

Thermodynamics for Class 11 Worksheet (Includes Answers)

Resources:

www.youtube.com/watch?v=iHedByuub4Y

www.somethingcalledscience.com/post/thermodynamics-chemistry-class-11

Name: _____ Date: _____

Part 1: Understanding the Basics

Thermodynamics Worksheet

Part A: Conceptual Understanding

1. Explain the difference between a system and its surroundings. Give an example of each from everyday life.
2. Classify each of the following as an open, closed, or isolated system:
 - a) A sealed pressure cooker on a hot stove
 - b) A person breathing
 - c) An insulated thermos containing hot coffee
 - d) A plant photosynthesizing in sunlight
3. State the First Law of Thermodynamics in your own words. Why is it also called the Law of Conservation of Energy?
4. What is the difference between heat (q) and work (w)? Give an example of each.
5. Explain why enthalpy (H) is particularly useful for studying chemical reactions that occur in open beakers in the laboratory.
6. What is the difference between an exothermic and endothermic reaction? For each type, state:
 - a) The sign of ΔH
 - b) Whether the reaction feels hot or cold
 - c) An example from everyday life

7. What is a "state function"? Is enthalpy a state function or a path function? What about heat and work?
8. Explain Hess's Law in your own words. Why is it useful in chemistry?
9. Why does water take longer to heat up than cooking oil, even when the same amount of heat is applied?
10. In calorimetry, we use the equation $-q(\text{reaction}) = q(\text{water})$. What principle does this represent?

Part B: Sign Conventions

11. For each process below, indicate whether q is positive or negative, whether w is positive or negative (if applicable), and whether ΔU increases or decreases:
 - a) A gas absorbs 150 J of heat and expands, doing 40 J of work
 - b) A gas releases 200 J of heat while being compressed with 75 J of work done on it
 - c) A system absorbs 300 J of heat at constant volume (no work done)

Part C: Calculations with the First Law

12. A gas absorbs 500 J of heat and does 200 J of work by expanding. Calculate the change in internal energy (ΔU).
13. A system's internal energy increases by 350 J while 120 J of work is done on the system (compression). How much heat was absorbed or released by the system?
14. During a chemical reaction, a gas releases 450 J of heat to its surroundings and has 180 J of work done on it. What is the change in internal energy?
15. A gas does 85 J of work during expansion. If the change in internal energy is +45 J, how much heat was absorbed or released?

Part D: Heat Capacity and Temperature Changes

16. How much heat is required to raise the temperature of 250 g of water from 20°C to 75°C? (Specific heat capacity of water = 4.18 J/g°C)
17. A 150 g piece of iron at 95°C is placed in 400 g of water at 22°C. If the final temperature of both is 25.8°C, calculate the specific heat capacity of iron. (Specific heat capacity of water = 4.18 J/g°C)

18. How much heat is released when 80 g of water cools from 60°C to 15°C? (Specific heat capacity of water = 4.18 J/g°C)

19. A 200 g sample of aluminum releases 2500 J of heat. If its initial temperature was 150°C, what is its final temperature? (Specific heat capacity of aluminum = 0.90 J/g°C)

Part E: Enthalpy Changes

20. The combustion of methane follows the equation: $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$ $\Delta H = -890 \text{ kJ/mol}$

- a) Is this reaction exothermic or endothermic?
- b) How much heat is released when 3.5 moles of methane are burned?
- c) How much heat is released when 32 g of methane (molar mass = 16 g/mol) are burned?

21. The dissolution of ammonium nitrate in water is endothermic: $\text{NH}_4\text{NO}_3(\text{s}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{NO}_3^-(\text{aq})$ $\Delta H = +25.7 \text{ kJ/mol}$

How much heat is absorbed when 40 g of NH_4NO_3 (molar mass = 80 g/mol) dissolves in water?

22. When 2.50 g of ethanol ($\text{C}_2\text{H}_5\text{OH}$) burns completely, 74.5 kJ of heat is released. Calculate the molar enthalpy of combustion of ethanol. (Molar mass of ethanol = 46 g/mol)

Part F: Hess's Law

23. Given the following reactions:

- $\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$ $\Delta H_1 = +68 \text{ kJ}$
- $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$ $\Delta H_2 = -114 \text{ kJ}$

Calculate ΔH for: $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$

24. Calculate the enthalpy change for the reaction: $2\text{C}(\text{s}) + 3\text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})$

Given:

- $\text{C}_2\text{H}_6(\text{g}) + 7/2\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$ $\Delta H_1 = -1560 \text{ kJ}$
- $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$ $\Delta H_2 = -394 \text{ kJ}$
- $\text{H}_2(\text{g}) + 1/2\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$ $\Delta H_3 = -286 \text{ kJ}$

25. Using the following equations:

- $\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$ $\Delta H_1 = -297 \text{ kJ}$
- $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g})$ $\Delta H_2 = -198 \text{ kJ}$

Calculate ΔH for: $2\text{S}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g})$

Part G: Calorimetry

26. When 5.00 g of a substance burns in a calorimeter containing 1500 g of water, the temperature rises from 20.0°C to 32.5°C. Calculate the heat released by the combustion. (Specific heat capacity of water = 4.18 J/g°C)

27. A 50.0 g sample of metal at 100.0°C is placed in 200 g of water at 22.0°C in an insulated calorimeter. The final temperature is 26.4°C. Calculate the specific heat capacity of the metal. (Specific heat capacity of water = 4.18 J/g°C)

28. In a coffee cup calorimeter, 100 mL of 1.0 M HCl is mixed with 100 mL of 1.0 M NaOH, both initially at 21.5°C. The final temperature is 28.2°C. Calculate the enthalpy of neutralization per mole of water formed. Assume the density and specific heat capacity of the solutions are the same as water (1.0 g/mL and 4.18 J/g°C).

Part H: Application and Analysis

29. A student performs an experiment where a reaction absorbs 250 J of heat and 80 J of work is done on the system. Another student performs the same reaction but in two steps: in step 1, the system absorbs 150 J and does 30 J of work; in step 2, the system absorbs 100 J and has 110 J of work done on it. Show that both methods give the same change in internal energy, demonstrating that internal energy is a state function.

30. Explain why the standard enthalpy of formation of O₂(g) is zero, but the standard enthalpy of formation of O₃(g) is not zero.

ANSWERS

Part A: Conceptual Understanding

1. The system is the specific part of the universe we're studying (e.g., the chemicals in a reaction flask). The surroundings are everything else outside the system (e.g., the flask itself, the air, the lab bench, the room). Example: If studying ice melting in a glass, the ice is the system and the glass, air, and everything else is the surroundings.

2.

- a) Closed system (energy can escape as heat but the matter stays inside)
- b) Open system (both matter—oxygen in, CO₂ out—and energy can be exchanged)
- c) Isolated system (ideally, neither energy nor matter can escape)
- d) Open system (matter—CO₂ in, O₂ out—and energy—sunlight in—can be exchanged)

3. The First Law states that energy cannot be created or destroyed, only converted from one form to another. The total energy of the universe remains constant. It's called the Law of Conservation of Energy because energy is always conserved—it just changes form.

4. Heat (q) is energy transferred due to a temperature difference (e.g., warmth from a fire heating your hands). Work (w) is energy transferred when a force causes movement (e.g., a gas expanding and pushing a piston).

5. Most lab reactions occur at constant atmospheric pressure in open containers. Enthalpy (H) is defined such that at constant pressure, the heat absorbed or released equals the change in enthalpy ($\Delta H = q_p$), making it particularly convenient for these conditions.

6.

- **Exothermic:** Releases heat to surroundings
 - a) ΔH is negative
 - b) Feels hot
 - c) Example: burning wood, hand warmers, combustion
- **Endothermic:** Absorbs heat from surroundings
 - a) ΔH is positive
 - b) Feels cold
 - c) Example: melting ice, evaporating water, instant cold packs

7. A state function depends only on the initial and final states, not on the path taken. Enthalpy is a state function—it doesn't matter how you get from reactants to products; ΔH is the same. Heat (q) and work (w) are path functions because they depend on how the process occurs.

8. Hess's Law states that the total enthalpy change for a reaction is the same regardless of the route taken. If a reaction occurs in multiple steps, the sum of the enthalpy changes for each step equals

the total enthalpy change. This is useful because we can calculate enthalpy changes for reactions that are difficult to measure directly by using known enthalpy changes from other reactions.

9. Water has a higher specific heat capacity ($4.18 \text{ J/g}^\circ\text{C}$) than cooking oil (about $2.0 \text{ J/g}^\circ\text{C}$). This means water requires more energy to raise its temperature by the same amount, so it takes longer to heat up.

10. This represents the conservation of energy. In an insulated calorimeter, no heat escapes to the outside environment, so all heat released by the reaction must be absorbed by the water (and vice versa). The total energy remains constant.

Part B: Sign Conventions

11.

- a) $q = +150 \text{ J}$ (absorbed), $w = +40 \text{ J}$ (done by system), ΔU increases ($\Delta U = 150 - 40 = +110 \text{ J}$)
- b) $q = -200 \text{ J}$ (released), $w = -75 \text{ J}$ (done on system), ΔU decreases ($\Delta U = -200 - (-75) = -125 \text{ J}$)
- c) $q = +300 \text{ J}$ (absorbed), $w = 0 \text{ J}$ (no work), ΔU increases ($\Delta U = +300 \text{ J}$)

Part C: Calculations with the First Law

12. $\Delta U = q - w = 500 - 200 = +300 \text{ J}$

13. $\Delta U = q - w$, so $q = \Delta U + w = 350 + (-120) = +230 \text{ J}$ (heat absorbed)

14. $\Delta U = q - w = (-450) - (-180) = -450 + 180 = -270 \text{ J}$

15. $\Delta U = q - w$, so $q = \Delta U + w = 45 + 85 = +130 \text{ J}$ (heat absorbed)

Part D: Heat Capacity and Temperature Changes

16. $q = mc\Delta T = 250 \times 4.18 \times (75 - 20) = 250 \times 4.18 \times 55 = 57,475 \text{ J or } 57.5 \text{ kJ}$

17. Heat lost by iron = Heat gained by water $-m(\text{iron})c(\text{iron})\Delta T(\text{iron}) = m(\text{water})c(\text{water})\Delta T(\text{water})$
 $-150 \times c \times (25.8 - 95) = 400 \times 4.18 \times (25.8 - 22) - 150 \times c \times (-69.2) = 400 \times 4.18 \times 3.8$
 $10,380c = 6353.6$
 $c = 0.612 \text{ J/g}^\circ\text{C}$ (close to the actual value of $\sim 0.45 \text{ J/g}^\circ\text{C}$; difference due to rounding and heat losses)

18. $q = mc\Delta T = 80 \times 4.18 \times (15 - 60) = 80 \times 4.18 \times (-45) = -15,048 \text{ J or } -15.0 \text{ kJ}$ (negative indicates heat released)

19. $q = mc\Delta T$, so $\Delta T = q/(mc) = -2500/(200 \times 0.90) = -2500/180 = -13.9^\circ\text{C}$ Final temperature = $150 + (-13.9) = \mathbf{136.1^\circ\text{C}}$

Part E: Enthalpy Changes

20.

- a) Exothermic (ΔH is negative)
- b) Heat released = $3.5 \text{ mol} \times 890 \text{ kJ/mol} = \mathbf{3115 \text{ kJ}}$
- c) Moles of $\text{CH}_4 = 32 \text{ g} \div 16 \text{ g/mol} = 2 \text{ mol}$ Heat released = $2 \text{ mol} \times 890 \text{ kJ/mol} = \mathbf{1780 \text{ kJ}}$

21. Moles of $\text{NH}_4\text{NO}_3 = 40 \text{ g} \div 80 \text{ g/mol} = 0.5 \text{ mol}$ Heat absorbed = $0.5 \text{ mol} \times 25.7 \text{ kJ/mol} = \mathbf{12.85 \text{ kJ}}$ or $\mathbf{12.9 \text{ kJ}}$

22. Moles of ethanol = $2.50 \text{ g} \div 46 \text{ g/mol} = 0.0543 \text{ mol}$ Molar enthalpy = $74.5 \text{ kJ} \div 0.0543 \text{ mol} = \mathbf{1371 \text{ kJ/mol}}$ or $\mathbf{1370 \text{ kJ/mol}}$ (Since heat is released, $\Delta H_c = \mathbf{-1370 \text{ kJ/mol}}$)

Part F: Hess's Law

23. We want: $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$

Given equations:

1. $\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g}) \Delta H_1 = +68 \text{ kJ}$
2. $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g}) \Delta H_2 = -114 \text{ kJ}$

Reverse equation 2: $2\text{NO}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \Delta H = +114 \text{ kJ}$

Now add equation 1 and reversed equation 2: $\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g}) \Delta H_1 = +68 \text{ kJ}$
 $2\text{NO}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \Delta H = +114 \text{ kJ}$

$\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) \Delta H = \mathbf{+182 \text{ kJ}}$

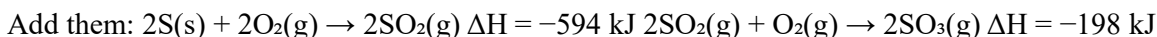
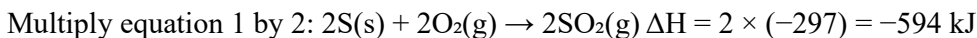
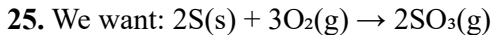
24. We want: $2\text{C}(\text{s}) + 3\text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})$

Reverse equation 1: $2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_2\text{H}_6(\text{g}) + 7/2\text{O}_2(\text{g}) \Delta H = +1560 \text{ kJ}$

Multiply equation 2 by 2: $2\text{C}(\text{s}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) \Delta H = 2 \times (-394) = -788 \text{ kJ}$

Multiply equation 3 by 3: $3\text{H}_2(\text{g}) + 3/2\text{O}_2(\text{g}) \rightarrow 3\text{H}_2\text{O}(\text{l}) \Delta H = 3 \times (-286) = -858 \text{ kJ}$

Add all three: $2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_2\text{H}_6(\text{g}) + 7/2\text{O}_2(\text{g}) \Delta H = +1560 \text{ kJ}$
 $2\text{C}(\text{s}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) \Delta H = -788 \text{ kJ}$
 $3\text{H}_2(\text{g}) + 3/2\text{O}_2(\text{g}) \rightarrow 3\text{H}_2\text{O}(\text{l}) \Delta H = -858 \text{ kJ}$



Part G: Calorimetry

26. $q(\text{water}) = mc\Delta T = 1500 \times 4.18 \times (32.5 - 20.0) = 1500 \times 4.18 \times 12.5 = \mathbf{78,375 \text{ J or } 78.4 \text{ kJ}}$

Since $-q(\text{reaction}) = q(\text{water})$, the combustion released **78.4 kJ**

27. Heat lost by metal = Heat gained by water
 $-m(\text{metal})c(\text{metal})\Delta T(\text{metal}) = m(\text{water})c(\text{water})\Delta T(\text{water})$
 $-50.0 \times c \times (26.4 - 100.0) = 200 \times 4.18 \times (26.4 - 22.0)$
 $-50.0 \times c \times (-73.6) = 200 \times 4.18 \times 4.4$
 $3680c = 3678.4$
 $c = \mathbf{1.00 \text{ J/g}^\circ\text{C}}$

28. Total volume = 200 mL, so mass = 200 g
 $q = mc\Delta T = 200 \times 4.18 \times (28.2 - 21.5) = 200 \times 4.18 \times 6.7 = 5601.2 \text{ J}$

Moles of water formed = $0.1 \text{ L} \times 1.0 \text{ M} = 0.1 \text{ mol}$ (since HCl and NaOH react 1:1)

$\Delta\text{H}(\text{neutralization}) = -5601.2 \text{ J} / 0.1 \text{ mol} = \mathbf{-56,012 \text{ J/mol or } -56.0 \text{ kJ/mol}}$ (Negative because heat is released in neutralization)

Part H: Application and Analysis

29. Method 1 (direct): $\Delta U = q - w = 250 - (-80) = 250 + 80 = \mathbf{330 \text{ J}}$

Method 2 (two steps): Step 1: $\Delta U_1 = 150 - 30 = 120 \text{ J}$ Step 2: $\Delta U_2 = 100 - (-110) = 100 + 110 = 210 \text{ J}$
 Total: $\Delta U = 120 + 210 = \mathbf{330 \text{ J}}$

Both methods give the same ΔU , demonstrating that internal energy is a state function—it depends only on the initial and final states, not on the path taken.

30. The standard enthalpy of formation is defined as the enthalpy change when one mole of a compound is formed from its elements in their standard states. For $\text{O}_2\text{(g)}$, it's already in its standard elemental form, so no formation reaction is needed, and by definition $\Delta\text{H}_f^\circ = 0$. For $\text{O}_3\text{(g)}$, the formation reaction is $3/2\text{O}_2\text{(g)} \rightarrow \text{O}_3\text{(g)}$, which requires energy input (it's endothermic), so ΔH_f° of ozone is positive (+142 kJ/mol).