneven charge distribution. This leads to the hydrogen atoms acquiring a small positive charge, $\delta+$ and the oxygen atom acquiring a small negative charge, $\delta-$. Thus, forming two O–H polar bonds, or dipoles. The vector sum of these two dipoles produces a net molecular dipole. Thus, the water molecule is a polar molecule.
- Dipole-dipole attractive forces: occurs between the polar water molecules. The positive hydrogen end or pole of one molecule is attracts the negative oxygen end or pole of a neighbouring molecule, and the negative oxygen end or pole of one molecules attracts the positive hydrogen end or pole of another molecule.
- Properties of water:
  - *Cohesion:*
    - Forces that act between 'like' particles eg. hydrogen bonding, dispersion forces and dipole-dipole forces
    - Strong bonding, strong forces of cohesion → eg. water – forms spherical droplets
  - *Adhesion:*
    - Results from the forces of attraction between unlike particles eg. forces between water molecules and the glass beaker holding it, or forces between water droplets and a solid surface
    - Competition between intermolecular forces within a liquid (cohesive forces) with those between a liquid and another surface (adhesive forces) → eg. hydrogen bonds between water molecules and glass (mainly silicon dioxide) are even stronger than the hydrogen bonds between water molecules in water itself → water has a concave meniscus
  - *Surface tension:*
    - Water molecules at surface of beaker of water are not surrounded by other water molecules → have overall force downwards into the rest of the water → creates tension on the surface of the water, so that it behaves like a tightly stretched skin → surface tension

        - Directly proportional to the strength of the forces between the particles of the liquid → water has a relatively strong surface tension
    - *Viscosity:* how easily a fluid flows → the resistance of a fluid to flow → slower, higher viscosity
        - Relates directly to the strength of the forces between the particles of the fluid and to the size of the particles → determine how easily the molecules of the fluid move past each other → stronger intermolecular forces (i.e. In water), high viscosity
        - Temperature increases, viscosity decreases → increased motion of particles
    - *Melting and boiling points:*
        - Stronger bonds (eg. hydrogen bonds between small polar molecules of water), higher melting and boiling points
        - Ice crystal – each molecule hydrogen bonded to four other water molecules → when ice melts, heat energy is provided to increase the kinetic energy of the molecules → they break free from hydrogen bonds
    - *Hardness and brittleness:*
        - Water (ice) – intermediate hardness – intermolecular hydrogen bonds:
            - → Strong than normal covalent bonds
            - → Weaker than ionic bonds and covalent network structures
        - Solid water (ice) is very brittle → hydrogen bonds are highly directional → when ice crystal is subjected to a force, it cannot deform its shape shatters → hydrogen bonds holding lattice are broken

Practicals

Investigating some of the properties of water

- Surface tension:
    - Use an eyedropper to determine how many drops of water can be placed on a 5 cent coin without the water flowing over. Repeated separately with a detergent-water mixture and methylated spirits → least to most drops (lowest to high surface tension): detergent-water mixture, methylated spirits, water
    - Carefully add paper clips one at a time to a test tube with water nearly overflowing, until the water overflows. Count how many can be added. → Water level rose until it overflowed
- Viscosity: Cover a 30cm long piece of board with a plastic sleeve. Draw a starting line at the top of the plastic and a finishing line at the end of the plastic. Place board at an angle of 45°. Place a drop of water at the starting line, using an eyedropper. Time how long it takes for the drop to move to the bottom, using a stopwatch. Repeat with glycerol, honey, acetone and lubricating oil. → fastest to slowest (lowest to highest viscosity): water, lubricating oil, acetone, glycerol, honey

3. Water is an important solvent

- Solubility of ionic compounds: In solutions where ionic compounds are soluble in water, charged ions are surrounded by water molecules
    - For cations, the surrounded water molecules are orientated with the negatively charged ends of the water molecules directed towards the positive ion
    - For anions, the positive end of the dipole is directed towards the ion
- Solubility of covalent molecular substances: Polar substances that are able to form hydrogen bonds with water are soluble, followed by polar substances that do not form hydrogen bonds, followed by non-polar substances

- Polar covalent molecular compounds – intermolecular forces between these substances and water are similar to those within in the separate solute and solvent (water)
- Covalent molecular substances unable to form hydrogen bonds with water molecules will only form dipole-dipole or dispersion forces with water → weaker than hydrogen bonds that form between water molecules in the solvent → substances is partially or slightly soluble
- Low solubility of covalent molecular gases is due to weak intermolecular forces, dipole-dipole forces and/or dispersion forces between gas and water molecules

o <u>Insoluble large molecules</u>: Very strong covalent bonds that form those large molecules cannot be broken by the weaker intermolecular forces (hydrogen bonds, dispersion forces) that could possibly be formed with the water molecules
  - large non-polar portion
  - need to break vast numbers of hydrogen bonds between the solvent water molecules

o <u>Solubility of covalent network substances:</u> Insoluble because the very strong covalent bonds which form the crystal lattices of these substances cannot be broken by the weaker intermolecular forces that would be formed with water molecules

o <u>Ionisation (of covalent molecular substances):</u> ions are formed due to the reaction with water
<u>Dissociation (of ionic compounds):</u> ions are already present in the solid and separate to move into the solution

Practicals

<u>Comparing the solubility of different substances in water</u>

Aim: To determine what types of substances are soluble in water

Apparatus: test tubes (one for each solute), eyedroppers, solutes

Results: water soluble substances were generally ionic or polar covalent substances; water insoluble substances were generally non-polar molecular covalent, network covalent

4. The concentration of salts in water will vary according to their solubility and precipitation can occur when the ions of an insoluble salt are in solution together

o <u>Solubility rules for ionic substances:</u> all sodium, potassium and ammonium salts are soluble

Soluble Anions	**Exceptions**
NO_3	None
Cl	Ag insoluble, Pb slightly soluble
Br	Ag insoluble, Pb slightly soluble
I	Ag , Pb insoluble
SO_4	Ba , Pb insoluble, Ca , Ag slightly soluble
Insoluble Anions	**Exceptions**
OH	Na , K , Ba soluble; Ca slightly soluble Note: NH_4OH exists as NH_3(aq)
S	Na , K , NH_4 soluble
CO_3	Na , K , NH_4 soluble
PO_4	Na , K , NH_4 soluble

o <u>Movement of ions when solution and precipitation occur:</u>
Eg. $AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

o <u>Spectator ions:</u> ions which are not involved in chemical reactions and remain as ions in solution

o <u>Ionic equations – rules:</u>
  - Strong electrolytes are written in ionic form
  - Weak electrolytes and non-electrolytes are written in molecular form

  - Insoluble substances are written as the formulas of the substances
  - Gases are written in molecular form
  - Net ionic equations should not include spectator ions
- <u>Reversible reactions and equilibrium systems:</u> Forward and reverse reactions happen at the same time, for example:
  - Excess solution in contact with saturated solution, eg. $AgCl_{(s)} = Ag_{(aq)} + Cl_{(aq)}$ → some of the solute continues to dissolve into the solution at the same rate as dissolved solute crystallises out of solution
  - Water in a closed system: without lid: rate of evaporation > rate of condensation; with lid further down: rate of condensation > rate of evaporation → eventually rates would once again be equal and equilibrium would be re-established → macroscopic properties would remain constant

  When the forward and backward reactions are going at the same rate, a ***<u>chemical equilibrium</u>*** is achieved → *dynamic equilibrium*
  <u>Note:</u> saturated – all the solute has dissolved
- <u>Concentration/molarity:</u> the number of moles of solute per litre of solution (mol/L)

  $$c = \frac{n}{V}$$
- <u>Other measurements of concentration:</u>
  - Mass per unit volume (m/V) eg. g/L: the mass of the solute in each volume unit of solution → used to describe how soluble a material is in a solvent
  - Percentage composition by mass: the mass of the solute, measured in grams, dissolved in 100g of solution, calculated as a percentage → used for solids dissolved in a solid
  - Parts per million (ppm): the mass of the solute, measured in milligrams, in one kilogram of solution → used for small concentrations
- <u>Importance of different measurements of concentrations:</u> A variety of units of concentration is necessary, instead of the terms 'dilute' and 'concentrated', in order to quantitatively describe the exact concentration of solutions made from several different chemical mixtures. This is important in determining how much solute or solvent must be added or removed to dilute or concentrate a solution used in chemical analysis, chemical manufacturing and the manufacture of commercial products, such as drugs and medicines. It determines how safe it is to handle chemicals in these situations and allows for accurate chemical reactions.
- <u>Dilutions:</u> A definite volume of the concentrated solution is measured out with a pipette and placed into a volumetric flask, sufficient solvent is then added to the flask to make to solution up the calibrated mark → does not alter the number of moles of solute present

  $$C_1V_1 = C_2V_2$$

Practicals

<u>Solubility of ionic compounds</u>

Method: Used an eyedropper to combine different solutions with different anions and cations to see if they formed a precipitate

Discussion: Some of the solutions contained impurities so they had precipitates in them before they were used.

<u>Making and diluting specific solutions</u>

Apparatus:

100mL volumetric flask
50mL volumetric flask
10mL graduated pipette
Pipette filler
$KMNO_4$ (potassium permanganate)
Small clean beaker
Pasteur pipette
Distilled water
Stirring rod
Spatula
Wash bottle

Method:

1. Weigh 1g of $KMNO_4$
2. Transfer to beaker
3. Add approx. 20-30mL of water and stir
4. Transfer to 100mL volumetric flask
5. Use the wash bottle to rinse out the beaker into the flask
6. Add distilled water to just below the mark on the flask
7. Use Pasteur pipette to add water right up to the mark
8. Place stopper in flask and invert flask several times
9. Using the pipette filler and 10mL graduated pipette, draw up to 5mL of the original solution
10. Transfer to 50mL volumetric flask
11. Repeat steps 6-8

Results: Find number of moles, concentration and dilution of $KMNO_4$

5. Water has a higher heat capacity than many other liquids

- Specific heat capacity: the amount of heat energy required to change the temperature of 1g of a substance by 1 kelvin → units: J/g/K
  Note: K = C° + 273

Substance	Specific heat capacity (J/g/K)
Water	4.18
Ethanol	1.41
Toluene (methylbenzene)	1.13
Phenol (hydroxybenzene)	1.11
Benzene	1.05
Aluminium	0.90
Chloroform	0.55
Carbon tetrachloride	0.54
Glass	0.50
Iron	0.45
Silver	0.23
Mercury	0.14

- $\Delta H = mC\Delta T$
  where H = enthalpy – heat energy absorbed or released (heat content), J
  m = mass of substances (g)
  C = specific heat capacity of substance
  T = temperature change of the substance ($T_f - T_i$), K
- Describing energy changes in chemical reactions:

  $\underbrace{\Delta H}_{\text{heat of reaction}} = \underbrace{H\text{ (products)}}_{\text{enthalpy of products}} - \underbrace{H\text{ (reactants)}}_{\text{enthalpy of reactants}}$

  - Negative for exothermic reactions – release energy causing surroundings to heat up eg. methane burnt in air
  - Positive for endothermic reactions – absorb energy from surroundings, causing them to cool down eg. carbon reacting with sulfur
- Enthalpy profile diagrams:
  Exothermic reaction: Endothermic reaction:

- Thermal pollution: the harmful rise in the temperature of water bodies, such as streams, rivers, lakes and oceans.
  - Major sources:

    - Industries, such as power plants and factories, which use water as steam to turn turbines in power generators and as a coolant to condense the steam. The water is then returned to the water body warmer than it was when it was drawn.
    - Deforestation and erosion, in which vegetation near the water body, which shades it and makes it cool, is removed, permitting sunlight to warm the water. This also leads to erosion of soil into the water, which makes it muddy. When the water is muddy, it absorbs more heat energy from the sun, therefore leading to further heating of the water. As warm water is denser than cool water, it forms a barricading layer on the surface that blocks the absorption of oxygen by the cooler water.
- Implications for life:
    - Reduction in the concentrations of dissolved gases eg. O2 and CO2, in water → higher temperatures, lower solubility of gases → insufficient to sustain particular life forms
    - Dead tissue of organisms use up oxygen to decomposes, further depleting the oxygen supply of the water
    - Organisms that depend on these dead organisms for food are also affected

Practicals

Finding molar heat of dissolution

Apparatus:

100mL measuring cylinder
electronic balance
spatula
thermometer
foam cup

Safety precautions

- Sodium hydroxide dissolves in water to form a very strong alkaline solution. It causes a burning sensation on skin and pollution. Take care when handling and do not dispose of down sink
- Ammonium chloride produces fumes of ammonium gas when dissolved in water. Do not breathe in
- Both substances are very damaging to the eyes. Wear safety goggles

Method:

1. Weigh 5g of NaOH on filter paper and record exact mass
2. Measure temperature of 100mL water. Record
3. Stir/dissolve NaOH into water
4. Observe changes in temperature
5. Record minimum/maximum temperature on thermometer
6. Use $\Delta H = mC\Delta T$ to find the enthalpy of water
7. convert the mass of solute into moles using $n = \frac{m}{M}$
8. Find molar heat of dissolution using $\frac{\Delta M}{n}$
9. Express answer as J/mol or kJ/mol
10. Repeat with ammonium chloride, NH_4Cl

Results:

- $\Delta H_{dissolution}$ of NaOH = negative number → exothermic reaction
- $\Delta H_{dissolution}$ of NH_4Cl = positive number → endothermic reaction

Conclusion:

- Heat escapes from cup to atmosphere/environment → not a perfect insulator → improve insulation of calorimeter and how much heat it absorbs eg. put lid on
- Some NaOH and NH_4Cl stayed stuck to the filter paper
- Compare theoretical values with values obtained to determine accuracy of experiment