

Periodic Table with Chemistry Formulas

CHEMISTRY FORMULAS

ATOMS AND SUBATOMIC PARTICLES

A. Atomic number: The number of protons in an atom of a given element, in the periodic table, elements are ordered by increasing atomic number.

Ex: ^{12}C $Z =$ atomic number

B. Mass number: The number of protons plus the number of neutrons in an atom of an element, measured in atomic mass units (amu). 1 amu = mass of one proton or neutron = 1.66×10^{-27} kg.

Ex: ^{12}C $12 =$ mass number

Name	Symbol	Mass	Charge
Proton	p^+	1 amu	+1
Neutron	n^0	1 amu	0
Electron	e^-	$1/1837$ amu	-1

C. Isotopes: Various atoms of an element that possess different numbers of neutrons, and therefore different mass numbers.

Ex: Boron has ^{10}B , Boron has ^{11}B

D. Calculating atomic mass: An element's atomic mass number in the periodic table is a weighted average of that element's isotopes.

Ex:

Isotope	Abundance	Mass #	Isotope Mass
Boron-10	20.26%	10	10.0129 amu
Boron-11	79.74%	11	11.0093 amu

Atomic mass = $(20.26)(10.0129) + (79.74)(11.0093)$
= 10.81 amu

MOLECULAR MASS

A. Avogadro's number: 6.022×10^{23}

B. The atomic mass number of an element equals the mass (in grams) of one mole of that element.

Ex: One mole of boron has a mass of 10.81 grams.

MOLECULAR FORMULA AND EMPIRICAL FORMULA

A. Molecular formula: The chemical formula of a compound that gives the actual number of atoms of each element in a molecule of the compound.

Ex: The molecular formula for glucose is $\text{C}_6\text{H}_{12}\text{O}_6$.

B. Empirical formula: A "simplified" expression of a chemical formula.

Ex: The empirical formula for glucose is $\text{C}_2\text{H}_4\text{O}$.

BALANCING CHEMICAL EQUATIONS

A. Check for diatomic molecules first: H_2 and O_2 .

- Balance metals.
- Balance nonmetals.
- Balance oxygen.
- Balance hydrogen.

B. You may change only coefficients, never the formula subscripts! Remember all atoms in the equation have equal numbers on both sides of the equation.

C. Reverse the equation so that coefficients are the smallest whole numbers possible.

PERCENTAGE AND PPM

A. Percentage composition: Tells how much mass any object or element contributes to a compound.

Percentage composition = $\frac{\text{mass of element}}{\text{mass of compound}} \times 100$

B. Parts per million (ppm): In a homogeneous solution, the number of grams of a

particular compound that would be present in a one million gram sample of the compound or solution.

TEMPERATURE CONVERSION

A. Converting between Fahrenheit and Celsius

$$\text{Fahrenheit} = \frac{9}{5}(\text{Celsius}) + 32^\circ\text{F}$$

$$\text{Celsius} = \frac{5}{9}(\text{Fahrenheit} - 32^\circ\text{F})$$

B. Kelvin: The temperature in K, absolute zero, is the theoretical lowest possible temperature, at which all molecular activity ceases.

$$0^\circ\text{K} = -273.15^\circ\text{C} = -459.67^\circ\text{F}$$

$$\text{K} = ^\circ\text{C} + 273.15 \quad ^\circ\text{C} = \text{K} - 273.15$$

IDEAL GAS LAWS

P = pressure

V = volume (expressed in liters)

T = temperature (expressed in Kelvin)

Standard temperature and pressure (STP):

$$T = 0^\circ\text{C} = 273.15\text{ K}$$

$$P = 1 \text{ atm} = 760 \text{ mmHg} = 101.3 \text{ kPa}$$

A. Boyle's law: $PV = \text{constant}$

$$P_1V_1 = P_2V_2$$

B. Charles' law: $\frac{V}{T} = \text{constant}$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

C. Gay-Lussac's law: $\frac{P}{T} = \text{constant}$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

D. Combined gas law: $\frac{PV}{T} = \text{constant}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

E. Ideal gas law: $PV = nRT$

$n =$ number of moles, $R =$ ideal gas constant
 $R = 8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}} = 0.0821 \frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$

F. Dalton's law of partial pressures:

$$\frac{P_{\text{gas}}}{P_{\text{total}}} = \frac{n_{\text{gas}}}{n_{\text{total}}}$$

$$\frac{P_{\text{gas}}}{P_{\text{total}}} = \frac{n_{\text{gas}}}{n_{\text{total}}}$$

$$P_{\text{total}} = P_{\text{gas}} + P_{\text{gas}} + \dots$$

SOLUBILITY

A. Molarity (M): Measures the number of moles (n) of solute per liter of solution:

$$M = \frac{n}{V_{\text{solution}}}$$

B. Molarity (m): Measures the number of moles (n) of solute dissolved in 1 kg of solvent:

$$m = \frac{n}{\text{kg of solvent}}$$

C. Dilution: The relationship between initial volume (V_1) and final volume (V_2) is inversely proportional to the relationship between initial molarity (M_1) and final molarity (M_2):

$$\frac{V_1}{V_2} = \frac{M_2}{M_1} \quad M_1V_1 = M_2V_2$$

COLLIGATIVE PROPERTIES

A. Boiling point: Original pressure. Boiling point when solute is added to pure solvent ΔT_b :

$$\Delta T_b = iK_bP_m$$

where P_m is the vapor pressure of A above B is added, K_b is the boiling point constant of A, and P_m is the vapor pressure of B above pure solvent A.

B. Freezing point depression:

$$\Delta T_f = K_f m$$

where ΔT_f is the amount the freezing point is lowered, K_f is a proportionality constant specific to the solvent, and m is the molarity of the solution.

C. Boiling point elevation:

$$\Delta T_b = K_b m$$

ΔT_b is the amount the boiling point is elevated, K_b is a proportionality constant specific to the solvent, and m is the molarity of the solution.

D. Osmotic pressure:

$$\pi = iMRT$$

where π is the osmotic pressure, R is the gas constant of the solution, T is the total gas constant, and T is the temperature in Kelvin.

CHEMICAL EQUILIBRIA

A. Equilibrium constant expression:

Method using molar concentration:
For the reaction: $aA + bB \rightleftharpoons cC + dD$

$$\text{Equilibrium constant } K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Method using gas partial pressures:

$$\text{For the reaction: } aA(g) + bB(g) \rightleftharpoons cC(g) + dD(g)$$

$$\text{Equilibrium constant } K_p = \frac{P_C^c P_D^d}{P_A^a P_B^b}$$

B. Acid-base equilibrium:

1. Acid dissociation reactions: The reaction of a weak acid with H_2O :



Acid dissociation constant: Equilibrium constant for the reaction:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

The greater the K_a value, the stronger the acid.

2. Base dissociation reactions: The reaction of a weak base with H_2O :



Base dissociation constant: Equilibrium constant for the reaction:

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

The greater the K_b value, the stronger the base.

C. Dissociation of water: The self-ionization of water:



$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

D. Conjugate acid-base pair:

$$K_a K_b = K_w = 1.0 \times 10^{-14}$$

E. pH and pOH: pH measures $10^{-\text{pH}}$ measures $10^{-\text{pOH}}$

$$\text{Acids: } \text{pH} < 7 \text{ and } \text{pOH} > 7$$

$$\text{Bases: } \text{pH} > 7 \text{ and } \text{pOH} < 7$$

$$\text{Water: } \text{pH} = \text{pOH} = 7$$

Calculating pH: $\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$

Calculating pOH: $\text{pOH} = -\log_{10}[\text{OH}^-]$

For all aqueous solutions:

$$\text{pH} + \text{pOH} = 14$$

For 25 the rate law is $k[A]^2[B]$, the reaction is second-order with respect to A, and first-order with respect to B.

For 25 the rate law is $k[A]^2[B]$, the reaction is

second-order with respect to A, meaning the reaction rate is proportional to $[A]^2$.

THEORETICAL AND PERCENTAGE YIELD

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

THERMODYNAMICS

A. Specific heat:

$q = mc\Delta T$
where q is the heat absorbed by the object, m is the mass of the object, ΔT is the change in temperature, and c is the specific heat of the substance.

B. Enthalpy (H): In kJ/mol

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

$$\Delta H < 0$$
 for exothermic reactions

$$\Delta H > 0$$
 for endothermic reactions

C. Entropy (S): In J/mol $\cdot$ K

$$\Delta S = S_{\text{products}} - S_{\text{reactants}}$$

D. Gibbs free energy (G): In kJ/mol

$$\Delta G = \Delta H - T\Delta S$$

$$\Delta G < 0$$
 when forward reaction is spontaneous

$$\Delta G > 0$$
 when reverse reaction is spontaneous

$$\Delta G = 0$$
 when system is at equilibrium

REDOX REACTIONS AND ELECTROCHEMISTRY

A. Standard reduction potential of a cell:

$$E_{\text{cell}} = E_{\text{cathode}} + E_{\text{anode}}$$

B. Normal equation:

$$E = E_{\text{cell}} - \frac{RT}{nF} \ln Q$$

where E is the cell potential, Q is the reaction quotient, R is the gas constant, T is the temperature in Kelvin, n is the number of electrons transferred in the half-reactions, F (Faraday's constant) = 96485 C/mol , and Q is the reaction quotient.

C. Standard cell voltage at equilibrium:

$$E_{\text{cell}} = \frac{RT}{nF} \ln K$$

where n is the total number of electrons transferred in the redox reaction.

D. Thermodynamics:

$$\Delta G^\circ = -nFE_{\text{cell}}^\circ$$

$$-nFE_{\text{cell}}^\circ = -RT \ln K_{\text{eq}}$$

$$E_{\text{cell}}^\circ = \frac{RT}{nF} \ln K_{\text{eq}}$$

where n_0 is the starting number of undecayed nuclei, k is the decay constant, and t is the time.

SIGNIFICANT FIGURES

A. Types of significant figures:

- Any nonzero digit is a significant figure.
- Zeros appearing between two nonzero digits are significant.
- Zeros appearing at the end of a number, but only if that number contains a decimal point.

B. Multiplication/division: The result of the calculation should have the same number of significant figures as the least precise number being multiplied or divided.

C. Addition/subtraction: The result of the calculation should have the same number of decimal places as the least precise number being added or subtracted.

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Table of Contents Periodic Table With Chemistry Formulas Sparkcharts

1. Understanding the eBook Periodic Table With Chemistry Formulas Sparkcharts
 - The Rise of Digital Reading Periodic Table With Chemistry Formulas Sparkcharts
 - Advantages of eBooks Over Traditional Books
2. Identifying Periodic Table With Chemistry Formulas Sparkcharts
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Periodic Table With Chemistry Formulas Sparkcharts
 - User-Friendly Interface
4. Exploring eBook Recommendations from Periodic Table With Chemistry Formulas Sparkcharts
 - Personalized Recommendations
 - Periodic Table With Chemistry Formulas Sparkcharts User Reviews and Ratings
 - Periodic Table With Chemistry Formulas Sparkcharts and Bestseller Lists
5. Accessing Periodic Table With Chemistry Formulas Sparkcharts Free and Paid eBooks

- Periodic Table With Chemistry Formulas Sparkcharts Public Domain eBooks
 - Periodic Table With Chemistry Formulas Sparkcharts eBook Subscription Services
 - Periodic Table With Chemistry Formulas Sparkcharts Budget-Friendly Options
6. Navigating Periodic Table With Chemistry Formulas Sparkcharts eBook Formats
 - ePub, PDF, MOBI, and More
 - Periodic Table With Chemistry Formulas Sparkcharts Compatibility with Devices
 - Periodic Table With Chemistry Formulas Sparkcharts Enhanced eBook Features
 7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Periodic Table With Chemistry Formulas Sparkcharts
 - Highlighting and Note-Taking Periodic Table With Chemistry Formulas Sparkcharts
 - Interactive Elements Periodic Table With Chemistry Formulas Sparkcharts
 8. Staying Engaged with Periodic Table With Chemistry Formulas Sparkcharts
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Periodic Table With Chemistry Formulas Sparkcharts
 9. Balancing eBooks and Physical Books Periodic Table With Chemistry Formulas Sparkcharts
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Periodic Table With Chemistry Formulas Sparkcharts
 10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
 11. Cultivating a Reading Routine Periodic Table With Chemistry Formulas Sparkcharts
 - Setting Reading Goals Periodic Table With Chemistry Formulas Sparkcharts
 - Carving Out Dedicated Reading Time
 12. Sourcing Reliable Information of Periodic Table With Chemistry Formulas Sparkcharts
 - Fact-Checking eBook Content of Periodic Table With Chemistry Formulas Sparkcharts
 - Distinguishing Credible Sources
 13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development

- Exploring Educational eBooks

14. Embracing eBook Trends

- Integration of Multimedia Elements
- Interactive and Gamified eBooks

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