Chapter 2 – "Basic Chemistry"

30 January 2012

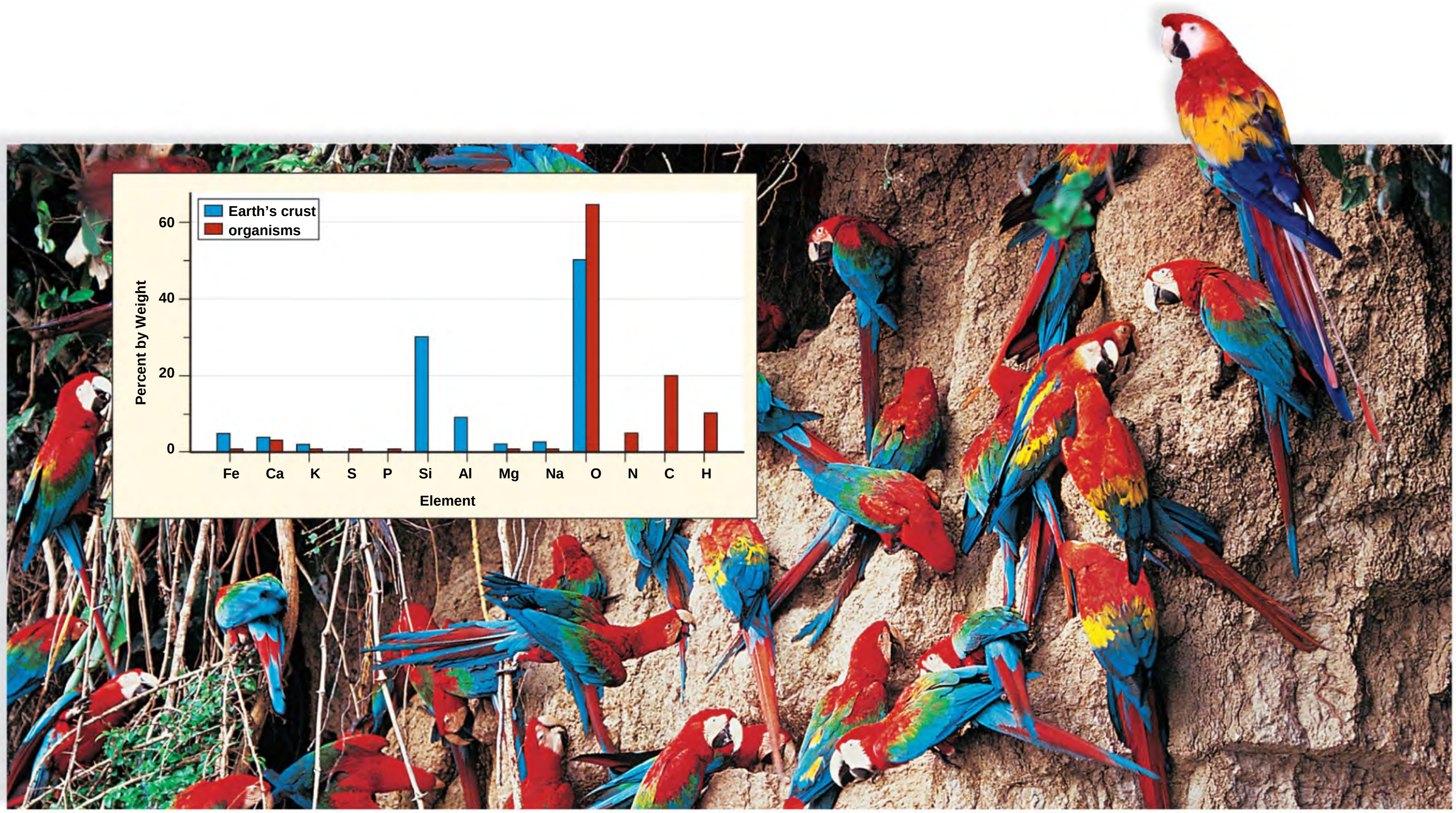
Outline

- Chemical Elements
  - Atoms
    - Atomic Mass and Atomic Number
  - Periodic Table
  - Isotopes
  - Electrons and Energy
  - Molecules and Compounds
- Chemical Bonding
  - Ionic and Covalent
  - Hydrogen
- Properties of Water
- Acids and Bases

Chemical Elements

- Matter:
  - Matter is defined as anything that has mass and occupies space
  - Matter exists in three states: **solid, liquid, and gas**
  - All matter (both living and non-living) is composed of 92 naturally-occurring **elements**
  - 98% of body weight of organisms are primarily composed of six elements (carbon, hydrogen, nitrogen, oxygen, phosphorus, and sulfur—acronym CHNOPS) make up 98% of the body weight of organisms.

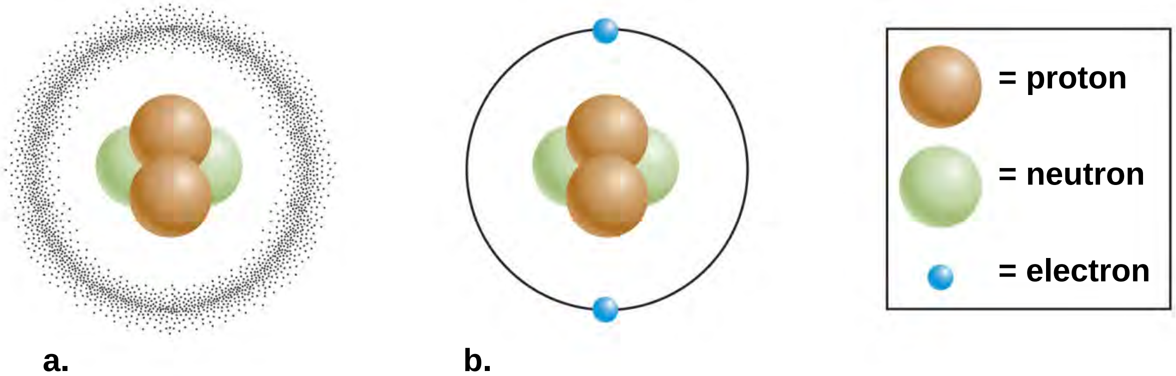
Composition of Earth's Crust versus Organisms



Atomic Structure

- Atom is the smallest unit of an element
- Atoms composed of subatomic particles:
  - Protons - positive charge; weight of approximately 1 Dalton, found in the nucleus
  - Neutrons - no charge; weight similar to protons, found in the nucleus
  - Electrons - negative charge; weigh 1/1836th Dalton; found in electron shell
- Atoms contain specific numbers of **protons, neutrons,** and **electrons.**

Subatomic Particles



Subatomic Particles			
Particle	Electric Charge	Atomic Mass Unit (AMU)	Location
Proton	+1	1	Nucleus
Neutron	0	1	Nucleus
Electron	–1	0	Electron shell

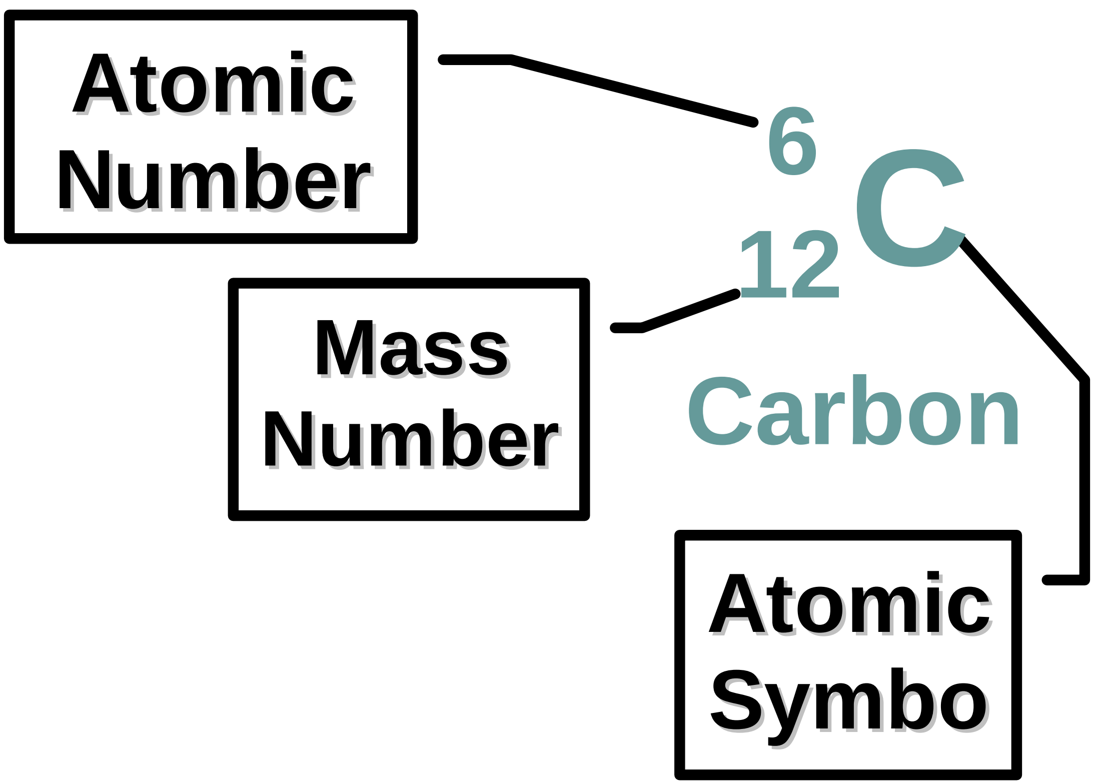
c.

Atomic Symbols

- Each element is represented by one or two letters to give them a unique atomic symbol
  - H = hydrogen, Na = Sodium, C = Carbon
- Each atom has its own specific mass (atomic mass)
- Atomic mass of an atom depends on the presence of subatomic particles
  - Atomic number = proton number;
  - Atomic mass or mass number = protons and neutrons

Atomic Symbol

The atomic number is above the atomic symbol and the atomic mass is below the atomic symbol



Periodic Table

- Elements grouped in periodic table based on characteristics
  - Vertical columns = groups; chemically similar
  - Horizontal rows = periods; larger and larger
- Atomic mass increases as you move down a group or across a period.

Periodic Table

Key: 1 — atomic number; H — atomic symbol; 4.003 — atomic mass

Periods / Groups	I	II	III	IV	V	VI	VII	VIII
	1 H 1.008							2 He 4.003
	3 Li 6.941	4 Be 9.012	5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
	1 Na 22.99	12 Mg 24.31	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
	19 K 39.10	20 Ca 40.08	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.60

Groups

Periodic Table

atomic number — atomic symbol — atomic mass

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Groups

Isotopes

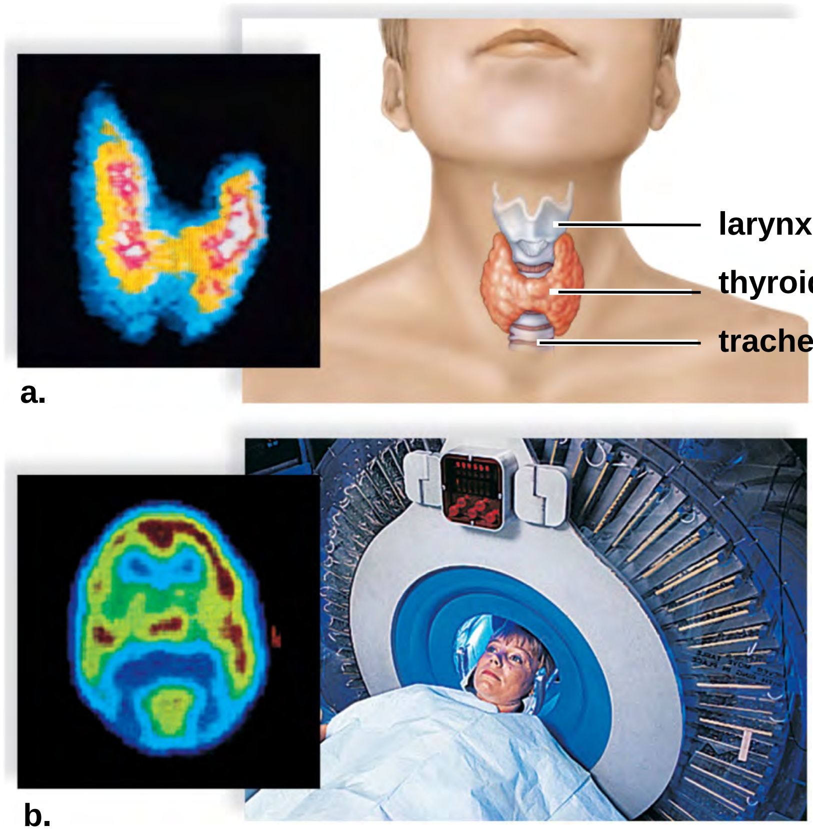
- Isotopes:
  - Atoms of the same element with a differing numbers of neutrons (and therefore have different atomic masses). e.g. see carbon below

  Some isotopes spontaneously decay
    - Radioactive
    - Give off energy in the form of rays and subatomic particles
    - Can be used as tracers
    - Mutagenic – Can cause cancer

$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
Carbon 12	**Carbon 13**	**Carbon 14**

Some Medical Uses for Low Level Radiation

- MRI, CT, Nuclear Medicine



a: © Biomed Commun./Custom Medical Stock Photo; b(Right): © Hank Morgan/Rainbow; b(Left): © Mazzlota et al./Photo Researchers, Inc

Some Medical Uses for High Level Radiation

- Radiation can kill cancer cells
- Radiation can preserve food longer

a. **b.**

a: (Peaches): © Tony Freeman/PhotoEdit; b: © Geoff Tompkinson/SPL/Photo Researchers, Inc.

Electrons and Energy

- Atoms normally have as many electrons as protons
- Opposite charges balance leaving atom neutral
- Electrons are attracted to the positive nucleus
  - Revolve around nucleus in orbitals
  - Can be pushed into higher orbitals with energy
  - Release that energy when they fall back to lower orbital
  - Different energy levels referred to as electron shells

Animation


Electron Configurations
28
Ni
: [Ar]
4s2
3d8
Main menu
The remaining 3 electrons are paired up in three of the d orbitals with opposite spin. This is the ground-state orbital diagram of nickel.
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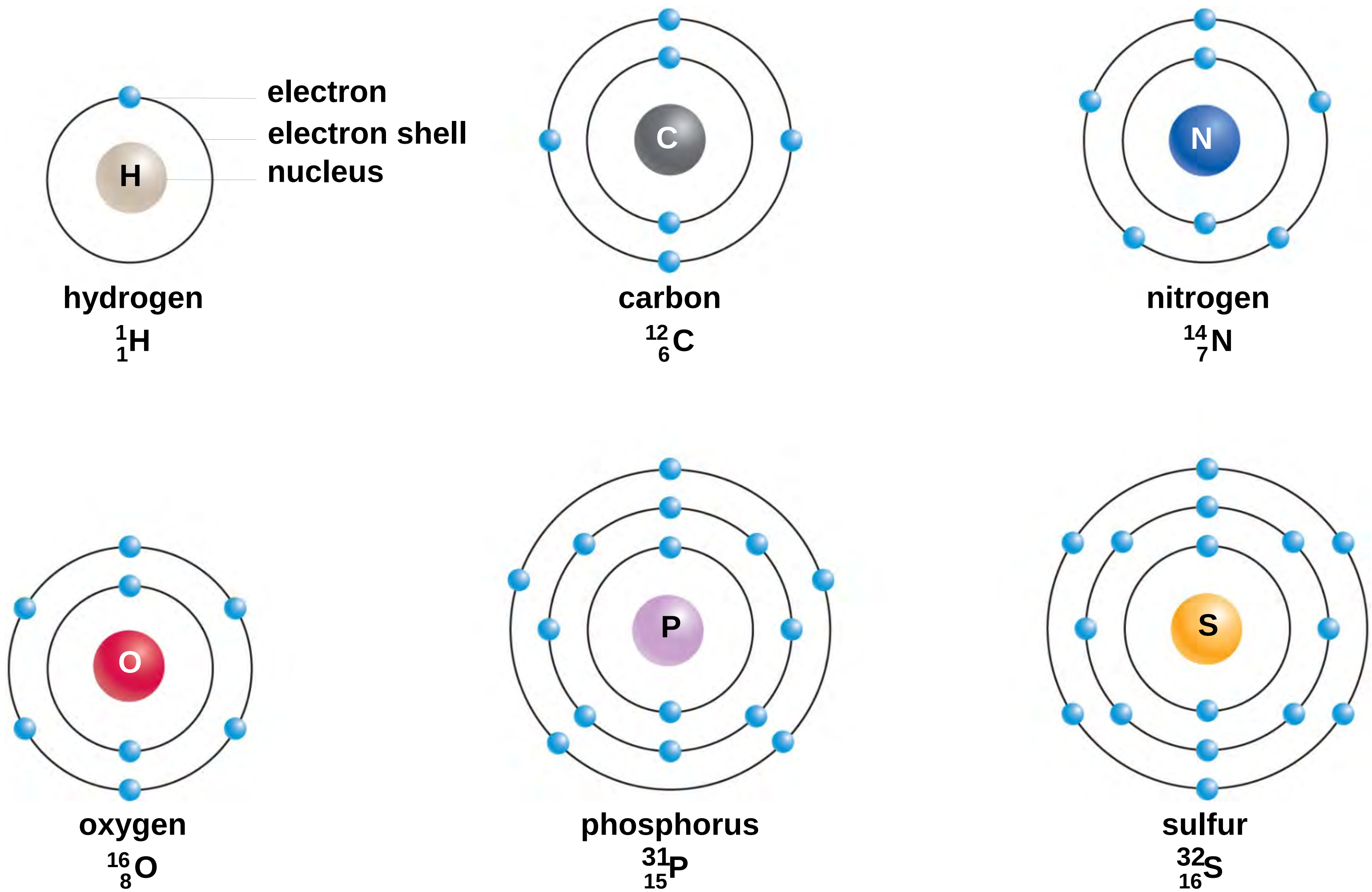
The Octet Rule for Distribution of Electrons

- Bohr models show electron shells as concentric circles around nucleus
  - Each shell has two or more electron orbitals
    - Innermost shell has two orbitals
    - Others have 8 or multiples thereof
- The outermost electron shell determines the reactivity of the element
  - If 3 or less – Tendency to donate electrons
  - If 5 or more – Tendency to receive electrons

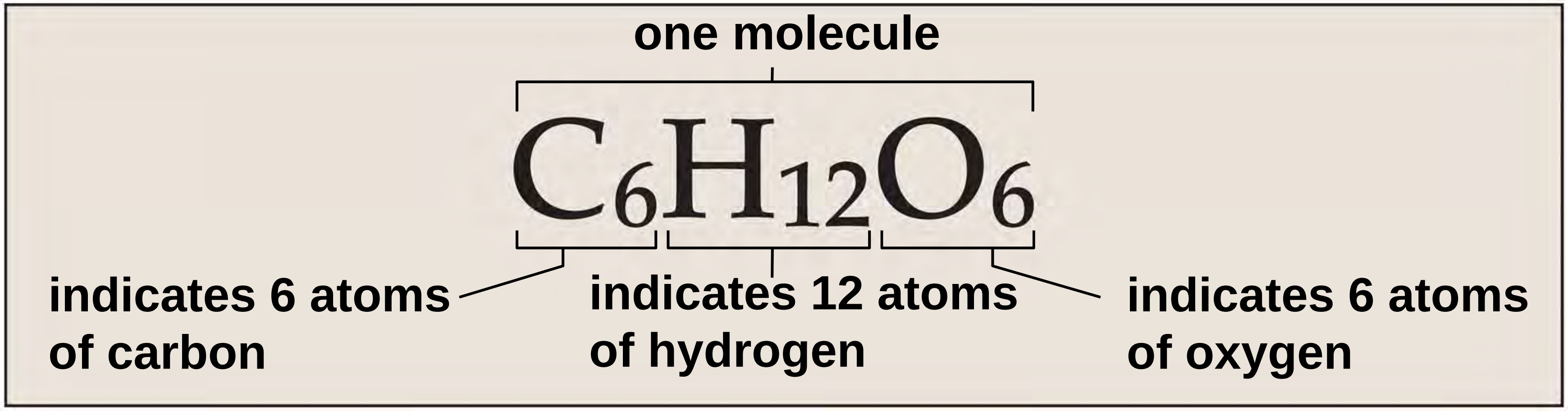
Bohr Models of Atoms



Compounds and Molecules

- Compound - when atoms of two or more different elements bond together
  - CO_2, H_2O, $C_6H_{12}O_6$, etc.
  - Characteristics dramatically different from constituent elements
- Molecule and compound is used interchangeably
  - In Biology molecule is used e.g. molecule of water (H_2O) molecule of glucose ($C_6H_{12}O_6$)
- Bonds that exist between atoms in molecules contain energy

Compounds and Molecules



Chemical Bonding

- Bonds between atoms are caused by electrons in outermost shells
- The process of bond formation is called a reaction
- The intensity of simple reactions can be predicted by the periodic table
  - If two elements are horizontally close in the table, they usually react mildly
  - If they are horizontally far apart, they usually react vigorously

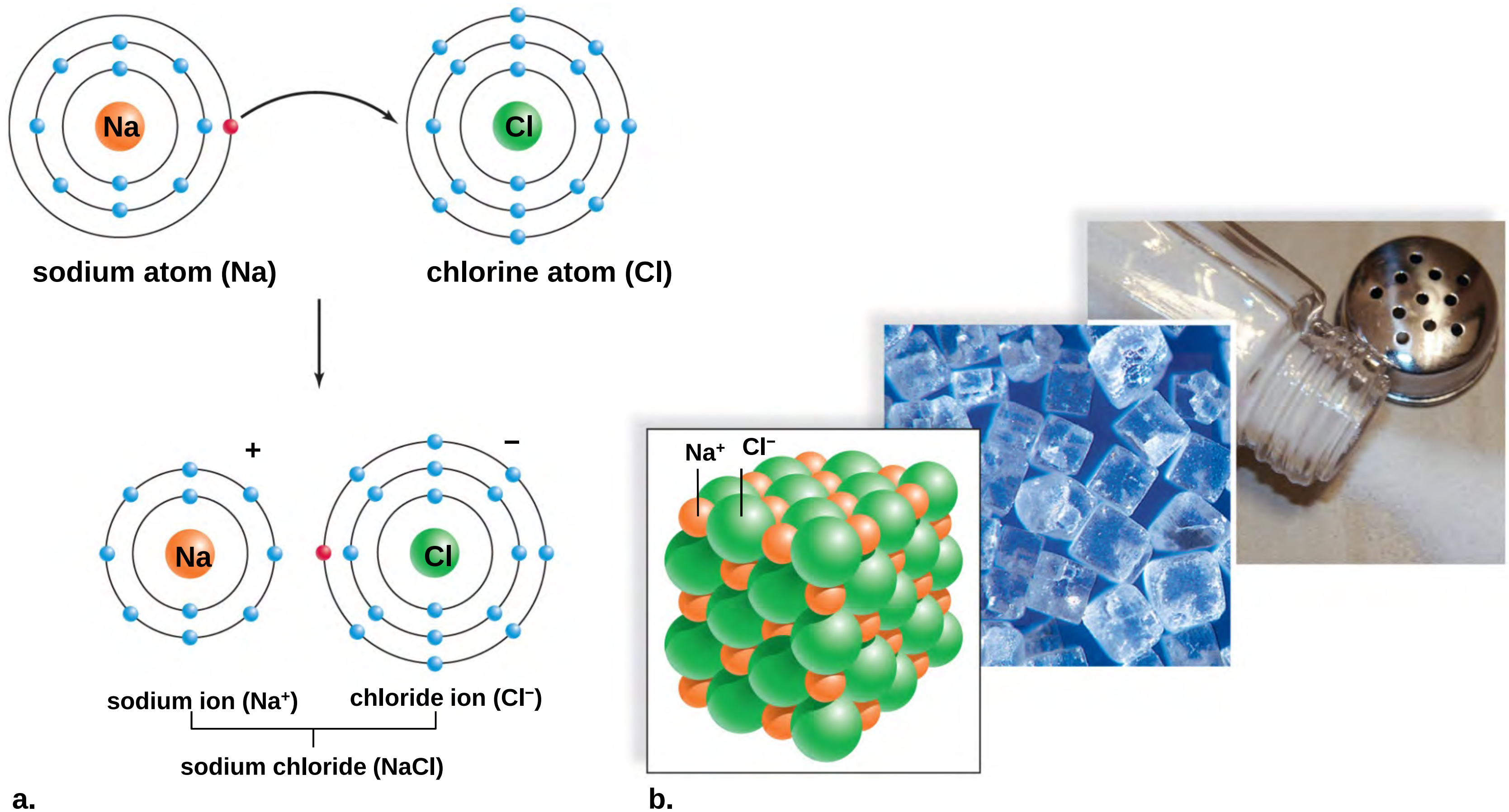
Types of Bonds: Ionic Bonding

- **Ionic bond** - forms when electrons are transferred from one atom to another atom.
- Octet rule – atoms lose or gain electrons to fill their outer shells and become more stable
  - Atoms "want" 8 electrons in outer shell
    - If have < 4, desire to donate electrons
    - If have > 4, desire to receive electrons
- Consider two elements from opposite ends of periodic table
    - Element from right side:
      - Has 7 electrons in outer shell
      - "Desperately wants" one more (7+1=8)
    - Element from left side:
      - Has only 1 electron in outer shell
      - "Desperately wants" to donate it (1-1=0=8)

Types of Bonds: Ionic Bond Example

- Sodium (Na):
  - Has only 1 electron in its outermost shell
- Chlorine (Cl):
  - Has 7 electrons in its outermost shell
- In a reaction between Na and Cl
  - Na loses an electron and becomes a positive ion (Na^+)
  - Cl gains an electron and becomes a negative ion (Cl^-)
  - Attraction of oppositely charged ions holds the two atoms together in an **ionic bond**

Formation of Sodium Chloride



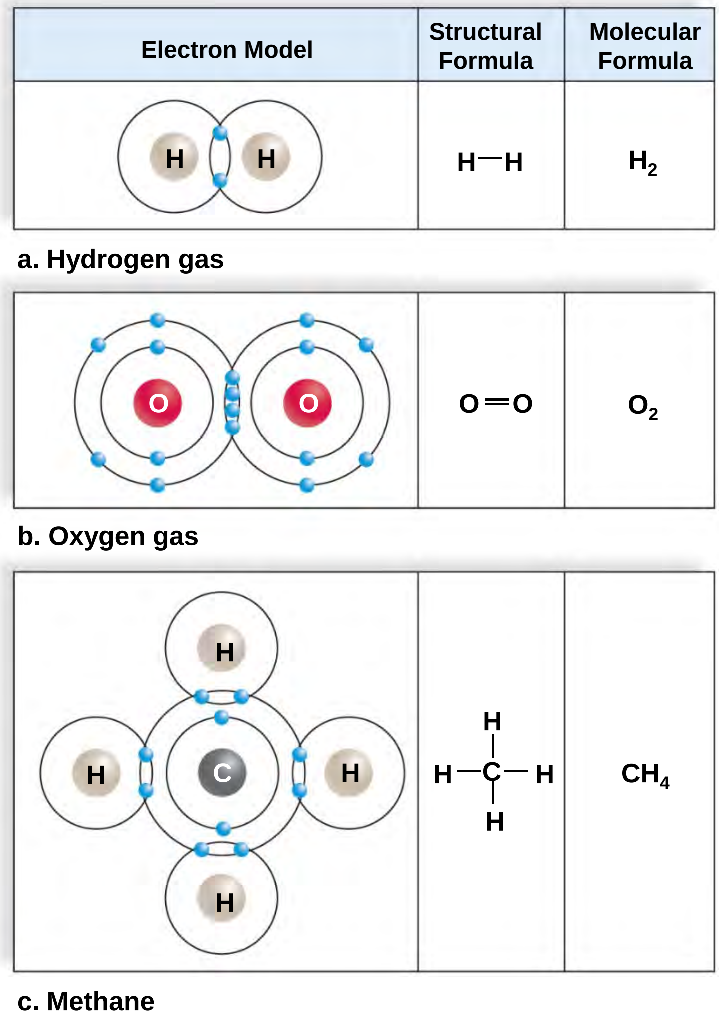
(Crystals): © Charles M. Falco/Photo Researchers, Inc.; (Salt shaker): © Erica S. Leeds

Types of Bonds: Covalent Bonds

- Covalent bonds result when two atoms share electrons so each atom has an octet of electrons in the outer shell (in the case of hydrogen, 2 electrons).
- When atoms are horizontally closer together in the periodic table
  - The electrons are not permanently transferred from one atom to the other like in NaCl
  - A pair of electrons from the outer shell will “time share” with one atom and then the other
  - This also causes the atoms to remain together
  - Known as covalent bonding
- Double covalent bond - when two pairs of electrons are shared between atoms

Covalently Bonded Molecules

- The **structural formula** of a molecule indicates a shared pair of electrons by a line between the two atoms e.g. single covalent bond (H–H), double covalent bond (O=O), and triple covalent bond (N ≡ N).



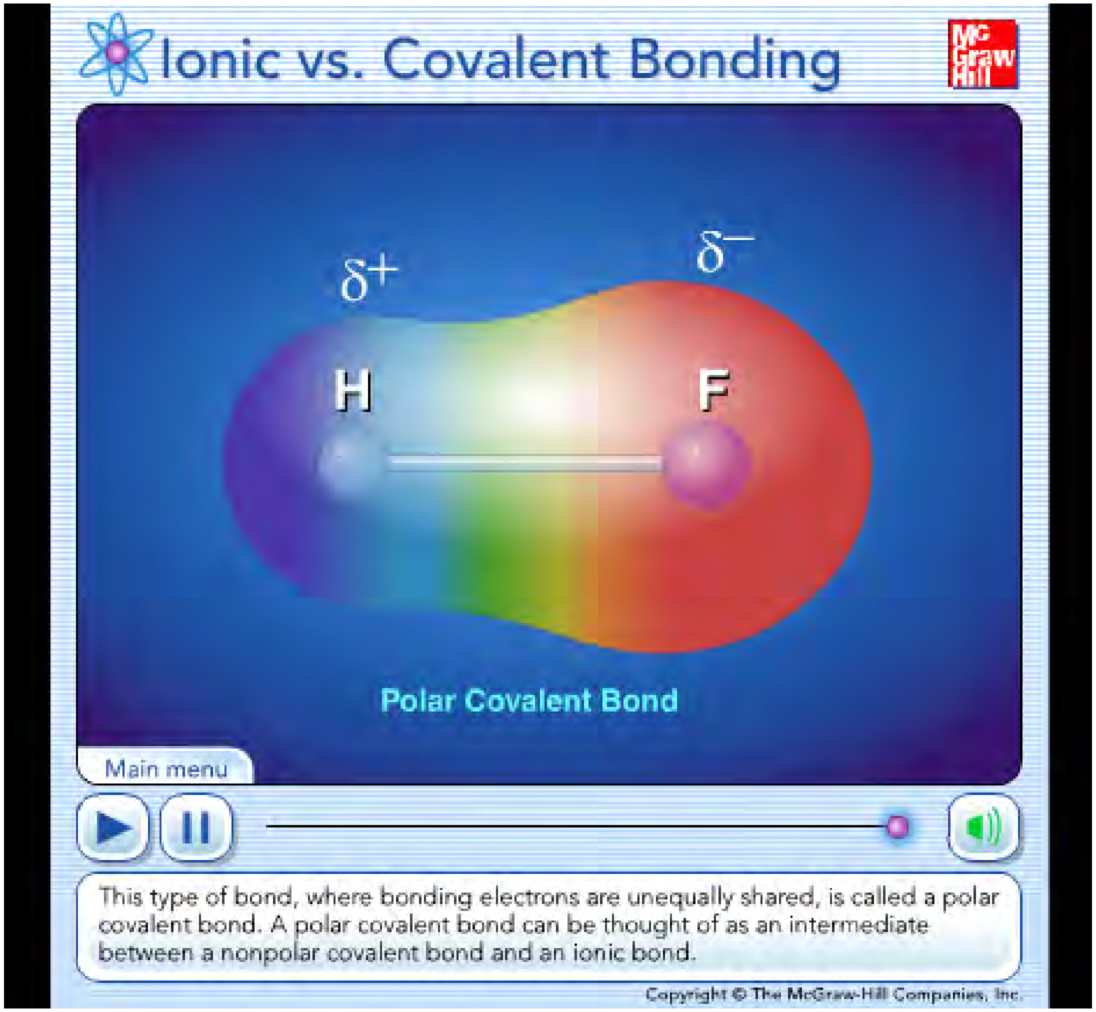
Nonpolar Covalent Bonds

- In **nonpolar covalent bonds**, sharing of electrons is equal, i.e. the electrons are not attracted to either atom to a greater degree
  - One atom “wants” (with a specific intensity) to donate electron(s)
  - The other atom “wants,” (with the same intensity) to receive electron(s)
  - The bond electrons will spend about equal time with both atoms

Polar Covalent Bonds

- With **polar covalent bonds**, the sharing of electrons is unequal i.e. atoms will have unequal affinity for electrons
  - One atom “wants” to donate or receive electron(s) with a specific intensity
  - The other atom “wants” to donate or receive electron(s) with a different intensity
    - In H2O - sharing of electrons by oxygen and hydrogen is not equal; the oxygen atom with more protons attracts the electrons closer therefore assumes a partial negative charge i.e. the atom that gets the most time with the electrons will be slightly negative

Animation


Ionic vs. Covalent Bonding
Mc Graw Hill
δ+
δ−
H
F
Polar Covalent Bond
Main menu
This type of bond, where bonding electrons are unequally shared, is called a polar covalent bond. A polar covalent bond can be thought of as an intermediate between a nonpolar covalent bond and an ionic bond.
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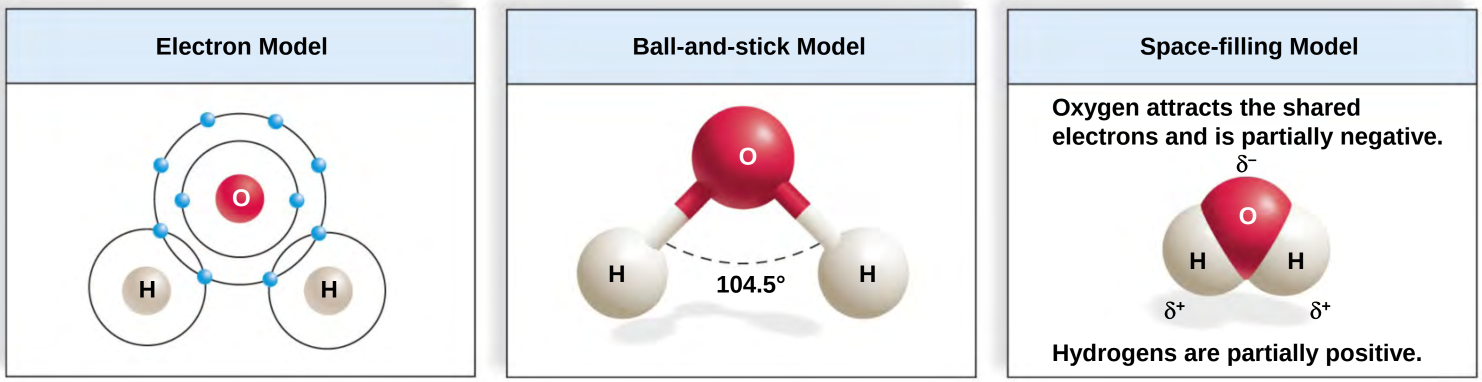
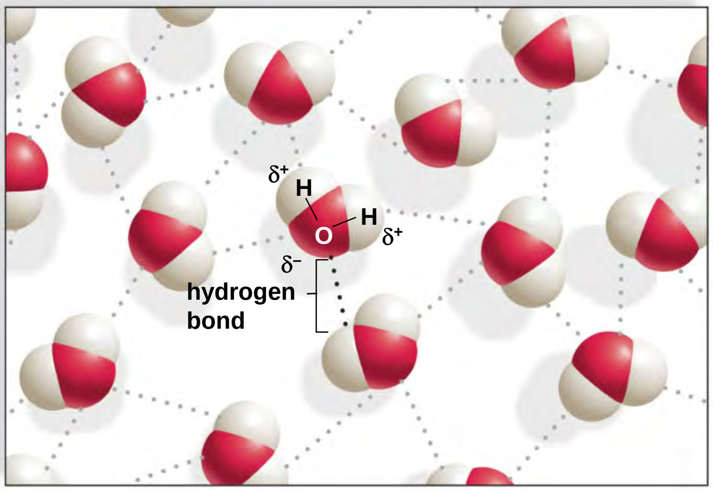
Types of Bonds: Hydrogen Bonds

- Water (H_2O or H–O–H) is a polar molecule
  - Electrons spend more time with O than H's
  - H's become slightly +, O slightly –
- When polar molecules are dissolved in water
  - The H's of water molecules are attracted to the negative parts of the solute molecules and form a hydrogen bond
  - This bond is a weak attractive force between the slightly positive charge of the hydrogen atom of one molecule and slightly negative charge of another atom
  - Easily broken, but many together can be quite strong
- Help to maintain the proper structure and function of complex molecules such as proteins and DNA.

Water Molecule



a. Water (H_2O)



b. Hydrogen bonding between water molecules

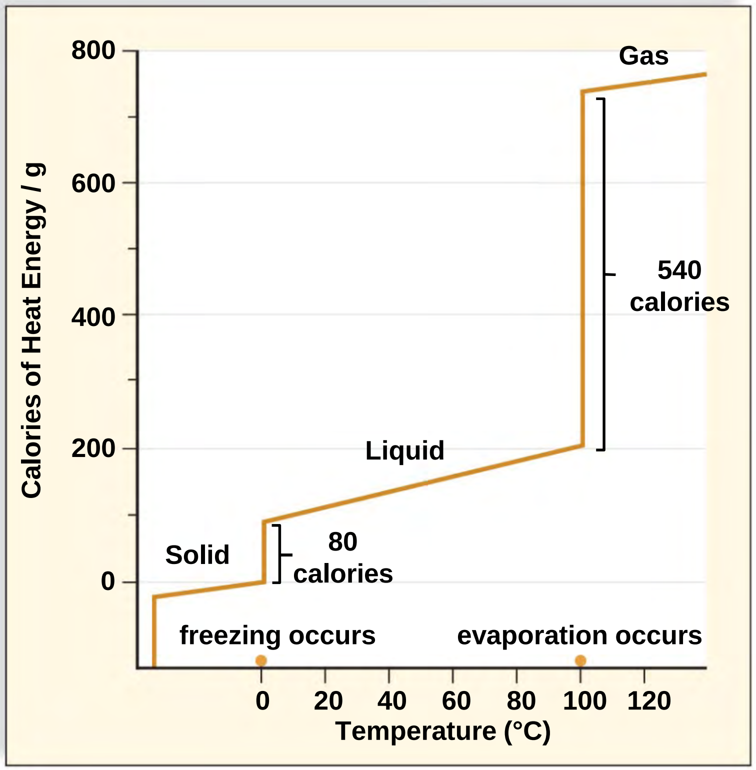
The Chemistry of Water: Heat Capacity

- Water has a high heat capacity
  - Temperature = rate of vibration of molecules
  - Apply heat to liquid
    - Molecules bounce faster
    - Increases temperature
  - But, when heat applied to water
    - Hydrogen bonds restrain bouncing
    - Temperature rises more slowly per unit heat
    - Water at a given temp. has more heat than most liquids
- Thermal inertia – resistance to temperature change
  - More heat required to raise water one degree than most other liquids (1 calorie per gram)
  - Also, more heat is extracted/released when lowering water one degree than most other liquids

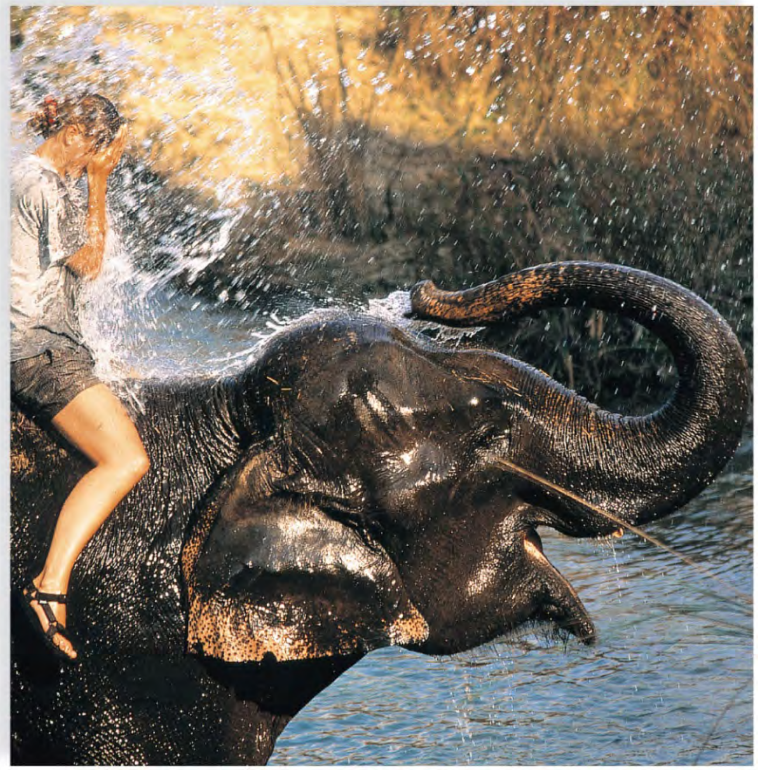
Properties of Water: Heat of Vaporization

- High heat of vaporization
  - To raise water from 98 to 99 ºC; ~1 calorie
  - To raise water from 99 to 100 ºC; ~1 calorie
  - However, large numbers of hydrogen bonds must be broken to evaporate water
  - To raise water from 100 to 101 ºC; ~540 calories!
- This is why sweating (and panting) cools
  - Evaporative cooling is best when humidity is low because evaporation occurs rapidly
  - Evaporative cooling works poorest when humidity is high because evaporation occurs slowly

Evaporative Cooling of Animals



a. Calories lost when 1 g of liquid water freezes and calories required when 1 g of liquid water evaporates.

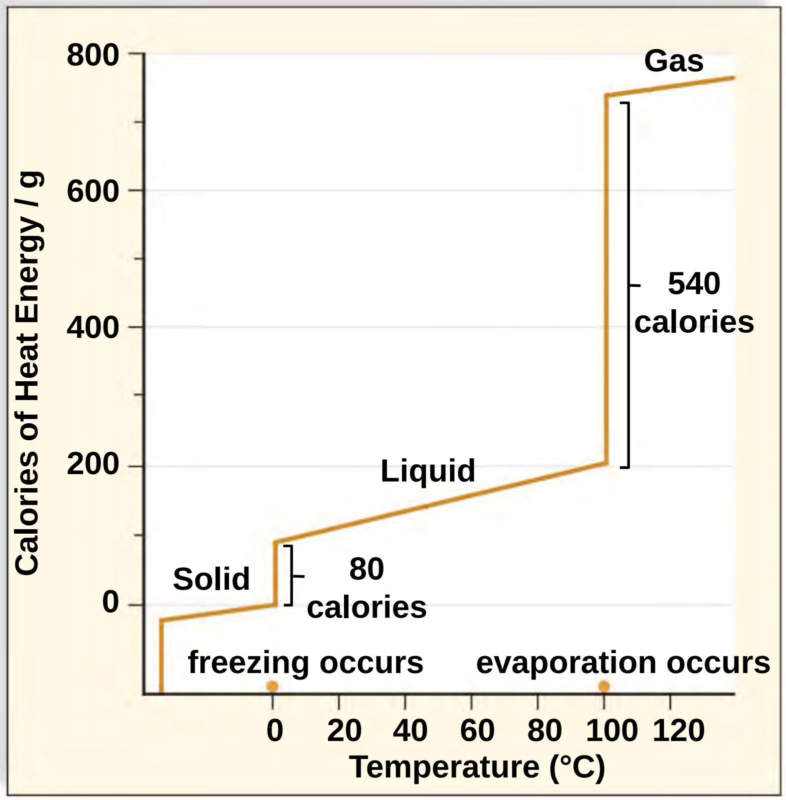


b. Bodies of organisms cool when their heat is used to evaporate water.

Properties of Water: Heat of Fusion

- Heat of fusion (melting)
  - To raise ice from -2 to -1 ºC; ~1 calorie
  - To raise water from -1 to 0 ºC; ~1 calorie
  - To raise water from 0 to 1 ºC; ~80 calories!
- This is why ice at 0 ºC keeps stuff cold MUCH longer than water at 1 ºC
- This is why ice is used for cooling
  - NOT because ice is cold
  - But because it absorbs so much heat before it will warm by one degree

Heat Content of Water at Various Temperatures

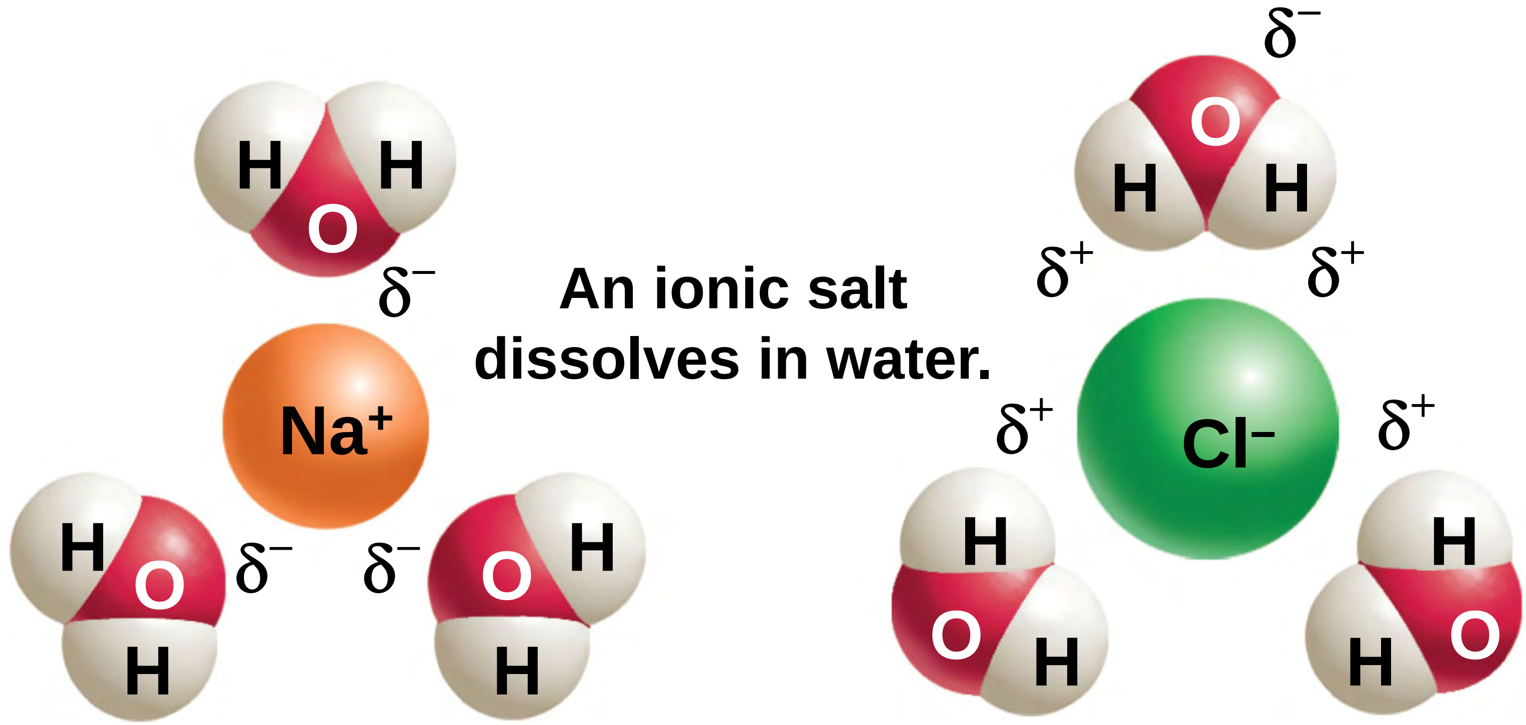


a. Calories lost when 1 g of liquid water freezes and calories required when 1 g of liquid water evaporates.

Properties of Water: Water as a Solvent

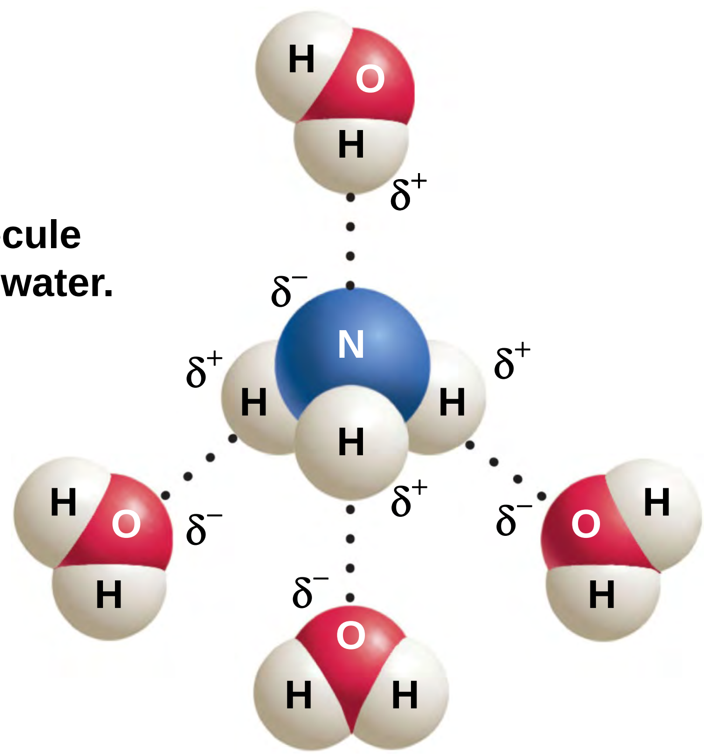
- Solutions consist of:
  - A solvent (the most abundant part) and
  - A solute (less abundant part) that is dissolved in the solvent
- Polar compounds readily dissolve; hydrophilic
- Nonpolar compounds dissolve only slightly; hydrophobic
- Ionic compounds dissociate in water
  - Na^+
    - Attracted to negative (O) end of H_2O
    - Each Na^+ completely surrounded by H_2O
  - Cl^-
    - Attracted to positive (H_2) end of H_2O
    - Each Cl^- completely surrounded by H_2O

Properties of Water: Water as a Solvent



Properties of Water: Water as a Solvent

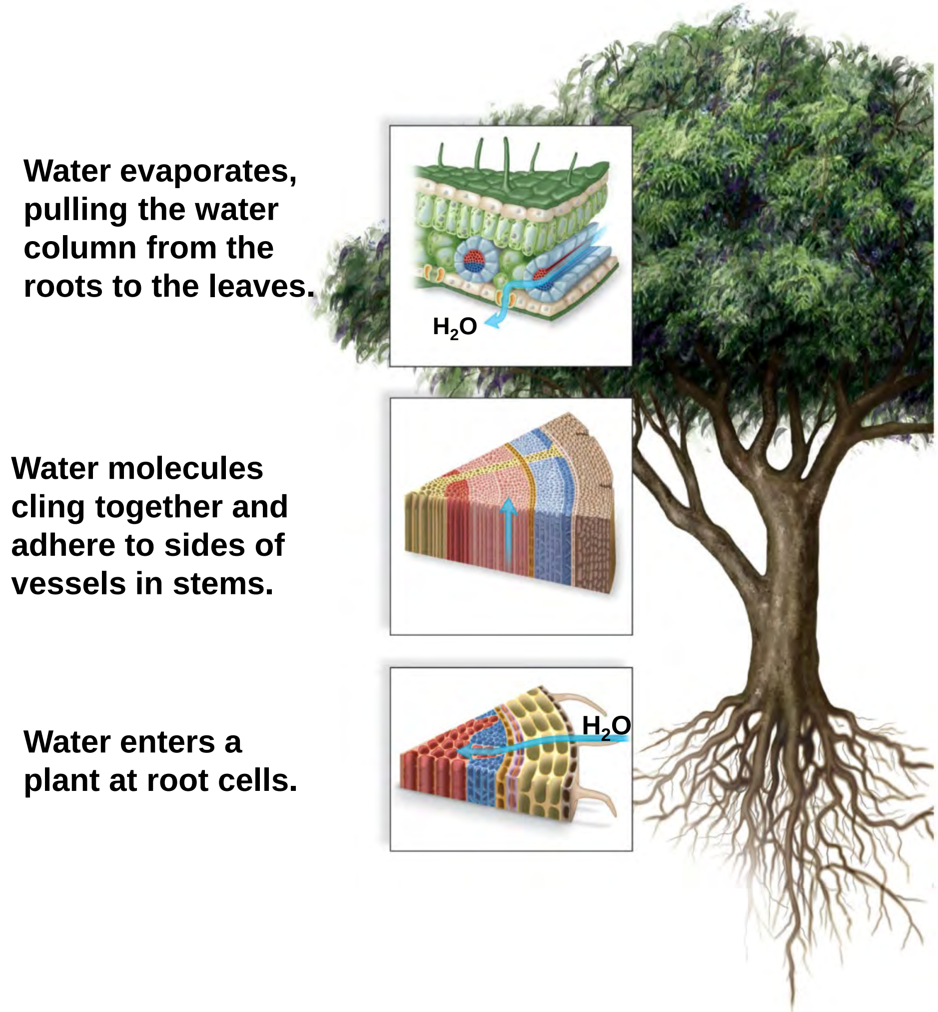
A polar molecule dissolves in water.



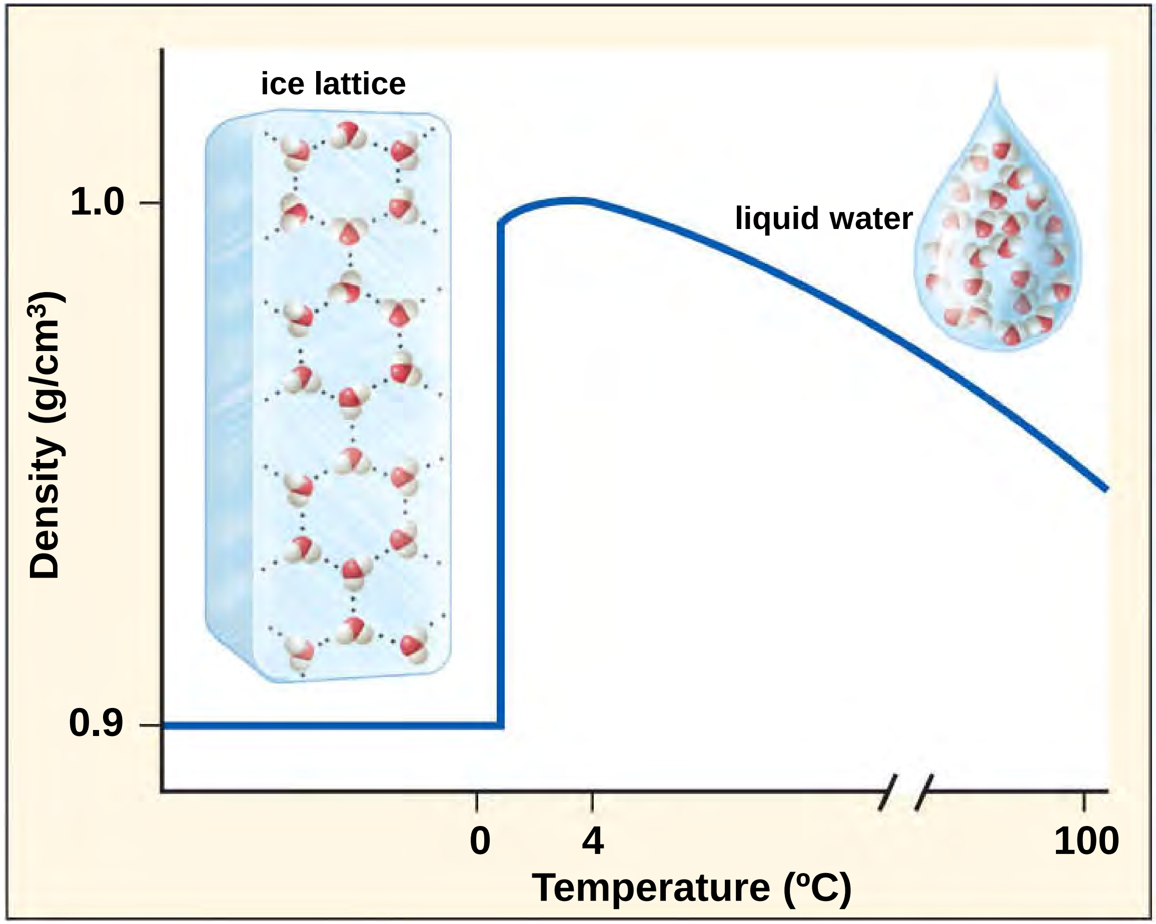
Properties of Water: Uniqueness of Ice

- Frozen water less dense than liquid water
  - Otherwise, oceans and deep lakes would fill with ice from the bottom up
  - Ice acts as an insulator on top of a frozen body of water
  - Melting ice draws heat from the environment

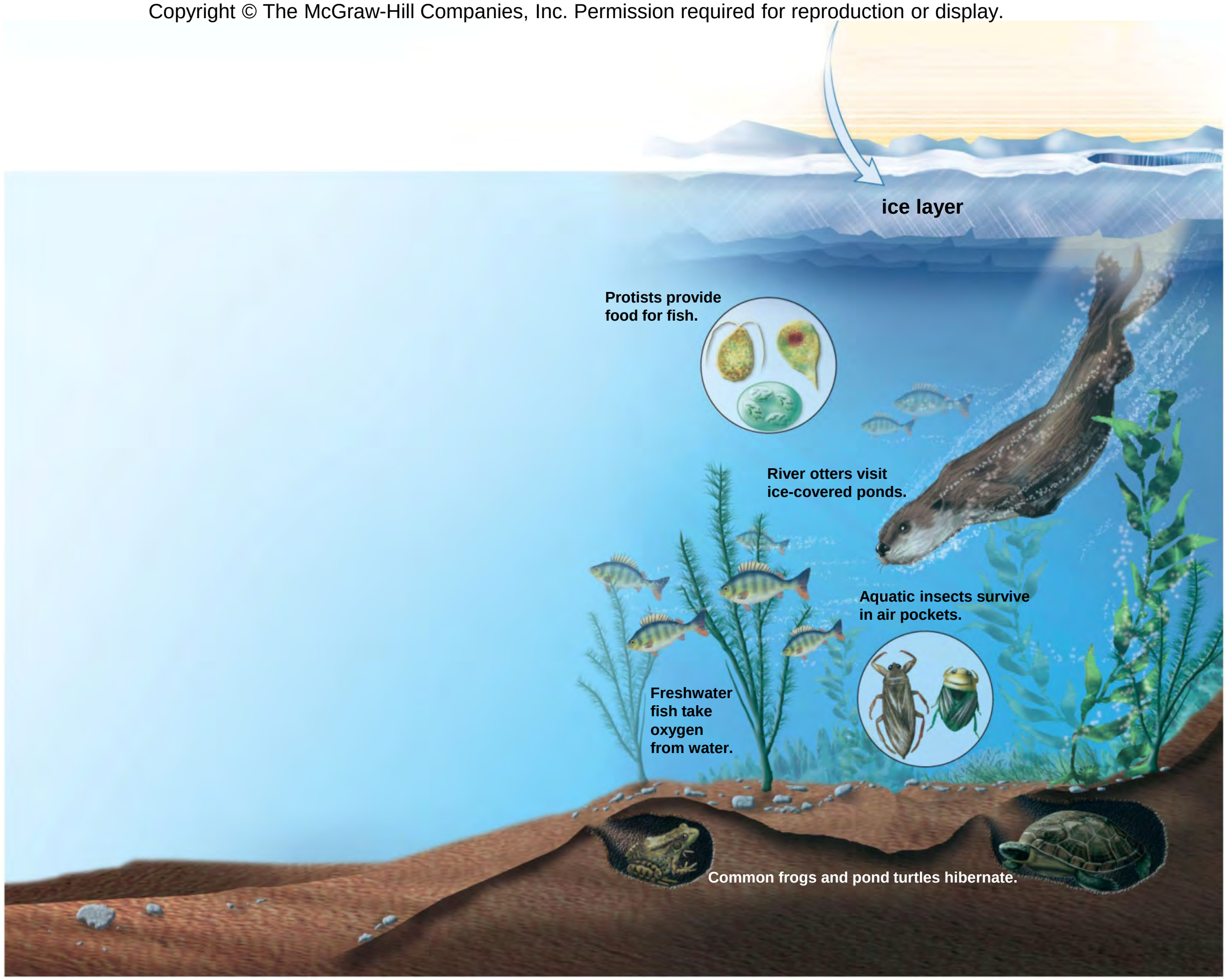
Water as a Transport Medium



Density of Water at Various Temperatures



A Pond in Winter



Properties of Water: Cohesion & Adhesion

- Cohesive and Adhesive
  - Cohesion – Hydrogen bonds hold water molecules tightly together i.e. allows water to flow freely without molecules separating.
  - Adhesion – Hydrogen bonds for between water and other polar materials
  - Allow water be drawn many meters up a tree in a tubular vessel
- High Surface Tension
  - Water molecules at surface hold more tightly than below surface
  - Amounts to an invisible "skin" on water surface
  - Allows small nonpolar objects (like water strider) to sit on top of water http://youtu.be/RphuMEUY3Og

pH of Water: Acids

- pH is a measure of the concentration of hydrogen ions
- When water ionizes or dissociates, it releases a small but equal number of hydrogen (H+) ions and hydroxide (OH-) ions
- Acids donate hydrogen ions
  - Dissociate in water and release hydrogen ions (H^+) e.g. HCl → H+ + Cl
    - Dissociation of HCl is almost total, therefore it is a strong acid
  - Sour to taste

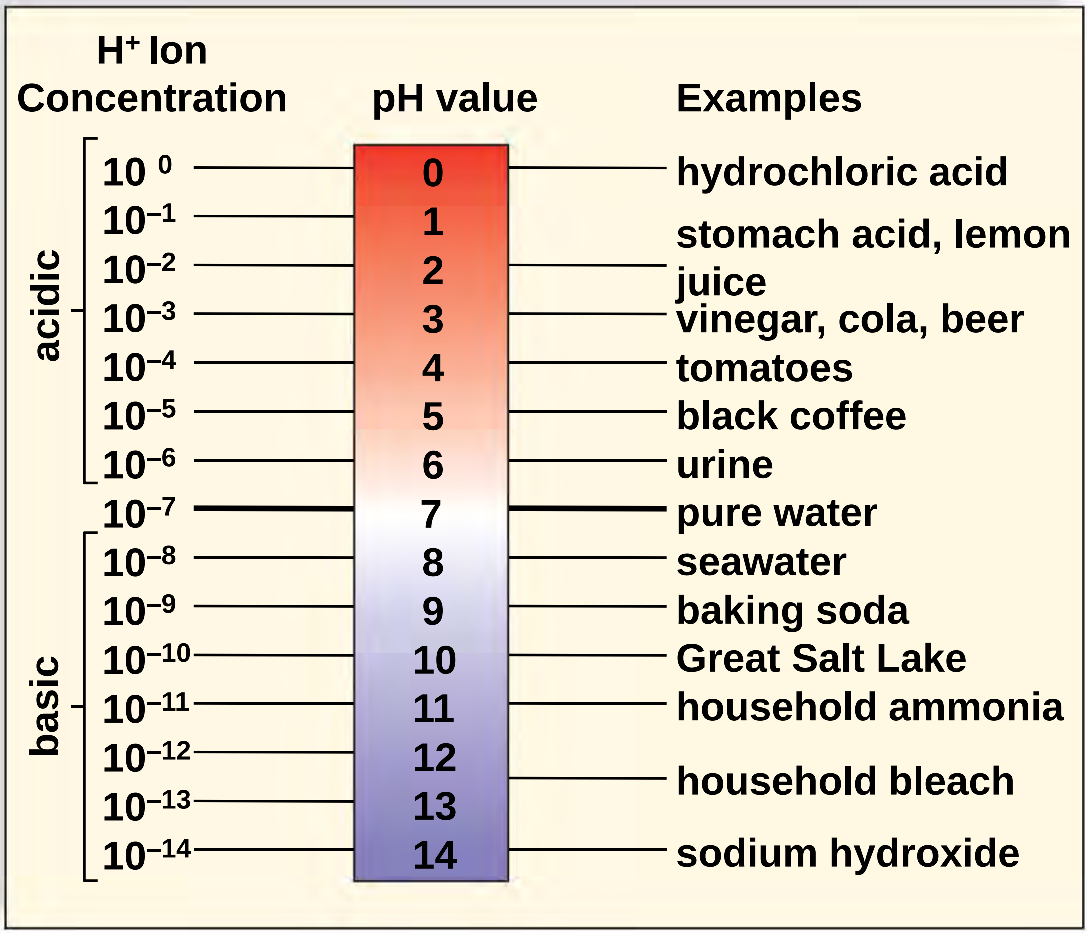
pH of Water: Bases

- Bases remove hydrogen ions
  - Either take up hydrogen ions (H^+) or release hydroxide ions (OH^-)
  - Bitter to taste
  - Sodium hydroxide is a solid with symbol NaOH
    - In water, it dissociates into Na^+ and OH^-
    - Dissociation of NaOH is almost total, therefore it is a strong base

pH Scale

- pH scale used to indicate acidity and alkalinity of a solution.
  - Values range from 0-14
    - 0 to <7 = Acidic
    - 7 = Neutral
    - >7 to 14 = Basic (or alkaline)
  - Logarithmic Scale
    - Each unit change in pH represents a change of 10X
    - pH of 4 is 10X as acidic as pH of 5
    - pH of 10 is 100X more basic than pH of 8

The pH Scale


H+ Ion Concentration
pH value
Examples
acidic
basic
10 0 — 0 — hydrochloric acid
10–1 — 1 — stomach acid, lemon juice
10–2 — 2
10–3 — 3 — vinegar, cola, beer
10–4 — 4 — tomatoes
10–5 — 5 — black coffee
10–6 — 6 — urine
10–7 — 7 — pure water
10–8 — 8 — seawater
10–9 — 9 — baking soda
10–10 — 10 — Great Salt Lake
10–11 — 11 — household ammonia
10–12 — 12
10–13 — 13 — household bleach
10–14 — 14 — sodium hydroxide

Buffers and pH

- When H^+ is added to pure water at pH 7, pH goes down and water becomes acidic
- When OH^- is added to pure water at pH 7, pH goes up and water becomes alkaline
- Buffers are solutes in water that resist change in pH
  - When H^+ is added, buffer may absorb, or counter by adding OH^-
  - When OH^- is added, buffer may absorb, or counter by adding H^+

Buffers in Biology

- Health of organisms requires maintaining pH of body fluids within narrow limits
  - Human blood normally 7.4 (slightly alkaline)
  - Many foods and metabolic processes add or subtract H^+ or OH^- ions
    - Reducing blood pH to 7.0 results in acidosis
    - Increasing blood pH to 7.8 results in alkalosis
    - Both life threatening situations
  - Bicarbonate ion ($^-HCO_3$) in blood buffers pH to 7.4

Review

- Chemical Elements
  - Atoms
  - Isotopes
  - Molecules and Compounds
- Chemical Bonding
  - Ionic and Covalent
  - Hydrogen
- Properties of Water
- Acids and Bases