



# Introduction into Biochemistry

Summer 2023



- Recommended textbooks

- **Marks' Basic Medical Biochemistry: A Clinical Approach 5th Edition, by Michael Lieberman (Author), Alisa Peet MD (Author), 2018**

- Instructors

- **Prof. Mamoun Ahram**
- **Dr. Diala Abu Hassan**



- Biochemistry:

- Introduction
- Acids and bases, pH, buffer
- Macromolecules (carbohydrates, lipids, amino acids, proteins)
- Protein structure-function relationship (part I: fibrous proteins: collagen, elastin, and keratins; part II: globular proteins (myoglobin, hemoglobin, and immunoglobulins))

## ***Midterm***

- Enzymes (introduction, kinetics, mechanisms of regulation, cofactors)
- Protein analysis
- Molecular biology (online/recorded):
  - Restriction endonucleases recombinant DNA technology; polymerase chain reaction (PCR); DNA sequencing, protein-protein interactions; CRISPR-CAS9; DNA microarrays; RNA-seq)



- Main References:

- The lectures
- Marks' Basic Medical Biochemistry by M. Lieberman A. Marks and Alisa Peet, Lippincott, Williams and Wilkins, 2017. 5th edition.
- The Cell: A Molecular Approach, Geoffrey M. Cooper and Robert E. Hausmann, 8th edition, Sinauer Associates, 2019.

- Extra:

- NCBI Bookshelf:
  - (<http://www.ncbi.nlm.nih.gov/entrez/query.fcgi?db=Books>)
- The Medical Biochemistry Page:
  - (<https://themedicalbiochemistrypage.org/>)

# Office hours for Prof. Ahram



- Location: School of Medicine, building 1, first floor (Sunday and Wednesday 3:30-4:30)
- Note: If I am not in my office, try the lab on the first floor. The map to the lab is on the office's door. Simply ask for me.



# Biochemistry & chemical composition of living organisms

# Biochemistry = understanding life

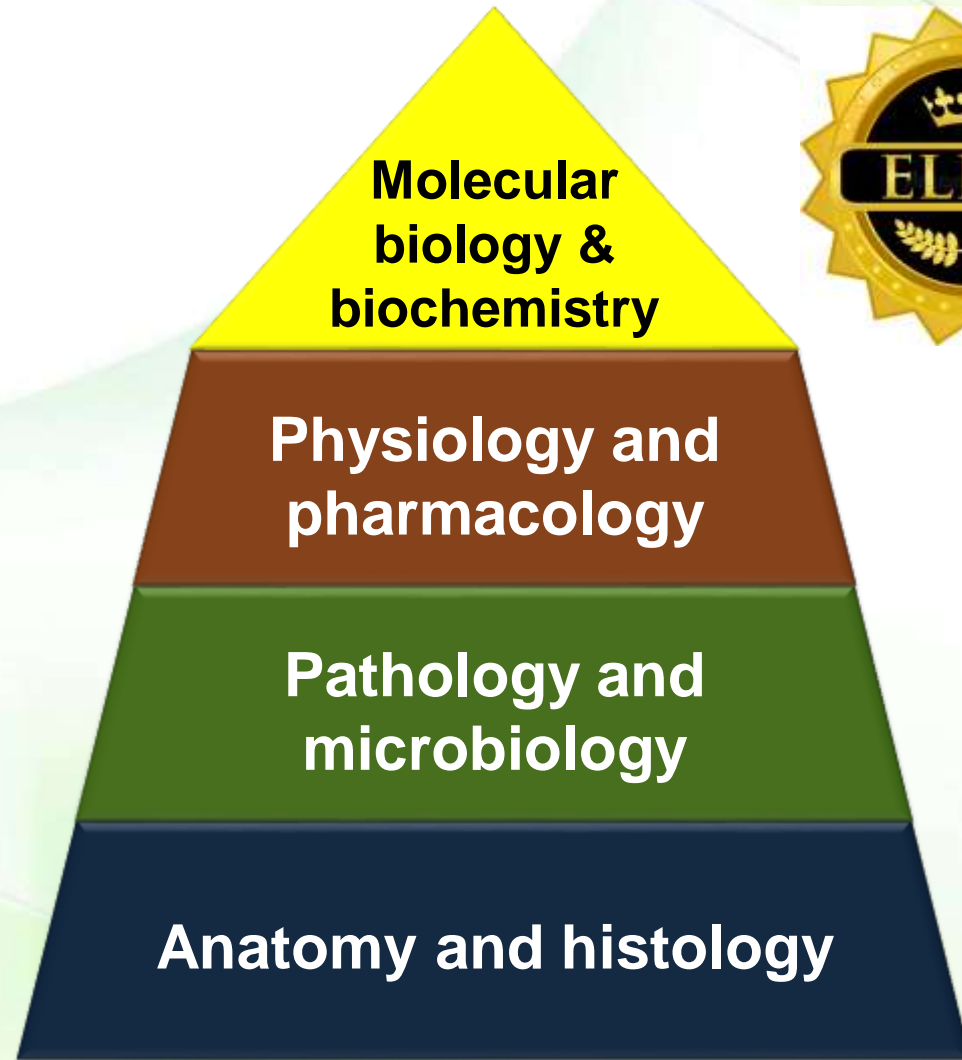


- Know the chemical structures of biological molecules
- Understand the biological function of these molecules
- Understand the interaction and organization of different molecules within individual cells and whole biological systems
- Understand bioenergetics (the study of energy flow in cells)

## ***Biochemistry in medicine:***

- ***explains all disciplines***
- ***diagnose and monitor diseases***
- ***design drugs (new antibiotics, chemotherapy agents)***
- ***understand the molecular bases of diseases***

# Biochemistry is for the elite





# Chemical elements in living creatures



- The human body is composed mainly of ~30 elements.
- Four primary elements: carbon, hydrogen, oxygen, and nitrogen (96.5% of an organism's weight)
- Then, calcium and phosphorus (that's 98.5%).
- Others exist in trace amounts but are essential, elements (mostly metals).

Bulk biological elements
  Trace elements believed to be essential for bacteria, plants or animals
  Possibly essential trace elements for some species

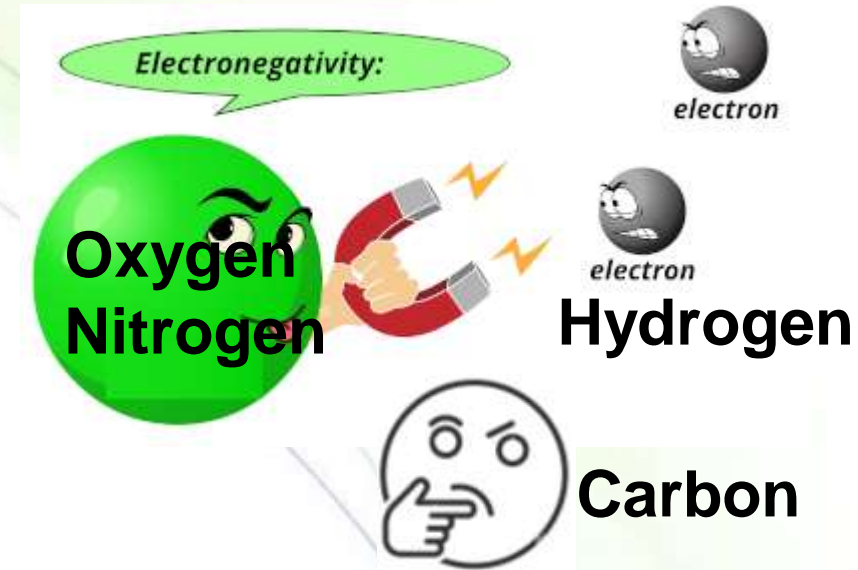
**TABLE 2.1** Elements of the Human Body

Name	Symbol	Percentage of Body Weight	
<b>Major Elements (Total 98.5%)</b>			
Oxygen	O	65.0	
Carbon	C	18.0	
Hydrogen	H	10.0	
Nitrogen	N	3.0	
Calcium	Ca	1.5	
Phosphorus	P	1.0	
<b>Lesser Elements (Total 0.8%)</b>			
Sulfur	S	0.25	
Potassium	K	0.20	
Sodium	Na	0.15	
Chlorine	Cl	0.15	
Magnesium	Mg	0.05	
Iron	Fe	0.006	
<b>Trace Elements (Total 0.7%)</b>			
Chromium	Cr	Molybdenum	Mo
Cobalt	Co	Selenium	Se
Copper	Cu	Silicon	Si
Fluorine	F	Tin	Sn
Iodine	I	Vanadium	V
Manganese	Mn	Zinc	Zn

# Important terms



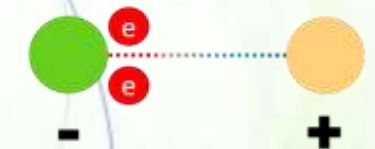
- Electronegativity
- Covalent bonds
  - Polar vs. non-polar covalent bonds
  - Single vs. multiple
- Non-covalent interactions
  - electrostatic interactions
  - hydrogen bonds (donor and acceptor)
  - van der Waals interactions
  - Hydrophobic interactions
    - Hydrophobic versus hydrophilic molecules
    - Backbone of a molecule
- Nucleophile versus electrophile



Covalent Bonds:  
Molecules



Ionic Bonds:  
Ionic Compounds

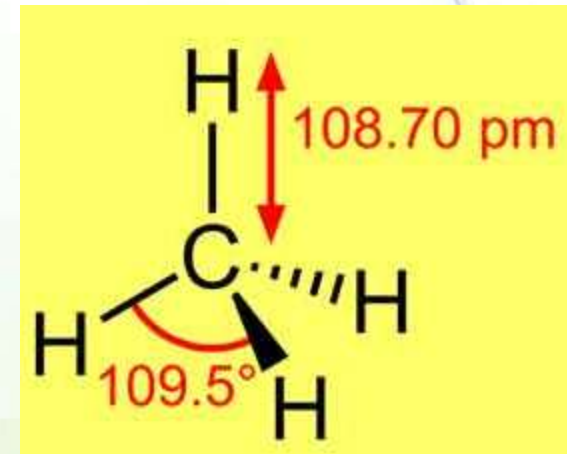


# Important properties of bonds



- Bond strength (amount of energy that must be supplied to break a bond)
- Bond length: the distance between two nuclei
- Bond orientation: bond angles determining the overall geometry of atoms
- The three-dimensional structures of molecules are specified by the bond angles and bond lengths for each covalent linkage.

<b>Bond</b>	H—F	H—Cl	H—Br	H—I
<b>Length</b>	0.917 Å	1.275 Å	1.415 Å	1.609 Å
<b>Strength</b>	136 kcal/mol 571 kJ/mol	103 kcal/mol 432 kJ/mol	87 kcal/mol 366 kJ/mol	71 kcal/mol 298 kJ/mol

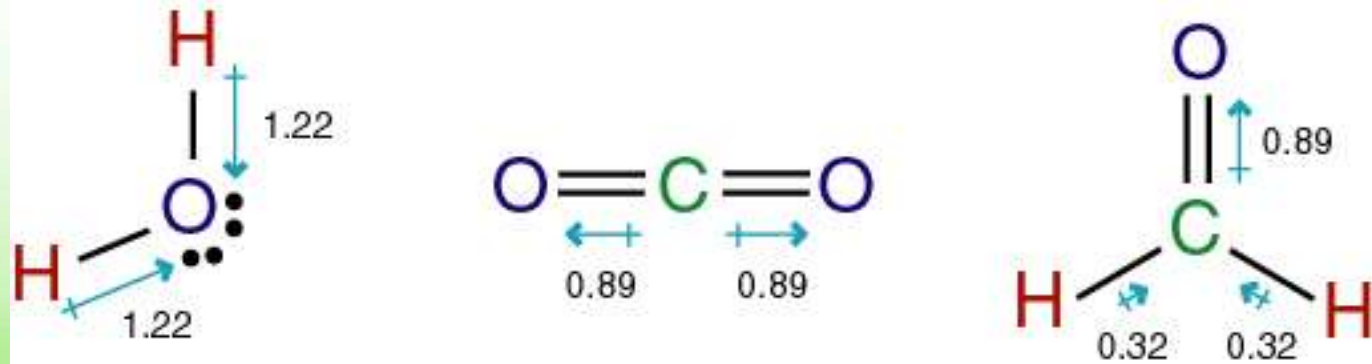


# Polarity of covalent bonds



- Covalent bonds in which the electrons are shared unequally in this way are known as polar covalent bonds. The bonds are known as “dipoles”.
- Oxygen and nitrogen atoms are electronegative
- Oxygen and hydrogen
- Nitrogen and hydrogen
- Not carbon and hydrogen

**Water is an excellent example of polar molecules, but not CO<sub>2</sub>.**



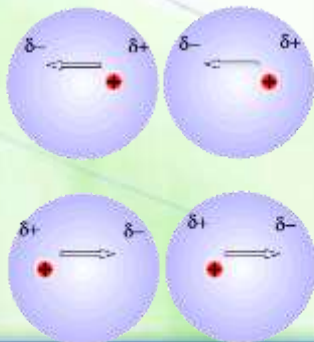
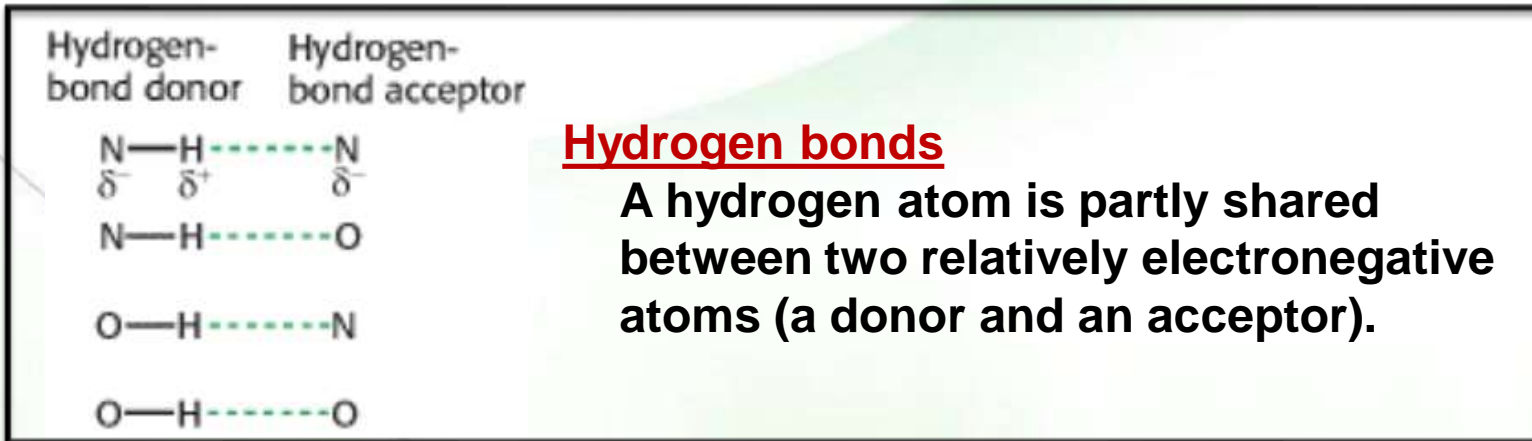
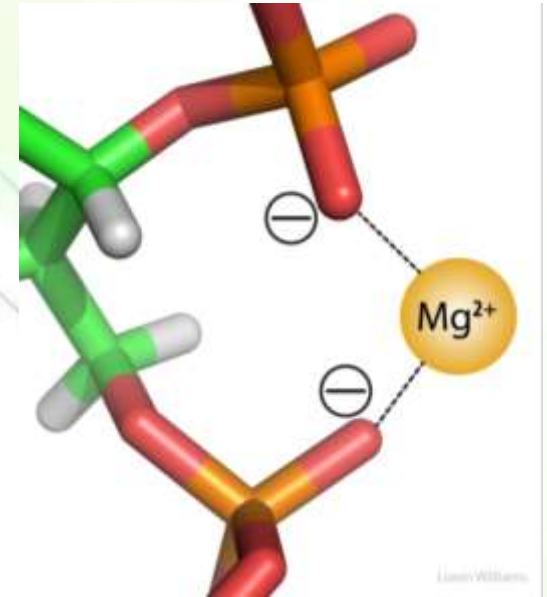
# What are the non-covalent interactions?



- They are reversible and relatively weak.

## Electrostatic interactions (charge-charge interactions):

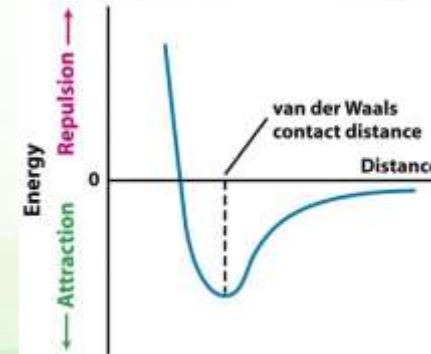
- They are formed between two charged particles.
- These forces are quite strong in the absence of water .



## van der Waals interactions

Unequal distribution of electronic charge around an atom changes with time.

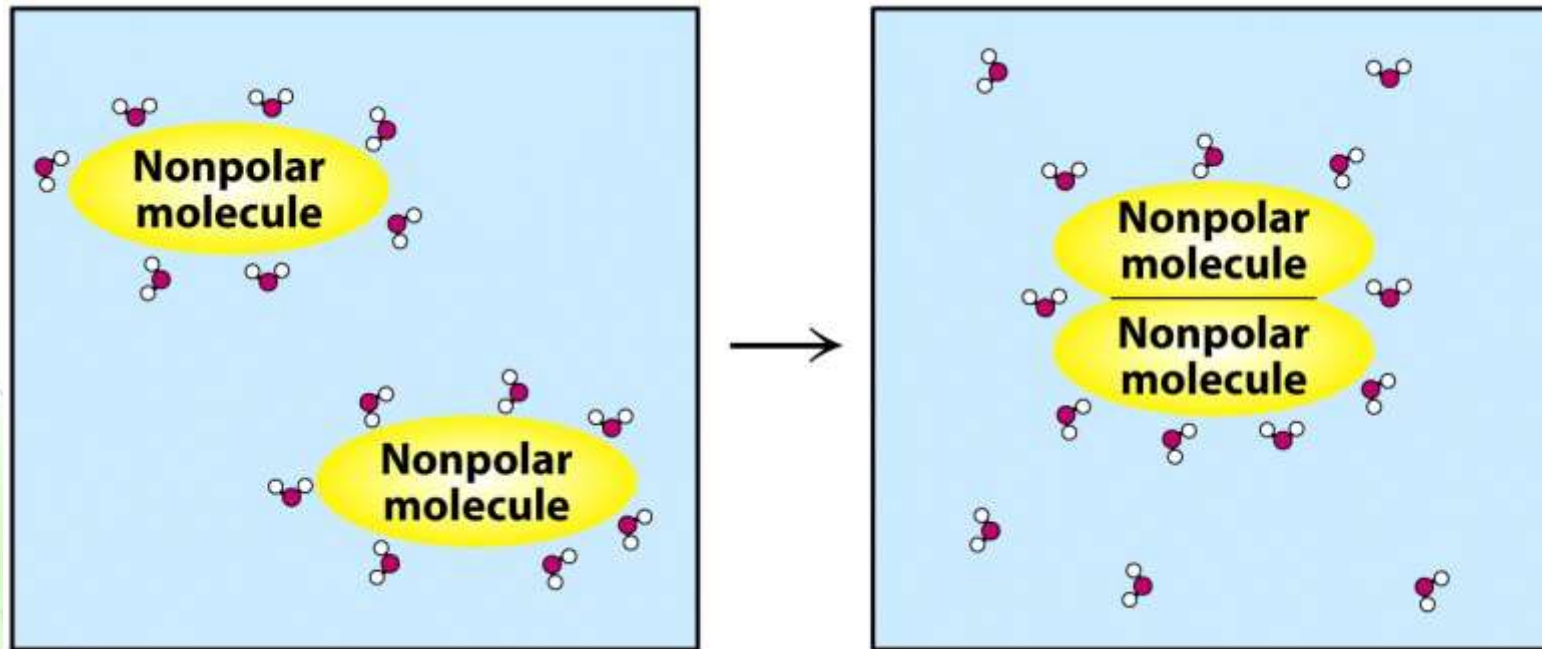
The strength of the attraction is affected by distance.



# Hydrophobic interactions



- Self-association of nonpolar compounds in an aqueous environment
- Minimize unfavorable interactions between nonpolar groups and water



# Properties of noncovalent interactions



- Reversible
- Relatively weak
- Molecules interact and bind specifically.
- Noncovalent forces significantly contribute to the structure, stability, and functional competence of macromolecules in living cells.
- Can be either attractive or repulsive
- Involve interactions both within the biomolecule and between it and other molecules of the surrounding environment



# Carbon

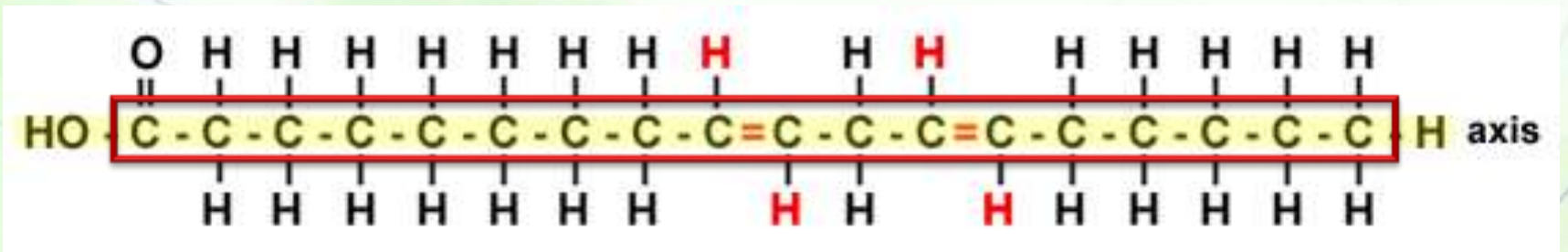
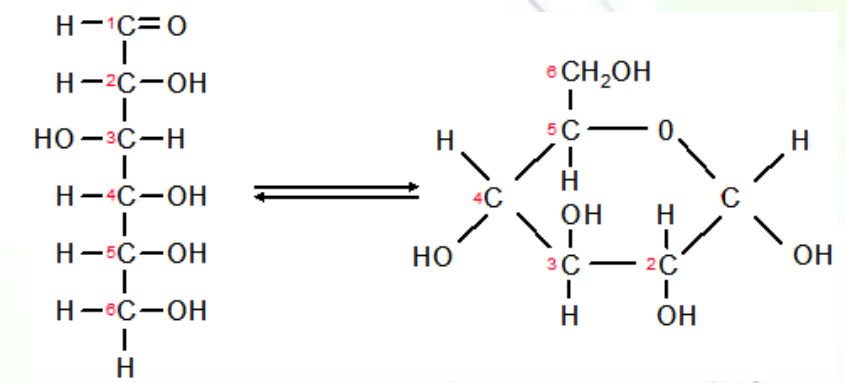
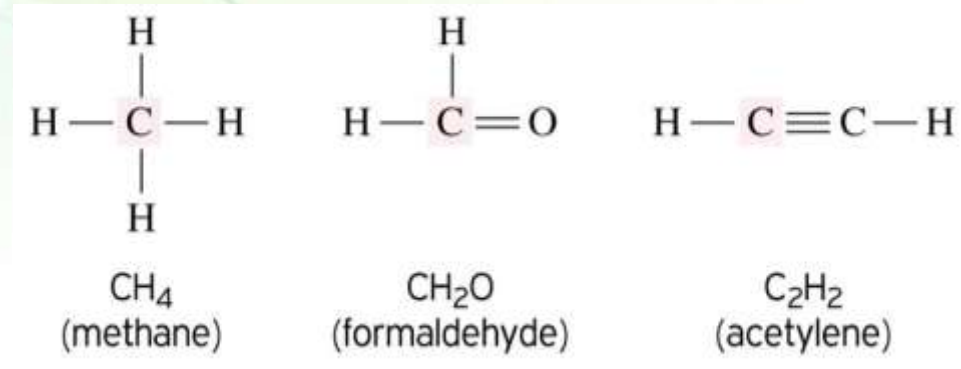
The road to diversity and stability



# Properties of carbon (1)



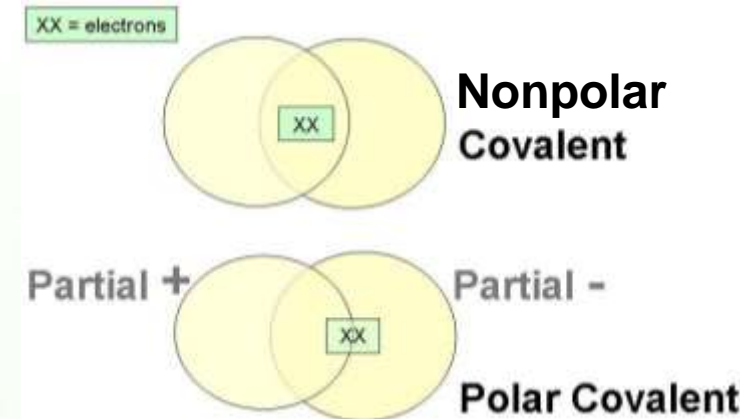
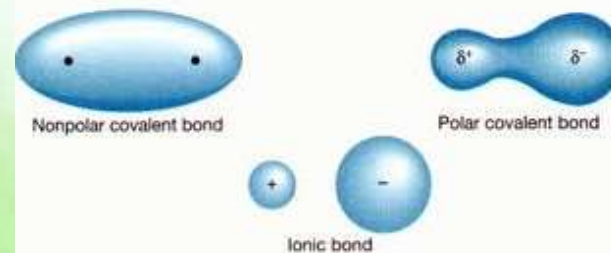
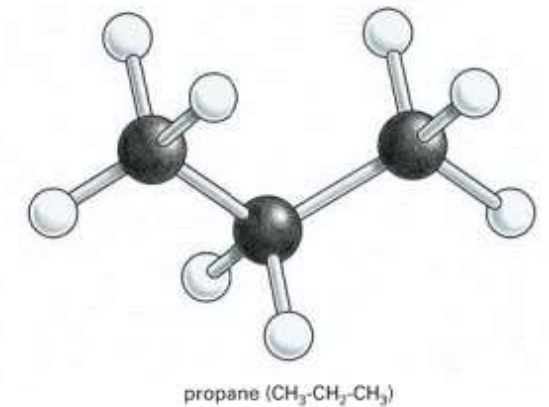
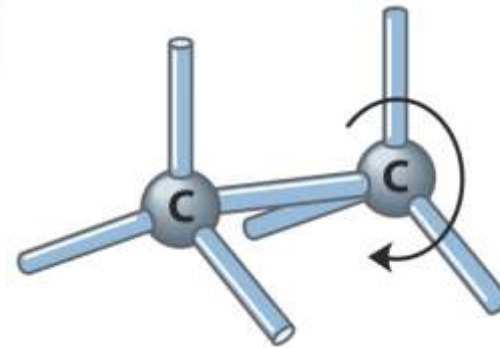
- It can form four bonds, which can be single, double, or triple bonds.
- Each bond is very stable.
  - strength of bonds: triple > double > Single)
- They link C atoms together in chains and rings.
  - These serve as a backbones.



# Properties of carbon (2)



- Carbon bonds have angles giving molecules three-dimensional structures.
- In a carbon backbone, some carbon atoms rotate around a single covalent bond producing molecules of different shapes.
- The electronegativity of carbon is between other atoms.
  - It can form polar and non-polar molecules.
- Pure carbon is not water soluble, but when carbon forms covalent bonds with electronegative elements like O or N, the molecule becomes soluble.



# Functional groups



## (Groups attached to a carbon skeleton)

Class of Compound	General Structure <sup>a</sup>	Functional Group Structure	Functional Group Name	Example
Alkane	$\text{RCH}_2\text{—CH}_3$	$\begin{array}{c}   &   \\ \text{—C—C—} \\   &   \\ \text{H} & \text{H} \end{array}$	Carbon–carbon and carbon–hydrogen single bonds	$\text{H}_3\text{C—CH}_3$
Alkene	$\text{RCH=CH}_2$	$\begin{array}{c} \diagup & \diagdown \\ \text{C}=\text{C} \\ \diagdown & \diagup \end{array}$	Carbon–carbon double bond	$\text{H}_2\text{C=CH}_2$
Alcohol	$\text{ROH}$	$\text{—OH}$	Hydroxyl group	$\text{CH}_3\text{OH}$
Thiol	$\text{RSH}$	$\text{—SH}$	Thiol or sulfhydryl group	$\text{CH}_3\text{SH}$
Ether	$\text{R—O—R}$	$\text{—O—}$	Ether group	$\text{CH}_3\text{—O—CH}_3$
Amine <sup>b)</sup>	$\text{RNH}_2$ $\text{R}_2\text{NH}$ $\text{R}_3\text{N}$	$\begin{array}{c} \diagup \\ \text{—N—} \\ \diagdown \end{array}$	Amino group	$\text{H}_3\text{C—NH}_2$
Imine <sup>b</sup>	$\text{R=NH}$	$\begin{array}{c} \diagup \\ \text{C=N—H} \\ \diagdown \end{array}$	Imino group	$\begin{array}{c} \text{H}_3\text{C} \\ \diagdown \\ \text{C=NH} \\ \diagup \\ \text{H}_3\text{C} \end{array}$
Aldehyde	$\begin{array}{c} \text{O} \\    \\ \text{R—C—H} \end{array}$	$\begin{array}{c} \text{O} \\    \\ \text{—C—H} \end{array}$	Carbonyl group	$\begin{array}{c} \text{O} \\    \\ \text{CH}_3\text{C—} \\ \diagdown \\ \text{H} \end{array}$
Ketone	$\begin{array}{c} \text{O} \\    \\ \text{R—C—R} \end{array}$	$\begin{array}{c} \text{O} \\    \\ \text{—C—} \end{array}$	Carbonyl group	$\begin{array}{c} \text{O} \\    \\ \text{CH}_3\text{CCH}_3 \end{array}$
Carboxylic acid <sup>b</sup>	$\text{R—COOH}$	$\begin{array}{c} \text{O} \\    \\ \text{—C—OH} \end{array}$	Carboxyl group	$\begin{array}{c} \text{O} \\    \\ \text{CH}_3\text{C—} \\ \diagdown \\ \text{OH} \end{array}$



Ester	$\text{R}-\overset{\text{O}}{\parallel}{\text{C}}-\text{OR}$	$-\overset{\text{O}}{\parallel}{\text{C}}-\text{OR}$	Ester group	$\text{CH}_3\overset{\text{O}}{\parallel}{\text{C}}-\text{OCH}_3$
Amide	$\text{R}-\overset{\text{O}}{\parallel}{\text{C}}-\text{NH}_2$	$-\overset{\text{O}}{\parallel}{\text{C}}-\underset{\text{H}}{\overset{\text{H}}{\text{N}}}$	Amide group	$\text{CH}_3\overset{\text{O}}{\parallel}{\text{C}}-\text{NH}_2$
Phosphoric acid <sup>b</sup>	$\text{HO}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$	$\text{HO}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$	Phosphoric acid group	$\text{HO}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$
Phosphoric acid ester <sup>b</sup>	$\text{R}-\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$	$-\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$	Phosphoester group or phosphoryl group	$\text{CH}_3\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$
Phosphoric acid anhydride <sup>b</sup>	$\text{R}-\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2-\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$	$-\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2-\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$	Phosphoric anhydride group	$\text{CH}_3\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2-\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$
Carboxylic acid-phosphoric acid mixed anhydride <sup>b</sup>	$\text{R}-\overset{\text{O}}{\parallel}{\text{C}}-\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$	$-\overset{\text{O}}{\parallel}{\text{C}}-\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$	Acyl-phosphoryl anhydride	$\text{CH}_3\overset{\text{O}}{\parallel}{\text{C}}-\text{O}-\overset{\text{O}}{\parallel}{\text{P}}(\text{OH})_2$

<sup>a</sup> R refers to any carbon-containing group.

<sup>b</sup> These molecules are acids or bases and are able to donate or accept protons under physiological conditions. They may be positively or negatively charged.



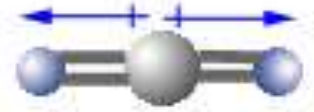
**Water**

# Properties of water (1)



- Water is a polar molecule as a whole because of:
  - the different electronegativities between Hydrogen and oxygen
  - It is angular.
- Water is highly cohesive.
- Water molecules produce a network.
- Water is an excellent solvent because:
  - it is small.
  - it weakens electrostatic forces and hydrogen bonds among polar molecules.

CO<sub>2</sub>

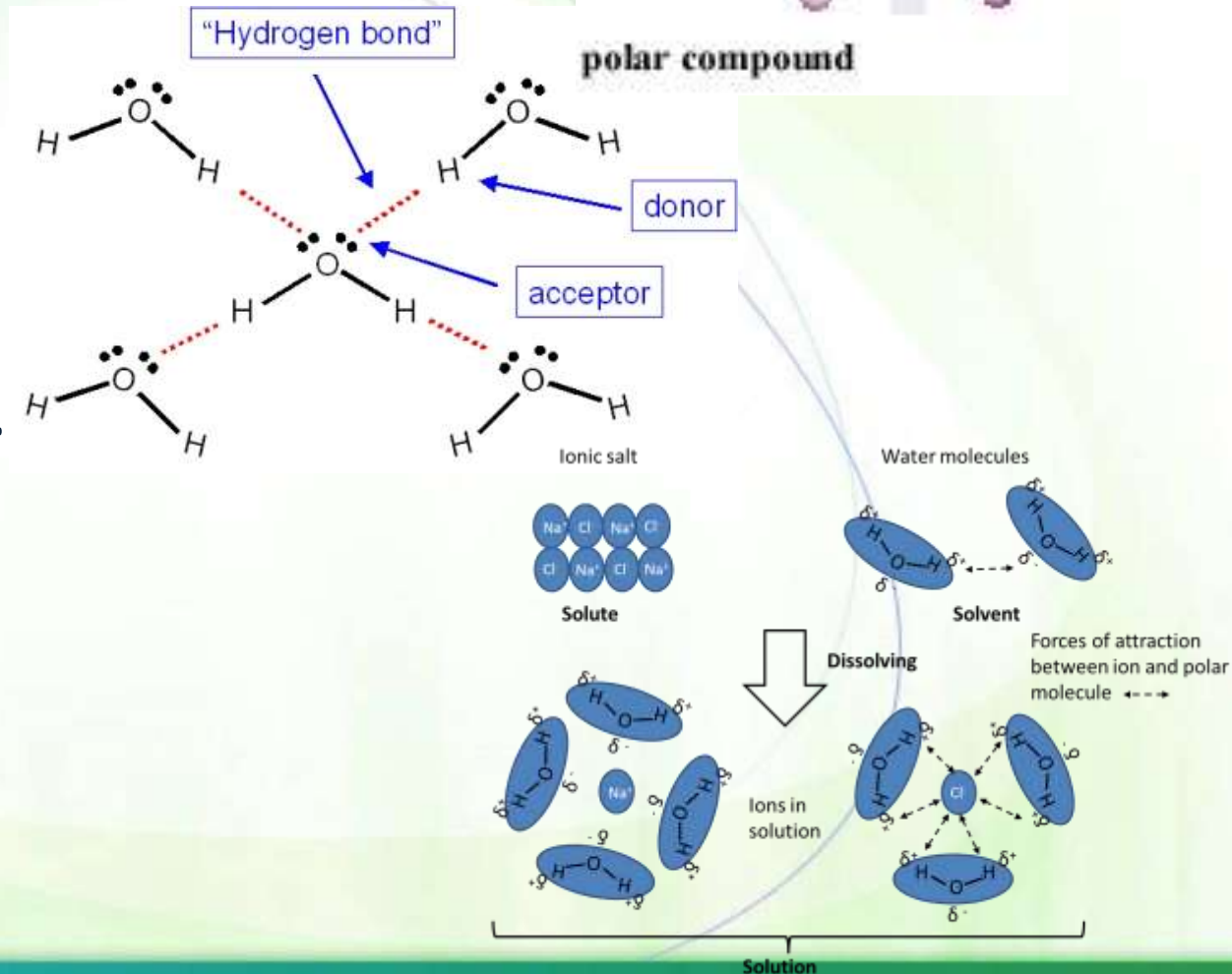


nonpolar compound

H<sub>2</sub>O

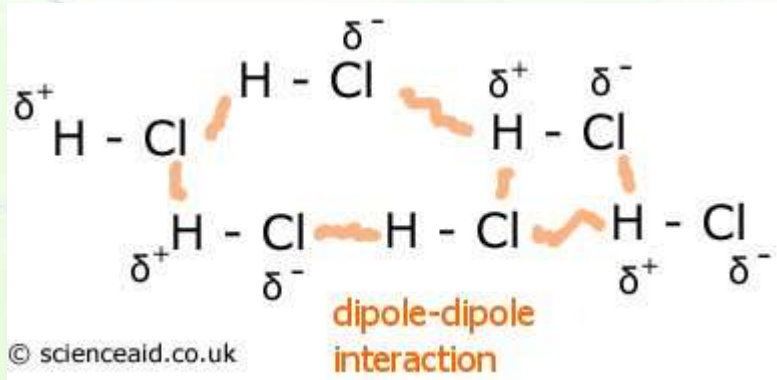
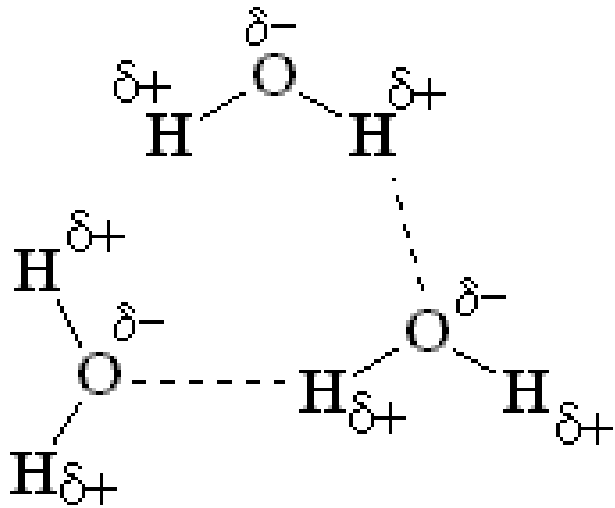


polar compound



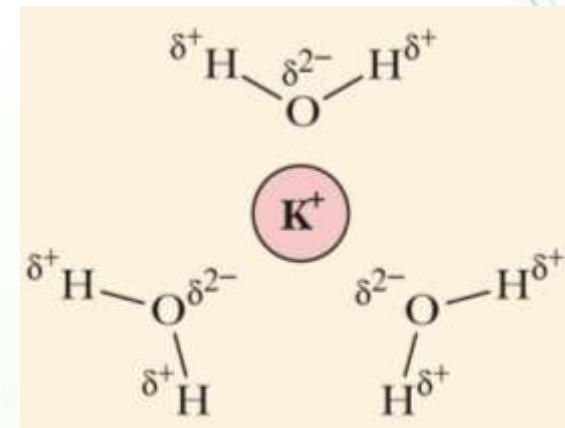


## Dipole-dipole interaction



## Dipole-charge interaction

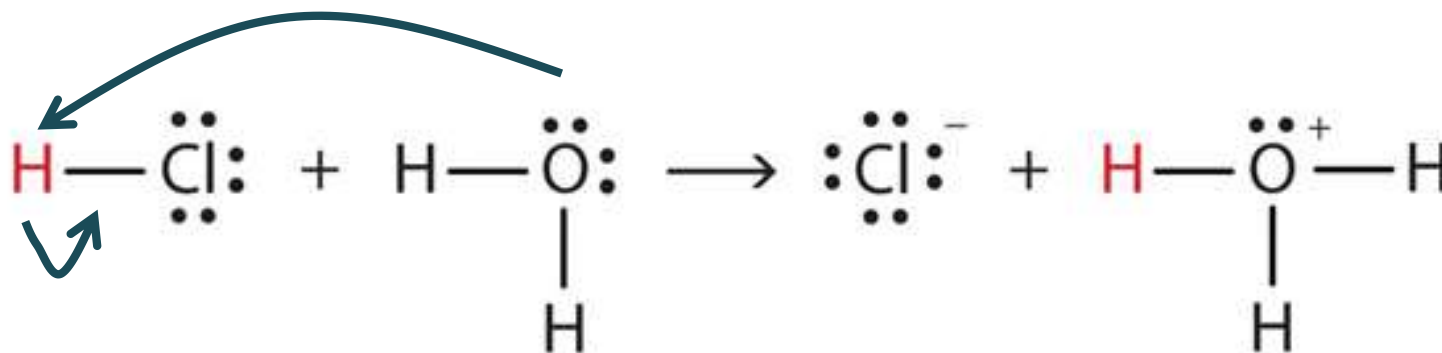
**A** Ion-dipole interactions with water.



# Properties of water (3)



- It is reactive because it is a nucleophile.
  - A nucleophile is an electron-rich molecule that is attracted to positively-charged or electron-deficient species (electrophiles).

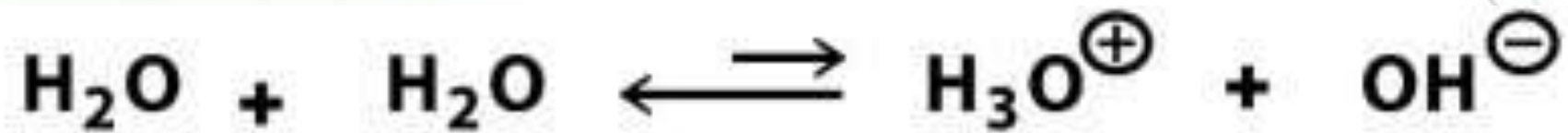




# Properties of water (4)



- Water molecules are ionized to become a positively-charged hydronium ion (or proton), and a hydroxide ion:

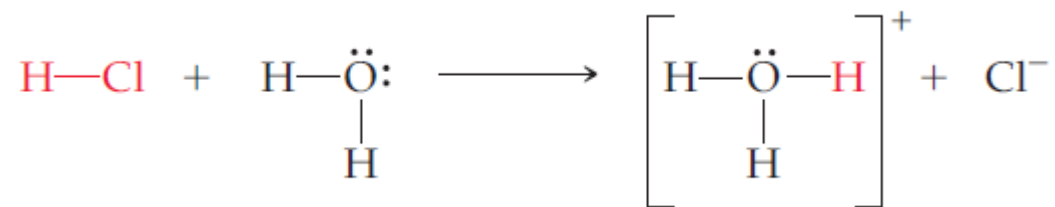


**Note:  $\text{H}_3\text{O}^+ = \text{H}^+$**

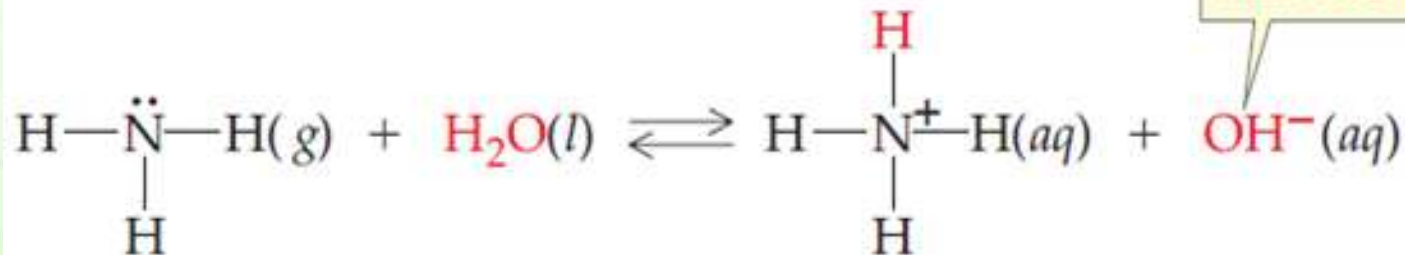
# Brønsted-Lowry acids and bases



- Acid: a substance that produces  $\text{H}^+$  when dissolved in water
  - $\text{H}^+$  Reacts with water producing hydronium ion ( $\text{H}_3\text{O}^+$ ).



- Base: a substance that produces  $\text{OH}^-$  when dissolved in water.



This  $\text{OH}^-$  ion comes from  $\text{H}_2\text{O}$ .

# Types of acids and bases



- The Brønsted-Lowry acid: any substance (proton donor) able to give a hydrogen ion or a proton ( $\text{H}^+$ ) to another molecule.
  - Monoprotic acid:  $\text{HCl}$ ,  $\text{HNO}_3$ ,  $\text{CH}_3\text{COOH}$
  - Diprotic acid:  $\text{H}_2\text{SO}_4$
  - Triprotic acid:  $\text{H}_3\text{PO}_3$
- Brønsted-Lowry base: any substance that accepts a proton ( $\text{H}^+$ ) from an acid.
  - $\text{NaOH}$ ,  $\text{NH}_3$ ,  $\text{KOH}$

# Water = amphoteric



- Substances that can act as an acid in one reaction and as a base in another are called **amphoteric substances**.

- **Example: water**

- With ammonia (NH<sub>3</sub>), water acts as an acid because it donates a H<sup>+</sup> to NH<sub>3</sub>.



- With hydrochloric acid, water acts as a base.



***Ampho = 'both' or 'dual'***

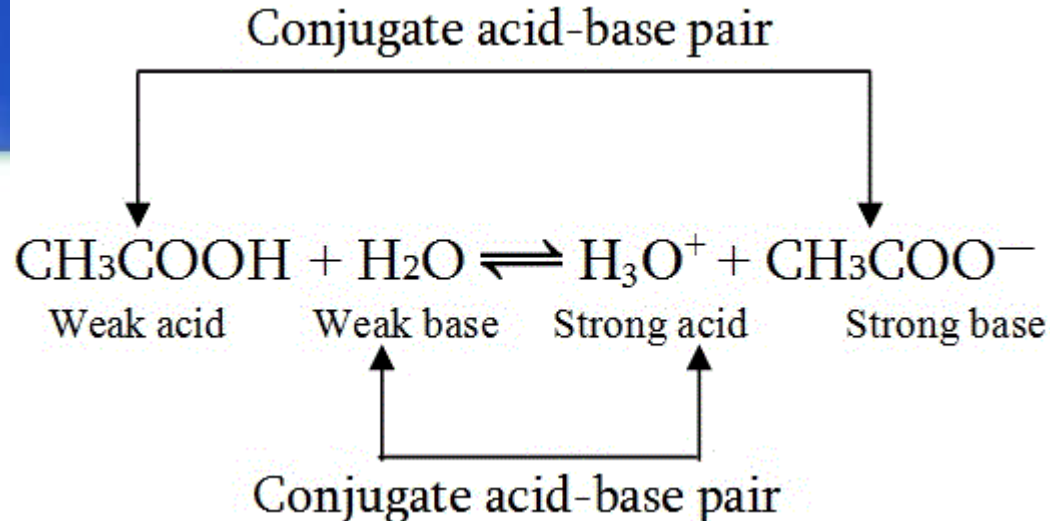
# Acid/base strength



- Acids differ in their ability to release protons.
  - Strong acids dissociate 100%.
- Bases differ in their ability to accept protons.
  - Strong bases have a strong affinity for protons.
- For multi-protic acids ( $\text{H}_2\text{SO}_4$ ,  $\text{H}_3\text{PO}_4$ ), each proton is donated at different strengths.

		ACID	BASE		
100 percent ionized in $\text{H}_2\text{O}$	Strong	HCl	$\text{Cl}^-$	Negligible	Base strength increases
		$\text{H}_2\text{SO}_4$	$\text{HSO}_4^-$		
		$\text{HNO}_3$	$\text{NO}_3^-$		
	Weak	$\text{H}^+$ (aq)	$\text{H}_2\text{O}$	Weak	
		$\text{HSO}_4^-$	$\text{SO}_4^{2-}$		
		$\text{H}_3\text{PO}_4$	$\text{H}_2\text{PO}_4^-$		
		HF	$\text{F}^-$		
		$\text{HC}_2\text{H}_3\text{O}_2$	$\text{C}_2\text{H}_3\text{O}_2^-$		
		$\text{H}_2\text{CO}_3$	$\text{HCO}_3^-$		
		$\text{H}_2\text{S}$	$\text{HS}^-$		
Negligible	$\text{H}_2\text{PO}_4^-$	$\text{HPO}_4^{2-}$	Strong		
	$\text{NH}_4^+$	$\text{NH}_3$			
	$\text{HCO}_3^-$	$\text{CO}_3^{2-}$			
100 percent protonated in $\text{H}_2\text{O}$	$\text{HPO}_4^{2-}$	$\text{PO}_4^{3-}$	Strong		
	$\text{H}_2\text{O}$	$\text{OH}^-$			
	$\text{HS}^-$	$\text{S}^{2-}$			
Negligible	$\text{OH}^-$	$\text{O}_2^-$	Strong		
	$\text{H}_2$	$\text{H}^-$			

# Rule

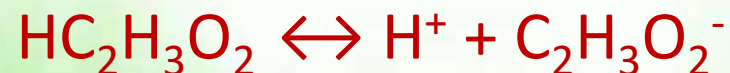


- The stronger the acid, the weaker the conjugate base.
- Strong vs. weak acids

- Strong acids and bases are one-way reactions



- Weak acids and bases do not ionize completely



# Equilibrium constant



- Acid/base solutions are at constant equilibrium.
- We can write equilibrium constant ( $K_{eq}$ ) for such reactions



$$K_a = \frac{[H_3O^+] \cdot [A^-]}{[HA]}$$

**Note:**  $H_3O^+ = H^+$

- The value of the  $K_a$  indicates the direction of the reaction.
  - When  $K_a$  is greater than 1 the product side is favored (more acidic).
  - When  $K_a$  is less than 1 the reactants are favored (less acidic).

# What is pKa?



$$\text{p}K_a = -\log K_a$$

The lower the pKa is,  
the more acidic it is

**TABLE 2.4** Dissociation constants and  $\text{p}K_a$  values of weak acids in aqueous solutions at 25°C

Acid	$K_a(\text{M})$	$\text{p}K_a$
HCOOH (Formic acid)	$1.77 \times 10^{-4}$	3.8
CH <sub>3</sub> COOH (Acetic acid)	$1.76 \times 10^{-5}$	4.8
CH <sub>3</sub> CHOHCOOH (Lactic acid)	$1.37 \times 10^{-4}$	3.9
H <sub>3</sub> PO <sub>4</sub> (Phosphoric acid)	$7.52 \times 10^{-3}$	2.2
H <sub>2</sub> PO <sub>4</sub> <sup>⊖</sup> (Dihydrogen phosphate ion)	$6.23 \times 10^{-8}$	7.2
HPO <sub>4</sub> <sup>2⊖</sup> (Monohydrogen phosphate ion)	$2.20 \times 10^{-13}$	12.7
H <sub>2</sub> CO <sub>3</sub> (Carbonic acid)	$4.30 \times 10^{-7}$	6.4
HCO <sub>3</sub> <sup>⊖</sup> (Bicarbonate ion)	$5.61 \times 10^{-11}$	10.2
NH <sub>4</sub> <sup>⊕</sup> (Ammonium ion)	$5.62 \times 10^{-10}$	9.2
CH <sub>3</sub> NH <sub>3</sub> <sup>⊕</sup> (Methylammonium ion)	$2.70 \times 10^{-11}$	10.7



**TABLE | 9.4  $K_A$  AND  $pK_A$  VALUES FOR SELECTED ACIDS**

Name	Formula	$K_a$	$pK_a$
Hydrochloric acid	HCl	$1.0 \times 10^7$	-7.00
Phosphoric acid	$H_3PO_4$	$7.5 \times 10^{-3}$	2.12
Hydrofluoric acid	HF	$6.6 \times 10^{-4}$	3.18
Lactic acid	$CH_3CH(OH)CO_2H$	$1.4 \times 10^{-4}$	3.85
Acetic acid	$CH_3CO_2H$	$1.8 \times 10^{-5}$	4.74
Carbonic acid	$H_2CO_3$	$4.4 \times 10^{-7}$	6.36
Dihydrogenphosphate ion	$H_2PO_4^-$	$6.2 \times 10^{-8}$	7.21
Ammonium ion	$NH_4^+$	$5.6 \times 10^{-10}$	9.25
Hydrocyanic acid	HCN	$4.9 \times 10^{-10}$	9.31
Hydrogencarbonate ion	$HCO_3^-$	$5.6 \times 10^{-11}$	10.25
Methylammonium ion	$CH_3NH_3^+$	$2.4 \times 10^{-11}$	10.62
Hydrogenphosphate ion	$HPO_4^{2-}$	$4.2 \times 10^{-13}$	12.38

# Expression



- Solutions can be expressed in terms of its concentration or molarity.
- Acids and bases can also be expressed in terms of their normality (N) or equivalence (Eq).

# Molarity of solutions



- Moles of a solution are the amount in grams in relation to its molecular weight (MW or a.m.u.).

$$\text{moles} = \text{grams} / \text{MW}$$

- A molar solution is where the number of grams equal to its molecular weight (moles) in 1 liter of solution.

$$M = \text{moles} / \text{volume (L)}$$

- Since (mol = grams / MW), you can calculate the grams of a chemical you need to dissolve in a known volume (L) of water to obtain a certain concentration (M) using the following formula:

$$\text{grams} = M \times \text{volume (L)} \times \text{MW}$$

# Exercise



- How many grams do you need to make 5M NaCl solution in 100 ml (MW 58.4)?

$$\text{grams} = 58.4 \times 5 \text{ M} \times 0.1 \text{ liter} = 29.29 \text{ g}$$

# Equivalents



- When it comes to acids, bases and ions, it is useful to think of them as equivalents.
- An equivalent is the amount of moles of hydrogen ions that an acid can donate .
  - or a base can accept.
- A 1 g-Eq of any ion is defined as the molar mass of the ion divided by the ionic charge.

# Examples



- For acids:
  - 1 mole HCl = 1 mole  $[H^+]$  = 1 equivalent
  - 1 mole  $H_2SO_4$  = 2 moles  $[H^+]$  = 2 equivalents
    - 1 eq of  $H_2SO_4$  =  $\frac{1}{2}$  mol (because 1 mole gives two  $H^+$  ions)
- For ions:
  - One equivalent of  $Na^+$  = 23.1 g
  - One equivalent of  $Cl^-$  = 35.5 g
  - One equivalent of  $Mg^{2+}$  =  $(24.3)/2 = 12.15$  g

***Remember: One equivalent of any acid neutralizes one equivalent of any base.***

# Molarity and equivalents



$$\text{Equivalents} = n \times M \times \text{volume (L)}$$

***One equivalent of any acid neutralizes one equivalent of a base.***

**Based on the equation above, since  $x$  eq of an acid is neutralized by the same  $x$  eq of a base, then  $(n \times M \times \text{vol})$  of an acid is neutralized by  $(n \times M \times \text{vol})$  of a base.**

**This is done in a process known as Titration.**

**Titration is the slow addition of one solution of a known concentration to a known volume of another solution of unknown concentration until the reaction reaches neutralization.**

# Problem 1



**10.92** Titration of a 12.0 mL solution of HCl requires 22.4 mL of 0.12 M NaOH. What is the molarity of the HCl solution?

- Note that each one produces 1 mole of  $H^+$  or  $OH^-$ , so 1M of HCl is equal to 1M of NaOH, so 1M HCl produces 1M of  $H^+$ , which is neutralized by 1M NaOH, which produces 1M  $OH^-$

Eq of acid = Eq of base

$$N \times M_1 \times Vol_1 = n \times M_2 \times Vol_2$$

$$1 \times 0.12 \times 22.4 = 1 \times M_2 \times 12$$

$$M_2 = (0.12 \times 22.4) / 12$$

$$M_2 = 0.224 \text{ M}$$



# Problem 2



**10.93** What volume of 0.085 M  $\text{HNO}_3$  is required to titrate 15.0 mL of 0.12 M  $\text{Ba}(\text{OH})_2$  solution?

- Note that 1 mole of  $\text{HNO}_3$  produces 1 mole of  $\text{H}^+$ , but 1 mole of  $\text{Ba}(\text{OH})_2$  produces 2 moles of  $\text{OH}^-$ . In other words, the  $n$  is different.
- Also, remember that **Equivalents =  $n \times M \times \text{volume (L)}$** , where  $n$  is the number of charges or the number of  $\text{H}^+$  (or  $\text{OH}^-$ ) the acid or base can produce or accept.
- Titration means that we add an acid to a base slowly. At one point during titration, the acid and the base neutralize or cancel each other. In other words, “to titrate” means “to neutralize”. At the point of neutralization, the concentration of  $\text{H}^+$  is equal to the concentration of  $\text{OH}^-$ . The best way to calculate how much acid is needed to neutralize a base (or the opposite) is to calculate the equivalents.

$$\text{Eq of acid} = \text{Eq of base}$$

$$N \times M_1 \times \text{Vol}_1 = n \times M_2 \times \text{Vol}_2$$

$$1 \times 0.085 \times \text{Vol} = 2 \times 0.12 \times 15$$

$$\text{Vol} = (2 \times 0.12 \times 15) / 1 \times 0.085$$

$$\text{Vol} = 42.35 \text{ mL}$$

# Ionization of water



- Water dissociates into hydronium ( $\text{H}_3\text{O}^+$ ) and hydroxyl ( $\text{OH}^-$ ) ions.
- For simplicity, we refer to  $\text{H}_3\text{O}^+$  as  $\text{H}^+$  and write the reaction equilibrium as:



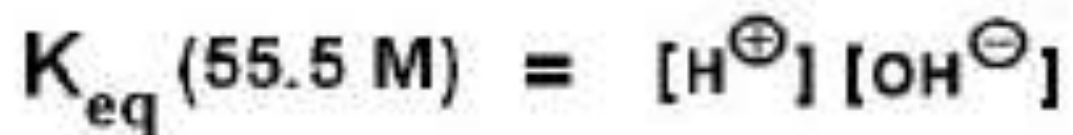
- The equilibrium constant  $K_{\text{eq}}$  of the dissociation of water is:

$$K_{\text{eq}} = \frac{[\text{H}^{\oplus}] [\text{OH}^{\ominus}]}{\text{H}_2\text{O}}$$

- The equilibrium constant for water ionization under standard conditions is  $1.8 \times 10^{-16} \text{ M}$ .



- Since there are 55.6 moles of water in 1 liter, the product of the hydrogen and hydroxide ion concentrations results in a value of  $1 \times 10^{-14}$  for:



- This constant, K<sub>w</sub>, is called the ion product for water

$$K_w = [\text{H}^{\oplus}] [\text{OH}^{\ominus}] = 1.0 \times 10^{-14} \text{ M}^2$$

# [H<sup>+</sup>] and [OH<sup>-</sup>]



- For pure water, there are equal concentrations of [H<sup>+</sup>] and [OH<sup>-</sup>], each with a value of  $1 \times 10^{-7}$  M.
- Since  $K_w$  is a fixed value, the concentrations of [H<sup>+</sup>] and [OH<sup>-</sup>] are inversely changing.
- If the concentration of H<sup>+</sup> is high, then the concentration of OH<sup>-</sup> must be low, and vice versa. For example, if [H<sup>+</sup>] =  $10^{-2}$  M, then [OH<sup>-</sup>] =  $10^{-12}$  M

