1 # Matter & Energy

EVPP 110 Lecture - Fall 2002

2 ## Chapter Sections

- ✓Atoms & molecules
- ✓Properties of Water
- ✓Rearrangements of Atoms

3 **Atoms**

- ✓much of what we know about the ecosphere comes from a reductionist approach
  - – a scientific approach based on the idea that
    - – the whole can best be understood by studying its parts
    - – properties of the ecosphere emerge from the organization of the parts into non-living and living components.

4 **Atoms**

- ✓to understand chemical structure and function in the ecosphere we must
  - – start small and see how the structures at each level are combined into each higher level
    - *note the hierarchical aspect of this idea*

5 **Atoms**

- ✓the matter in all non-living and living things is made up of chemicals
- ✓at every level, structure and function are strongly interrelated

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- ✓All substances, non-living and living, are composed of **matter**
- ✓matter is composed of chemical **elements**

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- ✓**Matter** is anything that occupies space and has mass
  - – **mass** – the amount of a substance
    - is the same regardless of location
    - mass on earth = mass on moon
  - – **weight** – the force gravity exerts of a substance

- varies with gravitation changes
- mass on earth > mass on moon

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✓chemical **element**

– a substance which cannot be broken down into any other substance by ordinary chemical means

– each element consists of one type of **atom**

– an element consists of atoms with the same atomic number and same chemical properties

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✓**naturally occurring elements**

– 92 naturally occurring elements are recognized
  - ranging from hydrogen, the lightest , to uranium, the heaviest

✓**man-made elements**

– an additional 12 - 17 man-made elements have been created in the lab
  - by bombarding elements with subatomic particles in devices known as particle accelerators

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✓distribution of elements (non-living vs. living)

– in crust of earth (non-living)
  - 9 elements constitute ~99% (by mass) of the earth's crust
    – except for oxygen, the most common elements inside organisms are not the elements most abundant in earth's crust
      » for example, silicon, aluminum and iron constitute ~40% of the earth's crust but are found only in trace amounts in the human body

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✓distribution of elements (non-living vs. living)

– in living organisms
  - of the 92 naturally occurring elements
    – about 25 are essential to life
    – 14 of which are found in organisms in any more than trace (>0.01%) amounts
      » **trace elements** are essential to at least some organisms, but only in minute quantities
      » some trace elements are required by all organisms and others are required by only certain organisms

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✓of the 25 elements essential to life

– 11 are found in organisms in > than trace
  - 4 make up ~96% of the human body
    – **Carbon (C)**
    – **Hydrogen (H)**

- Oxygen (O)
- Nitrogen (N)

- 7 make up remaining ~4% of human body
  - **calcium (Ca), phosphorus (P), potassium (K), sulfur (S), sodium (Na), chlorine (Cl), magnesium (Mg)**

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✓ Each element has a symbol
- the first letter or two of its English, Latin or German name
  - first letter is capitalized
  - examples
    - gold (Au) - from Latin word *aurum*
    - oxygen (O) - from English word *oxygen*

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✓ **element** = substance that can't be broken down into any other substance by ordinary chemical means; atoms with the same atomic number and same chemical properties

✓ **molecule** = a group of atoms held together by energy in a stable association
- ex., two atoms of the element oxygen combine chemically to form a molecule of oxygen (O_2)

✓ **compound** = a molecule containing atoms of 2 or more elements combined in a fixed ratio
- ex., water is a chemical compound consisting of molecules produced when 2 atoms of hydrogen combine with one atom of oxygen (H_2O)

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✓ **Compounds**
- are much more common than pure elements
  - few elements exist in a pure state in nature
- many consist of only two elements
  - ex., table salt (NaCl)
- most compounds in living organisms contain at least 3 or 4 different elements
  - mainly C, H, O, N

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✓ **Compounds** are described using a combination of symbols and numerals
- **chemical formula (**or **molecular formula)**
- **structural formula**

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✓ **chemical formula** or **molecular formula** = a shorthand method for describing the chemical composition of molecules and compounds
- consists of chemical symbols which indicate the types of atoms present and subscript numbers which indicate the ratios among the atoms
- ex., chemical formula for water is H_2O
  - note, when a single atom of one type is present it is not necessary to write 1

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✓**structural formula** = another type of formula which shows not only the types and numbers of atoms in a compound but also their arrangement in a molecule

– ex., structural formula for water is H-O-H

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✓Each element consists of one kind of **atom**, which is different from the atoms of other elements

– the name **"atom**" comes from a Greek word meaning "indivisible"

- an **atom** is the smallest unit of matter that still retains the properties of an element

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✓**atoms** are composed of many types of subatomic particles

– **protons**

– **neutrons**

– **electrons**

– some others, discussed primarily by physicists

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✓**protons (p)**

– type of charge = single positive

– where found = nucleus of the atom

– relative mass = 1.009 daltons, approximately 1 dalton

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✓**neutrons (n)**

– type of charge = neutral, no charge

– where found = nucleus of the atom

– relative mass = 1.009 daltons, approximately 1 dalton

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✓**electrons (e)**

– type of charge = single, negative

– relative mass = 1/1840 dalton, contribution to overall mass of atom considered <u>negligible</u>

– where found = <u>orbiting the nucleus</u>

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✓**Electron orbitals** can be

– various shapes, but usually illustrated as concentric circles for purposes of simplicity

- electrons orbit the nucleus at nearly the speed of light
  - its not possible to precisely locate the position of any individual electron at any given time
  - a particular electron can be anywhere at a given point in time, from close to the

nucleus to infinitely far away from it

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✓**Electron orbitals**

– arrangement of electrons in their orbits is the key to the chemical behavior of an atom
– will return to this point shortly

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✓All atoms of a particular element have the same unique number of protons

– this is the element's **atomic number**
  • **atomic number** = the number of protons in the atom's nucleus (also the number of electrons if an atom has a neutral charge)
    – top number in box for element in periodic table

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✓An atom's **mass number** or **atomic mass** is the sum of the number of its protons and neutrons

– **atomic mass** = equal to the sum of the masses of an atom's protons & neutrons
  • measured in **daltons** (also in AMU = atomic mass units)
    – 1 gram = 602 million million billion daltons
  • also referred to as atomic weight
  • bottom number in box for element in periodic table

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✓Some elements have variant forms called **isotopes**

– atoms of the same element that vary in neutron number and atomic mass
  • have same numbers of protons and electrons but different numbers of neutrons
– isotopes of carbon
  • carbon ^{12}C – nucleus consists of 6 protons, makes up ~99% of naturally occurring C
  • carbon ^{13}C - nucleus has 7 neutrons, makes up ~1% of naturally occurring C
  • carbon ^{14}C- nucleus has 8 neutrons, occurs only in minute quantities

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✓**Isotopes** can be

– **stable**
  • nuclei remain permanently intact
    – such as the ^{12}C and ^{13}C isotopes of carbon
– **unstable** (or **radioactive**)
  • nuclei decays spontaneously, giving off particles and energy
    – such as the ^{14}C isotope of carbon

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✓**Isotopes** can be

– **unstable** (or **radioactive**)
  • in unstable isotopes the nucleus tends to break up into elements with lower atomic

numbers

- this breakup emits a significant amount of energy and is called **radioactive decay**
  - isotopes that decay in this fashion are called **radioactive isotopes**

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✓**Radioactive isotopes** can be
- harmful to life
  - by damaging molecules in cells, especially in DNA
    - the particles and energy thrown off by radioactive atoms can
      - » break apart the atoms of molecules
      - » cause abnormal connections between atoms to form

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✓**Radioactive isotopes** can have
- beneficial uses
  - living cells can't distinguish between radioactive and nonradioactive isotopes
    - serve as biological spies, monitoring the fate of atoms in living systems
      - » medical uses such as x-ray, radiation therapy, PET imaging
    - carbon dating = determining the ratios of the different isotopes of carbon, use known half-life, calculate age

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Electrons

✓Electrons orbit the nucleus of the atom

✓The arrangement of electrons in their orbits, or energy shells, is the key to the chemical behavior of an atom

✓electrons vary in the amount of energy they possess
- the farther the electron is from the nucleus, the greater its energy

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✓electrons are quite far from the nucleus
- if nucleus was size of an apple, orbit of nearest electron would be 1600m away
- as a result
  - nuclei of 2 atoms never get close enough to interact with each other (in nature)
  - electrons, not protons or neutrons, determine chemical behavior
  - isotopes of an element (which have the same # & arrangement of electrons, but different # of neutrons) behave the same chemically

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✓electrons in an atom occur only at certain energy levels, called **electron shells** (or **electron energy levels**)

✓depending on their atomic number, atoms may have 1, 2 or more electron shells
- electrons in outermost shell = highest energy
- electrons in innermost shell = lowest energy

– each shell can accommodate up to a specific number of electrons

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✓For the purpose of this class, we'll consider the first four **electron energy shells**, which covers most biologically significant elements

– the first, innermost energy shell can accommodate only 2 electrons

- in atoms with more than 2 electrons, the remainder of the electrons are found in shells farther from the nucleus

– the second, third and fourth energy levels can each accommodate 8 electrons

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✓not all atoms will have the second and third and fourth electron shells

– the atom will have just enough electron shells to accommodate its number of electrons

- atom with 6 electrons (C) has 2 shells
  – 2 electrons in innermost shell
  – 4 electrons in outermost shell
- atom with 11 electrons (Na) has 3
  – 2 electrons in innermost shell
  – 8 electrons in second shell
  – 1 electron in outermost shell

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✓Energy is required to keep electrons in their orbits

– electrons have potential energy of position

- more potential energy in outermost shells than in innermost shells
  – to oppose the attraction of the nucleus and move the electron to a more distant (higher energy level) orbital requires an input of energy and results in an electron with greater potential energy
  – moving an electron to an orbital closer to the nucleus (lower energy level) results in a release of energy (usually as heat) and the electron has less potential energy

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✓it is the number of electrons in the outermost shell that determines the chemical properties of element

– atoms whose outer shells are not full tend to interact with other atoms

- to participate in chemical reactions

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✓for example

– hydrogen (H) is highly reactive

- has one shell with only one of two possible electrons
  – it wants to "react" with another atom so it can fill its outermost shell

– helium (He) is highly unreactive (inert)

- has one shell with two of two possible electrons
  – it doesn't need to react with another atom in order to fill its outermost shell

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✓How does a chemical reaction enable an atom to fill its outer electron shell?
- when 2 atoms w/incomplete outer shells react
  - each atom either gives up or acquires electrons so that both partners end up with completed outer shells
    - by either transferring or sharing outer electrons
      - » which usually results in the atoms staying close together, held by attractions called **chemical bonds**

42 ⊟ Chemical Bonds

✓strong chemical bonds
- ionic bonds
- covalent bonds

✓weak chemical bonds
- hydrogen bonds

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✓Since electrons are negatively charged particles
- electron transfer between 2 atoms moves 1 unit of negative charge from one atom to other
  - original atom now has one less negative charge than before, resulting in charge of +1
    - now has one more positively charged proton than negatively charged electrons
  - recipient atom now has one more negative charge than before, resulting in charge of -1
    - now has 1 more negatively charged electron than positively charged protons

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✓**ions** = atoms in which the number of electrons does not equal the number of protons, and they therefore carry a net electrical charge
- types of ions
  - cations
  - anions

45 ⊟

✓**cation** = an atom with a net positive charge (+) because it has more protons than electrons
- sometimes other forces (ex. a nearby atom looking to fill its outer shell) overcome the attraction of the electrons to the nucleus and an electron is lost
- for example, a sodium (Na) atom that has lost one electron becomes a sodium ion (Na^+) with a charge of +1

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✓**anion** = an atom with a net negative charge (-) because it has fewer protons than electrons
- sometimes an atom can gain an extra electron

– for example, a chlorine (Cl) atom that has gained one electron becomes a chlorine (Cl^-) ion with a charge of -1

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✓two ions with opposite charge attract each other

– when this attraction holds the two ions together, the attraction is called an **ionic bond**

- the resulting compound is electrically neutral

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✓**covalent bond** = occurs when two atoms share one or more pairs of outer shell electrons

– results in each of the two atoms having a filled outer electron shell

- produces a very stable bond

– two or more atoms held together by covalent bonds form a molecule

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✓why is a covalent bond so stable?

– for ex., a covalent bond connects the two hydrogen (H) atoms in the molecule H_2, a common gas in the atmosphere

- has no net electrical charge (now has two protons and two electrons)
- outer shell is filled by two electrons
- has no free electrons (bonds between the two atoms also pair the two free electrons)

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✓More than one covalent bond can form between two atoms

– single covalent bond

– double covalent bond

– triple covalent bond

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✓**single covalent bond**

– one pair of electrons is shared by two atoms

– represented by a single line between the two letters representing the atoms in the structural formula, H-H

– is the least strong of the covalent bonds

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✓**double covalent bond**

– two pairs of electrons are shared by two atoms

– represented by a double line between the two letters representing the atoms in the structural formula, O=O

– is stronger than the single covalent bond because more chemical energy is required to break a double bond than a single bond

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- **triple covalent bond**
  - three pairs of electrons are shared by two atoms
  - represented by a triple dash line between the two letters representing the atoms in the structural formula, N ≡ N
  - is the strongest of the covalent bonds because more chemical energy is required to break a triple bond than a double or single bond

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- covalent bond energy
  - forming a bond requires an input of energy
    - and that energy is then stored in that bond
  - breaking a bond results in a release of energy
    - and that released energy then becomes available to do work

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- covalent bond energy
  - forming a bond requires an input of energy
    - and that energy is then stored in that bond
  - breaking a bond results in a release of energy
    - and that released energy then becomes available to do work

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- Atoms in a covalently bonded molecule are in a constant tug-of-war for the shared electrons in the covalent bond
  - **electronegativity** is a measure of this attraction for the shared electrons in chemical bonds
    - the more electronegative an atom is, the more strongly it pulls the shared electrons towards itself

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- Beccause of this concept of electronegativity**, covalent bonds** can be divided into two categories
  - **nonpolar covalent bonds**
  - **polar covalent bonds**

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- **nonpolar covalent bond** = a covalent bond between atoms that have similar electronegativity

- as a result, electrons are shared exactly between the two atoms
- examples
  - O_2
  - H_2
  - CH_4 (methane)

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✓**polar covalent bond** = a covalent bond between atoms that differ in electronegativity

- electrons are pulled closer to the atomic nucleus with the greater electron affinity
- such a bond has two dissimilar ends, or “poles”, one with a partial positive charge and the other with a partial negative charge
- example
  - H_2O

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✓**Water** is a **polar molecule**

- characteristics of **water (H_2O)**
  - importance to life
  - chemical characteristics

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✓Importance of **water** to life

- covers ¾ of surface of earth
- is where life evolved
- essential to life on earth
- makes up 2/3 or more of mass of all organisms

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✓chemical characteristics of **water (H_2O)**

- two hydrogen atoms covalently bonded to one oxygen atom
  - resulting molecule is stable because
    - its outer electron shells are full
    - it has no net charge
    - it has no unpaired electrons

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✓**Water** is a **polar molecule**

- the O atom is more electronegative than the H atoms, so it attracts the electrons more strongly than do the H atoms

- therefore, the shared electrons in a water molecule are far more likely to be found near the O nucleus than near the H nuclei, resulting in
  - – partial – charge on the O atom
  - – partial + charge on each H atom

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✓**Water** is a **polar molecule**

– although the water molecule as a whole is neutral, the partial charges cause the molecule to have “poles”

- negative pole
  - – at the O end because of the partial – charge on the O atom
- positive poles
  - – at the H ends because of the partial + charge on each H atom

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✓**Water’s polarity** leads to unusual properties that make life, as we know it, possible

– **hydrogen bonds**

– **cohesion**

– **surface tension**

– **temperature moderation**

– **less dense as solid than liquid**

– **versatile solvent**

– **role in acid/base conditions**

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✓The **polarity** of **water molecules** makes them interact <u>with each other</u>

– the partial charges at the poles of one water molecule are attracted to the opposite partial charges at the poles of neighboring water molecules

- this attraction results in the formation of weak bonds called **hydrogen bonds**

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✓**hydrogen bonds** = result when polar molecules interact with one another

– partial – charge of one molecule is attracted to the partial + charge of another molecule

– in the case of water, the partial – charge of the O atom in one water molecule is attracted to the partial + charge of the H atoms in another water molecule

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✓**hydrogen bonds**

– each hydrogen bond individually is very weak, readily formed and broken, lasting on

average on 1/100,000,000,000 second
  - cumulative effects of large numbers of these bonds can be enormous
  - each water molecule can form hydrogen bonds with a maximum of four neighboring water molecules
  - hydrogen bonds extremely important to biological systems

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✓Like no other common substance on Earth, **water exists in nature in all three physical states** (or phases of matter)
  - solid (ice)
  - liquid (water)
  - gas (water vapor)

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✓**cohesion** is the attraction resulting from polar water molecules being attracted to other polar water molecules
  - water molecules have a strong tendency to stick together
    - **cohesion** is much stronger for water than for most other liquids
  - cohesion of water is important in the living world
    - ex., trees depend on cohesion to transport water from roots to leaves

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✓Related to cohesion is **surface tension**
  - is a measure of how difficult it is to stretch or break the surface of a liquid
  - it results because at the air-water interface, all the hydrogen bonds in water face downward, causing the molecules of the water surface to cling together
    - polar water molecules are “repelled” by the nonpolar molecules in the air

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✓**surface tension**

✓water has highest surface tension of any liquid except for liquid mercury

✓some insects can walk on water because the surface tension of the water is greater than the force that one foot brings to bear

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✓**adhesive properties** = adhesion is the attraction resulting from polar water molecules being attracted to other polar molecules (non-water)
  - water molecules have a strong tendency to stick to other polar molecules and to substances with surface electrical charges (such as glass)
    - **capillary action** = the tendency of water to rise in small tubes, as a result of cohesive and adhesive forces

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✓Because of **water’s hydrogen bonds**, it has a greater ability to **resist temperature change** than most other substances

– **heat** is the amount of energy associated with the movement of the atoms and molecules in a body of matter
– **temperature** is the intensity of heat
  • the average speed of the molecules rather than the total amount of heat in a body of matter

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✓Water resists temperature increases
– raising the temperature of a substance involves adding heat energy to make its molecules move faster
– in water, some of the H bonds must first be broken to allow the molecules to move more freely
– much of the energy added to the water is used up in breaking the H bonds, only a portion of the heat energy is available to speed the movement of the water molecules

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✓Water stores heat
– heat is absorbed as H bonds break
  • water absorbs and stores a large amount of heat while warming up only a few degrees

✓Water cools slowly
– as water cools, H bonds re-form
  • heat energy is released as H bonds form, thus slowing the cooling process

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✓Water resists temperature change
– enables organisms, which have a high water content, to maintain a relatively constant internal temperatures
  • the heat generated by chemical reactions within cells would destroy cells if not for the high specific heat of the water within the cells
– crucial in stabilizing temperatures on earth
  • by storing heat from sun during warm periods, releasing heat during cooler times

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✓Water resists tendency to evaporate or vaporize
– liquids vaporize when some of their molecules move fast enough to overcome the attractions that keep the molecules close together
  • heating a liquid increases vaporization by increasing the energy of the molecules
    – providing some of the molecules with enough energy to escape

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✓Water resists tendency to vaporize
– a large amount of energy required to change one gram of liquid water into a gas
– water's resistance to vaporization results from the hydrogen bonding of its molecules
  • the transition of water from a liquid to a gas requires the input of energy to break the many hydrogen bonds

– the source of such energy can be the surface of the substance on which the water is located

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✓the evaporation of water from surfaces causing a cooling of that surface

– this characteristic enables organisms to dispose of excess heat by **evaporative cooling**
  - the organism gives up some heat energy to break H bonds in the water molecules
    – such molecules then have enough heat energy to escape & they take that heat energy with them when they go

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✓most substances become more dense as the temperature decreases

✓water is most dense at 4°C and then begins becomes less dense as temperature decreases below that point

✓hydrogen bonds in liquid water are unstable because they constantly break and re-form

✓hydrogen bonds in ice are stable

– each molecule bonds to 4 neighbors forming a 3-D crystal

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✓liquid water expands (becomes less dense) as it freezes because

– the H bonds joining in the water molecules in the crystalline lattice keep the molecules far enough apart to give ice a density about 10% less than the density of water

– the less dense frozen water (ice) floats upon the more dense cold, unfrozen water

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✓Ice is less dense than liquid water

– frozen water floats on liquid water

– this property of water has been an extremely important factor in enabling life to appear, survive and evolve
  - if ice were more dense than water it would sink
    – and all ponds, lakes, oceans would freeze solid from the bottom to the surface making life impossible

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✓Ice is less dense than liquid water

– since ice floats on water instead of sinking
  - when a body of deep water cools, it freezes at the top, becoming covered with floating ice
    – ice insulates liquid water below it, prevents freezing solid, thus allowing a variety of animals and plants to survive below the icy surface

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✓ **solution** is a liquid, that is uniform throughout (homogeneous), consisting of a mixture of two or more substances

- **solvent** is the substance in a solution that serves as the dissolved agent
  - a substance (usually liquid) capable of dissolving one or more other substances
- **solute** is the substance which is dissolved by the solvent

✓ a solution that has water as its solvent is called an **aqueous solution**

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✓ **Water is a versatile solvent**

- dissolves an enormous variety of solutes necessary for life
  - water is the solvent in all cells
    - therefore, its the solvent of blood, tree sap, etc.
- results from the polarity of its molecules
  - solutes whose charges or polarity allow them to stick to water molecules will dissolve in water, forming an aqueous solution

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✓ **Water is a versatile solvent**

- consider how a crystal salt dissolves in water
  - the Na^+ and Cl^- ions at the surface of the salt crystal have affinities for different parts of the water molecules
    - Na^+ ions attract the - area of H_2O at O
    - Cl^- ions attract the + areas at H's
  - as a result, the water molecules surround and separate the Na^+ and Cl^- ions (hydration shell)
    - causing the salt crystal to dissolve

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✓ Most water molecules remain intact in aqueous solutions within living organisms

- but some water molecules actually break apart in a process called **dissociation** or **ionization**
  - formation of ions when covalent bonds in a water molecule break spontaneously
    - at 25°C, one out of every 550 million water molecules spontaneously undergoes this process

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✓ Two types of ions result from the dissociation of water molecules (H_2O)

- **hydrogen ions (H^+)** with + charge
- **hydroxide ions (OH^-)** with – charge

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✓ Two types of ions result from dissociation

- **hydrogen ions (H^+)** with + charge result
  - when one of the protons (from hydrogen atom nuclei) dissociate from the rest of

the molecule

- because the dissociated proton lacks the negatively charged electron it was sharing in the covalent bond with oxygen, its own positive charge is no longer counterbalanced, and it becomes a positively charged ion, H^+

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✓Two types of ions result from dissociation

– **hydroxide ions (OH^-)** with - charge results

- when one of the protons (from hydrogen atom nuclei) dissociate from the rest of the molecule
- the rest of the dissociated water molecule, which has retained the shared electron from the covalent bond, is negatively charged and forms a hydroxide ion, OH^-

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✓Hydrogen and hydroxide ions result from the spontaneous dissociation of water molecules in aqueous solutions

– the right balance of these two ions is required for the proper functioning of chemical processes within organisms

- we describe and measure this balance between these two ions in the terms of **acids**, **bases** and the **pH scale**

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✓**acid** = any substance that dissociates in water to increase the concentration of H^+ ions

✓**base** (or **alkali**)= is any substance that combines with H^+ ions when dissolved in water

✓**neutral** = a substance in which the concentrations of H^+ ions and OH^- ions are equal

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✓the **pH scale** is used to measure the acidity or alkalinity of a solution

– pH stands for potential hydrogen

– its the negative logarithm of the hydrogen ion ($[H^+]$) concentration in the solution (the negative logarithm of 10^{-7} equals 7, and therefore the pH of pure water is 7)

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✓An **acid** is any substance that dissociates in water to increase the concentration of H^+ ions

– the stronger an acid is, the more H^+ ions it produces

– acidic solutions have pH values below 7

– strongly hydrochloric acid (HCl), abundant in your stomach, ionizes completely in water to H^+ and Cl^- ions, has a pH of 1

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✓**base** = any substance that combines with H^+ ions when dissolved in water

– by combining with H^+ ions, a base lowers the H^+ ion concentration in the solution

– basic, or alkaline, solutions have pH values above 7

– strong bases, such as sodium hydroxide (NaOH), have pH values of 12 or more

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✓**neutral** = a substance in which the concentrations of H^+ ions and OH^- ions are equal

– neutral solutions have a pH value of 7
  - at 25°C, a liter of pure water contains 1/10,000,000 (or 10^{-7}) mole of H^+ ions
  - the negative logarithm of 10^{-7} equals 7, and therefore the pH of pure water is 7

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✓the pH inside almost all cells, and in the fluid surrounding cells, is fairly close to 7

– therefore, even a slight change in pH can be harmful
  - biological fluids contain **buffers** that resist changes in pH

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✓**buffer**

– a substance that resists changes in pH by
  - accepting H^+ ions when they're in excess
  - donating H^+ ions when they're depleted

– a substance that acts as a reservoir for hydrogen (H^+) ions
  - taking H^+ ions from the solution when their concentration increases
  - donating H^+ ions to the solution when their concentration falls

– however, buffers are not foolproof

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✓it is important that a cell maintain a constant pH level

– the pH of an organism is kept at a relatively constant pH by **buffers**
  - within organisms most buffers act as pairs of substances, one an acid and the other a base
  - ex., the key buffer in human blood is an acid-base pair consisting of carbonic acid (acid) and bicarbonate (base)

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✓because changes in pH can harm living organisms, changes in the pH of the environment can have drastic effects

– **acid precipitation** (rain, fog, snow) can cause changes in the pH of the environment
  - these pH changes can kill fish in lake, trees in forests, affect human health, erode buildings

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✓**acid precipitation** (rain, fog, snow)

– precipitation with a pH below 5.6
  - rain with pH of 2-3, more acidic than vinegar, recorded in eastern US

- fog with pH1.7, nearly acidic as human stomach digestive juices, recorded downwind from LA

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✓**acid precipitation** (rain, fog, snow)

- results mainly from the presence in the air of sulfur oxides and nitrogen oxides
  - which result mostly from the burning of fossil fuels in factories and automobiles
    - coal, oil and gas are fossil fuels
- is a complex environmental problem with no easy solution

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✓**chemical reactions** lead to chemical changes in matter

- are the essence of chemistry and life

✓all chemical reactions involve

- the shifting of atoms from one molecule or ionic compound to another
  - via the formation and breaking of chemical bonds
- without any change in the number or identity of the atoms

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✓all chemical reactions involve

- **reactants**
  - the original molecules before a chemical reaction starts
- **products**
  - the molecules resulting from the chemical reaction

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✓chemical reactions can be described by **chemical equations**

- **reactants** are generally written on the left side of the equation
- **products** are generally written on the right side of the equation
- an **arrow** (instead of =)between the “reactants” side and the “products” side
  - means ***“yields”***
  - indicates the direction in which the reaction tends to proceed

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✓**chemical equations**

- example: $2H_2$ + O_2 $\rightarrow$ $2H_2O$
- reactants products
- the same numbers of H and O atoms appear on both the left and right hand side of the arrow but are grouped differently
  - (H-H) + (H-H) + (O-O) = (H-O-H) + (H-O-H)
  - 4 H, 2 O = 4 H, 2 O
  - 2 molecules of H plus 1 molecule of O yields 2 molecules of water
- (note: organisms can’t make water from H & O)

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✓chemical equations

- can proceed in two directions
  - **forward** = to the right $\rightarrow$
  - **reverse** = to the left $\leftarrow$
- when the rates of the forward and reverse reactions are equal the reaction has reached **equilibrium**

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✓chemical reactions

- organisms carry out a great number of chemical reactions, most involving carbon, that rearrange matter in significant ways
- examples
  - photosynthesis
    - $6CO_2 + 6H_2O + \text{energy} \rightarrow 6O_2 + C_6H_{12}O_6$
    - 6C, 12H, 18O $\rightarrow$ 6C, 12H, 18O
  - production of vitamin A in human cells
    - $C_{40}H_{56} + O_2 + 4H \rightarrow 2C_{20}H_{30}O$
    - beta-carotene vitamin A
    - 40C, 2O, 60H $\rightarrow$ 40C, 2O, 60H

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